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Interaction of Pyridine- and 4-Hydroxypyridine-2,6-dicarboxylic Acids with Heavy Metal Ions in Aqueous Solutions

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ABSTRACT: *Interactions between pyridine-2,6-dicarboxylic acid and 4-hydroxypyridine-2,6-dicarboxylic acid with Cu(II), Pb(II), and Cd(II) ions were characterized in aqueous solutions (20°C; $I = 0.4$ (KNO_3)) by means of dc-polarography. In solutions with excess of ligand, Cu(II), Pb(II), and Cd(II) form 1:2 complexes with the tridentate dianion of pyridine-2,6-dicarboxylic acid ($dipic^{2-}$) from weak acid to alkaline solutions. The values of $\log \beta_2$ for Cu(II), Pb(II), and Cd(II) are 16.1, 11.8, and 11.0, respectively. The complexing ability of pyridine-2,6-dicarboxylic acid is higher in acid solutions and lower in alkaline solutions than that of 4-hydroxypyridine-2,6-dicarboxylic acid. This difference is attributed to the OH-group, which can deprotonate in basic pH. In acid solutions the OH-group acts as an electron acceptor and reduces the electron donation available to the nitrogen atom in 4-hydroxypyridine-2,6-dicarboxylic acid, whereas in alkaline solutions the OH-group is deprotonated, and the deprotonated O^- group acts as an electron donor and increases the coordination ability of the ligand. The triple-deprotonated anion of 4-hydroxypyridine-2,6-dicarboxylic acid ($chel^{3-}$) forms a stable diligand complex with Cu(II), the stability constant logarithm being 21.5 ± 0.2.* Heteroatom Chem 14:625–632, 2003; DOI 10.1002/hc.10203

INTRODUCTION

Many transition and heavy metal cations play an active role in a great number of biological processes, being components of several vitamins and drugs. Pyridinedicarboxylic acids and their derivatives belong to an interesting series of compounds with biological applications [1]. Pyridine-2,6-dicarboxylic acid (dipicolinic acid) is present in nature as an oxidative degradation product of vitamins, coenzymes, and alkaloids and is a component of fulvic acids. It has frequently been cited in the literature as a plant-sterilizing and water-germicidal agent, and an antioxidant for ascorbic acids in foods [2]. Pyridine-2,6-dicarboxylic acid is almost unique to bacterial spores and may constitute as much as 15% of their weight [3]. Pyridine-2,6-dicarboxylic acid is a desirable metal ion ligand because of its low toxicity and amphophilic nature. Pyridine-3-carboxylic acid (commonly known as niacin or vitamin B_3), which is closely related to pyridine-2,6-dicarboxylic acid, is a precursor for the coenzyme NAD and is required in the human diet. Pyridine-2,6-dicarboxylic acid is furthermore related to pyridine-2,3-dicarboxylic acid (quinolinic acid), which also is an intermediate in the tryptophan degradation pathway and is a precursor of NAD [4].

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The interaction of transition and heavy metal ions with naturally occurring ligands in living organisms such as pyridine-2,6-dicarboxylic acid (H_2dipic) or its isomer 4-hydroxypyridine-2,6-dicarboxylic acid (chelidamic acid) (H_3chel) is important to evaluate potential beneficial and deteriorative effects of these ions.

Most studies have focused on the synthesis and structural investigations of metal H_2dipic complexes in solid state. Much is known on the preparation, crystalline structure, and properties of the Cu(II) [5–8], Fe(II) [9], Fe(III) [9–12], Mn(II) [6,13], Mn(III) [14,15], Ag(I) [1,16], Co(II) [5,6], Co(III) [17], Pt(II) [18], Ti(III) [19], Ni(II) [5,6], and Zn(II) [5,6] complexes. The metal-to-ligand stoichiometry 1:1 was found for complexes of Cu(II) [8] and Fe(II) [9], dinuclear and polymeric complexes were determined in the case of Ag(I) complexes [1,16], whereas 1:2 complexes were observed for Cu(II) [5–7], Fe(II) and Fe(III) [9–12], Mn(II) and Mn(III) [6,13–15], Co(II) and Co(III) [5,6,17], Pt(II) [18], Ti(III) [19], Ni(II) [5,6], and Zn(II) [5,6].

Most of the complexes listed above contain double deprotonated anion dipic^{2-}, whereas monodeprotonated Hdipic^{-} forms complexes with Co(II), Ni(II), Cu(II), Zn(II) [5], and Mn(II) [5,13]. In addition, the Cu(II) complex containing one unit of protonated H_2Dipic and one unit of double deprotonated dipic^{2-} was synthesized [7].

Data on synthesis of metal ion complexes with 4-hydroxypyridine-2,6-dicarboxylic acid (H_3chel) are not numerous. The metal-to-ligand stoichiometry 1:3 was found for the complex of Cr(III), containing monodeprotonated anions of H_2chel^{-} [20].

Solution chemistry of H_2dipic has been investigated. The UV–vis and IR absorption and ^{13}C NMR spectra of pyridine-2,6-dicarboxylic acid and of all other possible isomers of pyridinedicarboxylic acid were determined [21]. The pK_a values for H_2dipic had previously been reported [22] and the following average values equal to 2.2 and 4.6 were obtained for pK_{a1} and pK_{a2}, respectively, documenting that in neutral and alkaline solutions H_2dipic exists as dianion dipic^{2-}. Investigations on electrochemical behavior of pyridine-2,6-dicarboxylic acid on hanging mercury drop electrode (HMDE) showed that H_2dipic is strongly adsorbed on the mercury electrode only in acidic solutions, i.e., double deprotonated anion dipic^{2-} does not adsorb on Hg [23].

Data on the aqueous solution speciation of metal ion complexes with double deprotonated pyridine-2,6-dicarboxylic acid (dipic^{2-}) are listed in Table 1. When analyzing the data presented in the table it is seen that most heavy metal ions form complexes with metal-to-ligand stoichiometry 1:1 as well as 1:2. The complexes formed are of medium stability, except with Hg(II), which forms a relatively stable complex, the stability constant logarithm being over 20 (Table 1).

TABLE 1 Stability Constants of Heavy Metal Ion Complexes with Pyridine-2,6-dicarboxylate (dipic^{2-}) in Aqueous Solutions

	$\log \beta_1$	$\log \beta_2$	
Cu(II)	10.0	16.3	[24]
	9.14	16.52	[25]
	9.14	16.51	[26]
	8.9	17.1	[27]
	8.88	16.17	[28]
		16.4	[29]
Pb(II)	5.1	8.2	[30]
	8.70	11.60	[25]
	8.66	11.55	[31]
	8.69	11.57	[32]
Cd(II)	5.7	10.0	[30]
	6.75	11.11	[25]
	6.51	10.77	[33]
Fe(II)	5.71	10.36	[25]
Fe(III)	10.91	17.13	[34]
		16.74	[35]
Co(II)	7.0	12.5	[24]
	6.65	12.70	[25]
Ni(II)	8.0	14.1	[24]
	6.95	13.50	[25]
Zn(II)	7.0	13.0	[24]
	6.35	11.88	[25]
Mn(II)	5.01	8.49	[25]
Hg(II)		20.28	[25]
Pd(II)	16.0		[36]
VO_2^+	15.79		[37]

Interaction of 4-hydroxypyridine-2,6-dicarboxylic acid with metal ions in aqueous solutions has been less investigated (Table 2). Comparison of the stability constants obtained with that of metal ion complexes with pyridine-2,6-dicarboxylic acid shows that in general 4-hydroxypyridine-2,6-dicarboxylic

TABLE 2 Stability Constants of Metal Ion Complexes with 4-Hydroxypyridine-2,6-dicarboxylate (chel^{3-}) in Aqueous Solutions

	$\log \beta_1$	$\log \beta_2$	
Cu(II)	12.2	22.1	[38]
Co(II)	8.4	16.2	[38]
Mn(II)	6.7		[38]
Ni(II)	9.2	17.3	[38]
Zn(II)	9.3	17.8	[38]
Ba(II)	3.98		[39]
Ca(II)	5.40		[39]
Mg(II)	3.68		[39]
Sr(II)	4.34		[39]

acid forms more stable complexes with heavy metal ions than pyridine-2,6-dicarboxylic acid (cf. Tables 1 and 2). It can be mentioned that 4-hydroxypyridine-2,6-dicarboxylic acid has stronger acidic properties in comparison with pyridine-2,6-dicarboxylic acid, i.e., pK_{a1} and pK_{a2} values of the carboxylic groups are 1.4 and 3.1, respectively [22]. In addition, at higher pH the ionization of the OH-group of 4-hydroxypyridine-2,6-dicarboxylic acid occurs, pK_{a3} value being 10.9 [22].

Given the recent results showing that some transition metal complexes with the pyridine-2,6-dicarboxylic acid have beneficial effects in normalizing elevated blood glucose levels in diabetic rats [40,41], fundamental information on solution chemistry of metal complexes with pyridinedicarboxylic acid ligands is important. The present study was carried out to specify equilibria in solutions of pyridine-2,6-dicarboxylic and 4-hydroxypyridine-2,6-dicarboxylic acids in the presence of the divalent metal ions Cu(II), Pb(II), and Cd(II) in a wide pH range, and to compare the results with that obtained earlier for the isomer of the pyridine-2,6-dicarboxylic acid, i.e., the pyridine-2,5-dicarboxylic acid [42]. One of the other aims of the work presented herein was to determine the complexation level (i.e., concentration of free (uncomplexed) metal ions) of metal ions mentioned in the presence of pyridinecarboxylic acids.

RESULTS AND DISCUSSION

Metal Ion Complexes with Pyridine-2,6-dicarboxylic Acid

The quasi-reversible, diffusion-controlled (limiting current temperature coefficient values are in the range 1.1–1.4%/deg, and the dependencies of the limiting current on the square root of the mercury column height are linear) polarographic waves of Cu(II), Pb(II), and Cd(II) reduction in the presence of pyridine-2,6-dicarboxylic acid are observed in the wide pH range from 1.5 to 12.0 (Table 3) and allow calculations of the reversible half-wave potential $E^{r}_{1/2}$ and, subsequently, the equilibrium characteristics of the systems under investigation.

Cu(II)–Pyridine-2,6-dicarboxylic Acid

The actual half-wave potential $E_{1/2}$ of the polarographic waves of Cu(II) reduction shifts to more negative potentials with the increase in pH from 1.5 up to 6. The $E_{1/2}$ as well as $E^{r}_{1/2}$ remains practically constant with further increase in alkalinity up to pH 12.0 (Table 3).

TABLE 3 Characteristics of the Process of Metal Ion Reduction on a DME in Solutions of Pyridine-2,6-dicarboxylic Acid ($[M(II)]_0 = 0.5$ mM; $[H_2dipic]_0 = 5$ mM; 20°C; $I = 0.4$)

	pH	$-E^{r}_{1/2}$ *(mV)*	$-\Delta E^{r}_{1/2}$ *(mV)*
Cu(II)	1.5	59	124
	1.8	79	144
	2.1	107	172
	2.4	138	203
	2.7	155	220
	3.0	177	243
	3.3	197	262
	3.6	223	288
	5.0	256	231
	6.0	269	327
	8.0	270	328
	10.0	268	327
	12.0	267	326
Pb(II)	1.5	399	59
	2.0	429	88
	3.0	470	128
	4.0	501	159
	5.0	535	194
	6.0	545	204
	7.0	544	204
	8.0	546	205
	10.0	544	204
	11.0	562	221
	12.0	587	247
	13.0	638	296
Cd(II)	1.5	553	19
	1.8	558	24
	2.1	567	34
	2.4	584	51
	2.7	598	64
	3.0	615	82
	3.3	620	87
	3.6	647	113
	4.2	679	145
	4.5	687	153
	4.8	699	165
	5.1	704	170
	6.0	714	180
	8.0	713	179
	10.0	714	180
	12.0	715	181

The calculated concentration of the uncomplexed (free) Cu^{2+} ions in the solution under investigation diminishes from $10^{-7.6}$ to $10^{-14.6}$ upon increasing the pH from 1.5 to 6.0 (Fig. 1, curve 1) and evidences a rising of the complexing ability of H_2dipic because of its deprotonation (pK_{a1} and pK_{a2} are 2.2 and 4.6, respectively [22]). The further constancy of the pCu values at pH over 6.0 is in good agreement with the pK_{a2} value of pyridine-2,6-dicarboxylic acid. No evidence of additional Cu(II) complexation is observed at pH 6.0–12.0, i.e., no hydroxy complexes are formed in the pH region mentioned.

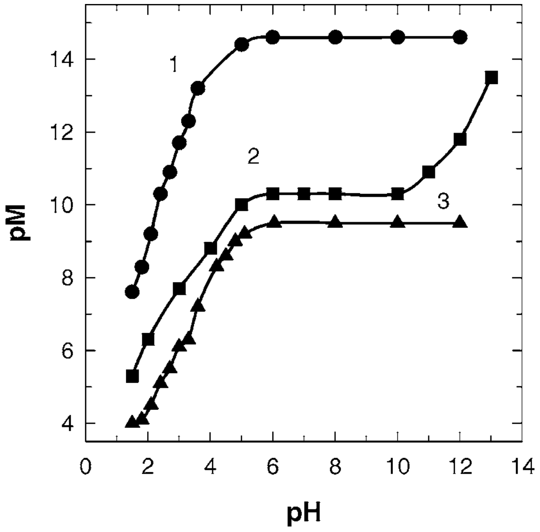

FIGURE 1 Dependence of the negative logarithm of free metal ion concentration (pM) on pH in pyridine-2,6-dicarboxylic acid (H_2dipic) containing solutions. Solution composition: $[M(II)]_0 = 0.5$ mM; $[H_2\text{dipic}]_0 = 5$ mM. **1**, Cu(II); **2**, Pb(II); **3**, Cd(II). 20°C; $I = 0.4$ (KNO_3).

The $\Delta E^{r}_{1/2}$ of Cu(II) reduction shifts by −61 mV with a 10-fold increase in ligand concentration (i.e., from 5 mM to 50 mM) at pH 10.0, suggesting participation of two double-deprotonated dipic^{2-} anions in complex formation. The value of the Cudipic$_2^{2-}$ complex stability constant logarithm calculated from data obtained at pH 6.0–12.0 is equal to 16.1 ± 0.1, showing Cu(II) complex formation of medium stability. The same mean log β_2 value (±0.4) was obtained from the data at pH 2.1–5.0, showing the Cudipic$_2^{2-}$ complex predominating in acid solution although protonated H_2dipic and Hdipic^{-} forms of pyridine-2,6-dicarboxylic acid predominate in solution bulk. Such a phenomenon—predominating of metal ion complex species with fully deprotonated anion of the ligand under the conditions of predominance of the protonated form of the ligand in the solution bulk—was observed earlier when investigating Cu(II) complex formation with polyaminopolycarboxylic acids (so-called complexones) [43–45].

The absorption maximum of Cudipic$_2^{2-}$ complex is at $\lambda = 775$ nm, the molar extinction coefficient being 60 l mol^{-1} cm^{-1} (Fig. 2, curve 2). The comparison of spectral characteristics of uncomplexed Cu(II) ions with those of the Cudipic$_2^{2-}$ complex shows the shift of the absorption maximum to shorter wavelengths from 815 to 775 nm and increased molar extinction coefficient from 12 to 60 l mol^{-1} cm^{-1} (Fig. 2, cf. curves 2 and 3).

Comparison of the results obtained with that of Cu(II)–pyridine-2,5-dicarboxylic acid system [48] shows significant differences. The stability of Cudipic$_2^{2-}$ complex is in two orders of magnitude higher than that of Cu(II)–pyridine-2,5-dicarboxylate complex of the same stoichiometry. The higher complexing ability of dipic^{2-} comparing with that of pyridine-2,5-dicarboxylate can be attributed to different denticity of ligands, i.e., pyridine-2,6-dicarboxylate is a tridentate ligand, whereas pyridine-2,5-dicarboxylate is didentate.

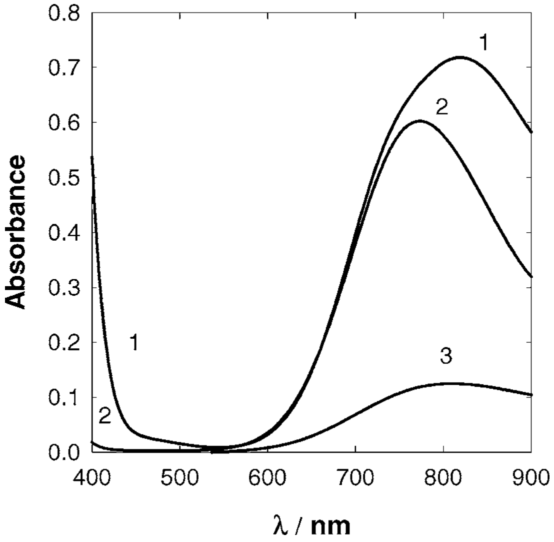

FIGURE 2 Light absorption spectra of Cu(II) in presence or absence of pyridinedicarboxylic acids. Solution composition: **1**−0.01 M $Cu(NO_3)_2$, 0.05 M 4-hydroxypyridine-2,6-dicarboxylic acid, pH 12.0; **2**−0.01 M $Cu(NO_3)_2$, 0.05 M pyridine-2,6-dicarboxylic acid, pH 10.0; **3**−0.01 M $Cu(NO_3)_2$, pH 4.0. $I = 0.4$ (KNO_3); 20°C; 1.0-cm path length quartz cells; pure water used as optical blanc solution.

The spectral characteristics of the complexes mentioned vary. The molar extinction coefficient value of Cudipic$_2^{2-}$ ($\varepsilon = 60$) is higher than that of Cu(II)–pyridine-2,5-dicarboxylate complex ($\varepsilon = 47$ [42]), absorption maximum being 775 and 650 nm [46], respectively.

Pb(II)–Pyridine-2,6-dicarboxylic Acid

The dependence of $\Delta E^{r}_{1/2}$ of the polarographic waves of Pb(II) reduction in the presence of pyridine-2,6-dicarboxylic acid on pH shows complex formation, being different from that of the Cu(II)–pyridine-2,6-dicarboxylate system (Table 3). The shape of dependencies is the same up to pH 10.0, $\Delta E^{r}_{1/2}$ values being more positive for Pb(II) (Table 3). Above pH 10 a difference is observed, i.e., for Pb(II) a shift of $\Delta E^{r}_{1/2}$ is observed with increasing pH, whereas $\Delta E^{r}_{1/2}$ remains constant in the Cu(II)–pyridine-2,6-dicarboxylate system.

The calculated concentrations of free Pb^{2+} ions reflects the changes in $\Delta E^{r}_{1/2}$ depending on solution pH (Fig. 1, curve 2). pPb decreases from $10^{-7.4}$ to $10^{-10.3}$ upon increasing pH from 1.5 to 6.0 (Fig. 1, curve 2) and remains constant from pH 6.0 to 10.0. The interpretation of the pPb–pH dependence is the same as for the Cu(II)–pyridine-2,6-dicarboxylate system. The further decrease in pPb at pH over 10.0 shows additional species form. The comparison of the pPb values at pH 10.0–13.0 with that obtained for $\mathrm{Pb(OH)_3^-}$ (i.e., in the precence of only $\mathrm{OH^-}$ ions) [46] shows similar pPb levels in the presence and in the absence of pyridine-2,6-dicarboxylate values consistent with the interpretation that Pb(II)–pyridine-2,6-dicarboxylate complex decomposes and converts to trihydroxyplumbate(II).

The $\Delta E_{1/2}^{\mathrm{r}}$ of Pb(II) reduction shifts by −60 mV with an increase in ligand concentration from 5 to 50 mM at pH 8.0, consistent with the coordination of two double-deprotonated $\mathrm{pdca^{2-}}$ anions in complex formation. The $\log\beta_2$ value of the $\mathrm{Pbdipic_2^{2-}}$ complex calculated from data obtained at pH 6.0–10.0 is equal to 11.8 ± 0.1 and is in the same range as previously determined [25,31,32]. The same mean $\log\beta_2$ value (±0.2) was obtained from the data at both pH 4.0 and 5.0, showing that the $\mathrm{Pbdipic_2^{2-}}$ complex is predominating in weak acid solutions.

Cd(II)–Pyridine-2,6-dicarboxylic Acid

The complexation of Cd(II) ions in pyridine-2,6-dicarboxylate containing solutions is least efficient from all metal ions investigated and ranges from $10^{-4.0}$ to $10^{-9.5}$ depending on pH (Fig. 1, curve 3).

The pattern of Cd(II) complex formation with pyridine-2,6-carboxylate is similar to that of Cu(II) complex formation, $E_{1/2}$, $E_{1/2}^{\mathrm{r}}$, $\Delta E_{1/2}^{\mathrm{r}}$ values being lower (Table 3).

The shift of $\Delta E_{1/2}^{\mathrm{r}}$ of Cd(II) reduction from −180 to −240 mV as the ligand increases from 5 to 50 mM at pH 10.0 shows that a 1:2 complex $\mathrm{Cddipic_2^{2-}}$ form, and the $\log\beta_2$ value was calculated to 11.0 ± 0.1 (from data used at pH 6.0–12).

Similar $\log\beta_2$ value (±0.3) was calculated from the data obtained at pH 1.8–5.1 and shows the $\mathrm{Cddipic_2^{2-}}$ complex predominate at pH over 1.5.

When comparing chelating affinities of stereoisomers, i.e., pyridine-2,6-dicarboxylate and pyridine-2,5-dicarboxylate, with regard to Cd(II) ions, under the same conditions (initial Cd(II) concentration 0.5 mM and total concentration of ligand 5 mM) at pH 5.0–11 the concentration free Cd(II) ions is equal to $10^{-9.5}$ and $10^{-5.4}$, respectively (cf. this work and [42]). The higher chelating affinity of pyridine-2,6-dicarboxylate ions can be explained by above-mentioned differences in denticity of ligands as well as by different stoichiometries of the complexes formed, i.e., Cd(II) forms diligand complex with pyridine-2,6-dicarboxylate, whereas in the case of pyridine-2,5-dicarboxylate the monoligand complex with $\log\beta_1 = 4.5$ is formed.

Diffusion Coefficients of Cu(II), Pb(II), and Cd(II) Complexes with Pyridine-2,6-dicarboxylate

The diffusion coefficients of 1:2 complexes of Cu(II), Pb(II), and Cd(II) with $\mathrm{dipic^{2-}}$ in 0.4 M $\mathrm{KNO_3}$ solution were calculated from polarographic data as described in Eq. (3). The diffusion coefficient values are listed in Table 4. The greatest diffusion coefficient was obtained for $\mathrm{Cddipic_2^{2-}}$, followed by $\mathrm{Pbdipic_2^{2-}}$ and $\mathrm{Cudipic_2^{2-}}$ species.

Diffusion coefficient value of Cd(II) ions in pyridine-2,6-dicarboxylate solutions (7.1×10^{-6} $\mathrm{cm^2\ s^{-1}}$, Table 4) is lower than that in pyridine-2,5-dicarboxylate solutions in the same conditions (7.8×10^{-6} $\mathrm{cm^2\ s^{-1}}$ [42]) and is in agreement with the complex stoichiometry.

Metal Ion Complexes with 4-Hydroxypyridine-2,6-dicarboxylic Acid

The diffusion-controlled, quasi-reversible polarographic waves of Pb(II) and Cd(II) reduction in the presence of 4-hydroxypyridine-2,6-dicarboxylic acid are observed only in acid solutions at pH 1.5–5.0 (Table 5). In the case of Cu(II) such waves are observed both at pH 1.5–5.0 and 12.0–13.0 (Table 5).

The $E_{1/2}$, $E_{1/2}^{\mathrm{r}}$, $\Delta E_{1/2}^{\mathrm{r}}$ as well as free metal ion concentration values are more positive in acidic 4-hydroxypyridine-2,6-dicarboxylate solutions than those obtained in pyridine-2,6-dicarboxylate solutions in the same conditions (cf. Figs. 1 and 3, and Tables 3 and 5). This suggests that the coordination ability of pyridine-2,6-dicarboxylic acid with OH-group is less. Since both ligands are tridentate and coordination of metal ions occurs in the same way through one nitrogen and two oxygen atoms, the

TABLE 4 Diffusion Coefficient Values of Metal Ion Species with Pyridine-2,6-dicarboxylate ($\mathrm{dipic^{2-}}$) Calculated from Polarographic Data ($I = 1.0$; 20°C; solution composition: $[\mathrm{M(II)}]_0 = 0.5$ mM, $[\mathrm{H_2dipic}]_0 = 5$ mM; predominating complex: $\mathrm{Mdipic_2^{2-}}$)

	pH	$D \times 10^6$ ($\mathrm{cm^2\ s^{-1}}$)
Cu(II)	6.0–12.0	5.2
Pb(II)	6.0–10.0	5.4
Cd(II)	6.0–12.0	7.1

TABLE 5 Characteristics of the Process of Metal Ion Reduction on a DME in Solutions of 4-Hydroxypyridine-2,6-dicarboxylic Acid ($[M(II)]_0 = 0.5$ mM; $[H_3chel]_0 = 5$ mM; 20°C; $I = 0.4$)

	pH	$-E^r_{1/2}$ *(mV)*	$-\Delta E^r_{1/2}$ *(mV)*
Cu(II)	1.5	−8	56
	2.0	30	94
	5.0	89	156
	12.0	418	481
	13.0	415	480
Pb(II)	1.5	364	22
	2.0	376	35
	3.0	385	43
	5.0	440	101
Cd(II)	1.5	542	8
	2.0	543	9
	3.0	558	25
	4.0	574	40
	5.0	580	48

lower complexing ability of 4-hydroxypyridine-2,6-dicarboxylic acid comparing with that of pyridine-2,6-dicarboxylic acid can be attributed to the influence of the substituent OH-group. The OH-group in 4-hydroxypyridine-2,6-dicarboxylic acid pulls electrons from the pyridine ring and reduces the coordination ability of this ligand.

In alkaline solutions the complexing ability of 4-hydroxypyridine-2,6-dicarboxylate is greater than that of pyridine-2,6-dicarboxylate. The $\Delta E^r_{1/2}$ value of Cu(II) reduction is ca. 150 mV more negative in the case of 4-hydroxypyridine-2,6-dicarboxylic acid at pH 12 (Tables 3 and 5). 4-Hydroxypyridine-2,6-dicarboxylic acid (H_3chel) exists in the solution as triple-deprotonated anion due to the deprotonation of OH-group in alkaline solutions ($pK_{a3} = 10.9$ [22]), and forms 1:2 complex with Cu(II), i.e., Cuchel$_2^{4-}$. The calculated stability constant ($\log \beta_2 = 21.5 \pm 0.2$) is high, and 4-hydroxypyridine-2,6-dicarboxylic acid can be attributed to the group of ligands forming stable complexes with Cu(II) ions, e.g., such as polyaminopolycarboxylic acids [43–45]. The diffusion coefficient calculated from the polarographic data for the complex equals 3.9×10^{-6} cm^2 s^{-1}.

The absorption maximum of Cuchel$_2^{4-}$ complex is at $\lambda = 820$ nm, molar extinction coefficient being 72 l mol^{-1} cm^{-1} (Fig. 2, curve 1). The λ_{max} is shifted to longer wavelengths with increase in molar extinction coefficient compared to Cudipic$_2^{2-}$ (cf. curves 1 and 2 in Fig. 2).

The stability of Cuchel$_2^{4-}$ and Cudipic$_2^{2-}$ discussed vary more than in four orders in magnitude, with the $\log \beta_2$ ranging from 21.5 to 16.1, respectively. The strengthening of 4-hydroxypyridine-2,6-dicarboxylic acid complexing ability in alkaline solutions comparing with that of pyridine-2,6-dicarboxylic acid can be explained by the influence of OH-group substitute. The deprotonated OH-group in 4-hydroxypyridine-2,6-dicarboxylic acid acts as a donor of electrons to the pyridine ring and increases the coordination ability of the ligand.

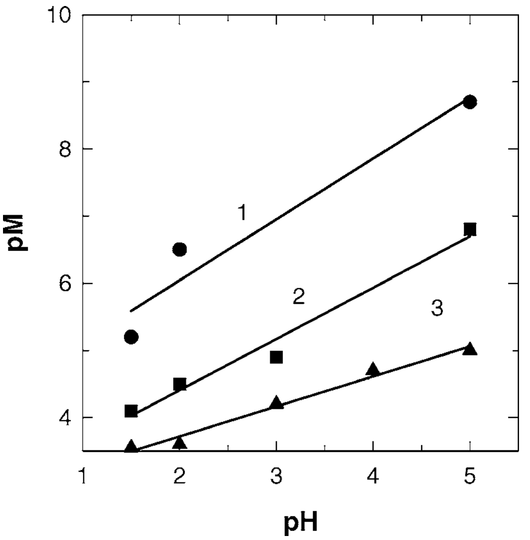

FIGURE 3 Dependence of the negative logarithm of free metal ion concentration (pM) on pH in 4-hydroxypyridine-2,6-dicarboxylic acid (H_3chel) containing solutions. Solution composition: $[M(II)]_0 = 0.5$ mM; $[H_3chel]_0 = 5$ mM. **1**, Cu(II); **2**, Pb(II); **3**, Cd(II). 20°C; $I = 0.4$ (KNO_3).

EXPERIMENTAL

Reagents

The pyridine-2,6-dicarboxylic acid (H_2dipic) was from Aldrich, and the 4-hydroxypyridine-2,6-dicarboxylic acid (H_3chel) was from TCI (Japan). Unless otherwise specified, analytical grade reagents were used. KNO_3 and KOH solutions were used to keep the ionic strength (I) of the solutions constant and equal to 0.4 M. The solutions were prepared in triple-distilled water.

Apparatus

The polarographic curves were recorded by a *PU–1* polarograph using a dropping mercury electrode in a thermostated three-electrode cell at 20 ± 0.1°C. The potential scanning rate was 100 mV min^{-1}. The capillary characteristics of the dropping mercury electrode (DME) were as follows: $m = 2.95$ mg s^{-1}, $t = 3.73$ s. The reference electrode was an Ag/AgCl electrode filled with saturated KCl. The solutions were deaerated by bubbling Ar through the solution before use.

The visible spectra were recorded with a Perkin-Elmer Lambda 35 UV–vis spectrometer at 20 ± 1°C in 1.0-cm path length quartz cells. The optical blank solution was pure water.

Procedures

The values of the actual half-wave potential $E_{1/2}$ were determined from a plot of $\log(\bar{i}/\bar{i}_{\mathrm{lim}} - \bar{i})$ against E. The values of the transition coefficient α and the reversible half-wave potential $E^{\mathrm{r}}_{1/2}$ were calculated using Eqs. (1) and (2) [47], where $E'_{1/2}$ is the half-wave potential determined by extrapolation of the linear part of the graph $\log(\bar{i}/\bar{i}_{\mathrm{lim}} - \bar{i})$ against E at sufficiently negative potentials, and $E_{1/11}$ is the potential where the current exceeds the value equal to 1/11 of the limiting current value. The diffusion coefficient values (D) were calculated using the Ilkovic Eq. (3) [48], where $\bar{i}_{\mathrm{dif}}$ is the limiting diffusion current (μA), n is the number of electrons involved in electrochemical reaction, c is the total concentration of Cu(II) ions (mM), m is the flow rate of Hg (mg s^{-1}), and t is the time of the drop formation (s). The corrections for a decrease in $\bar{i}_{\mathrm{lim}}$ were made in calculation of $\Delta E^{\mathrm{r}}_{1/2}$ for the case of complex formation as shown in Eq. (4) [48].

$$\alpha = 2.303\frac{RT}{nF}\left\{\frac{\Delta \log(\bar{i}/\bar{i}_{\mathrm{lim}} - \bar{i})}{\Delta(-E)}\right\}_{E\to-\infty}, \tag{1}$$

$$E^{\mathrm{r}}_{1/2} = E_{1/11} - 2.303\frac{RT}{nF} \times \left\{\log\left[10 - \exp\frac{\alpha nF}{RT}(E_{1/11} - E'_{1/2})\right]\right\}, \tag{2}$$

$$D^{1/2} = \frac{\bar{i}_{\mathrm{dif}}}{607\, ncm^{2/3}\, t^{1/6}}, \tag{3}$$

$$\Delta E^{\mathrm{r}}_{1/2} = \left(E^{\mathrm{r}}_{1/2}\right)_{\mathrm{compl}} - \left(\Delta E^{\mathrm{r}}_{1/2}\right)_{\mathrm{free}} - 2.303\frac{RT}{nF}\log\sqrt{\frac{\bar{i}_{\mathrm{lim\,free}}}{\bar{i}_{\mathrm{lim\,compl}}}} \tag{4}$$

Concentrations of free (uncomplexed) M^{2+} ions were calculated from $\Delta E^{\mathrm{r}}_{1/2}$ as shown in Eq. (5), where $[\mathrm{M(II)}]_0$ is the total M^{2+} concentration. The stability constants (β) of M(II) complexes with pyridinedicarboxylic acids were obtained during iterative approximation by minimizing the least-squares functional obtained from Eq. (6), where [L] is an equilibrium concentration of the ligand form taking part in complex formation (M), and a is an activity of OH^- ions. The stability constants were calculated assuming that the determination of the reversible half-wave potential was within ±2 mV. The accuracy of the stability constants can be calculated using the ±2 mV limit. Species were fit to the polarographic data, but if the calculation diverged to infinity, the species was not considered further.

$$\mathrm{pM} = nF/2.303RT\left(-\Delta E^{\mathrm{r}}_{1/2}\right) - \log[\mathrm{M(II)}]_0, \tag{5}$$

$$\sum_{x_i=0}^{N} \beta_{x_i}[\mathrm{L}^{n-}]^{x_i} = \exp\left[(nF/RT)\left(-\Delta E^{\mathrm{r}}_{1/2}\right)\right] - 1 \tag{6}$$

The molar extinction coefficient values of the individual complexes were calculated at $\lambda_{\max}$ according to Eq. (7), where c is the concentration (M) of the complexes, $A_{\lambda\max}$ is the light absorbance at wavelength $\lambda_{\max}$, and l is the cell width (cm).

$$\varepsilon_{\lambda\max} = A_{\lambda\max}/cl \tag{7}$$

CONCLUSIONS

1. In solutions with excess of ligand, Cu(II), Pb(II), and Cd(II) form 1:2 complexes with the tridentate dianion of pyridine-2,6-dicarboxylic acid (dipic^{2-}) from weak acids to alkaline solutions. The values of $\log\beta_2$ for Cu(II), Pb(II), and Cd(II) are 16.1, 11.8, and 11.0, respectively. Tenfold excess of pyridine-2,6-dicarboxylic acid reduces the concentration of free Cu(II), Pb(II), and Cd(II) ions from 1 to 11 orders in magnitude, depending on pH and stability constant of the complex formed. The values of the diffusion coefficients for $\mathrm{Cu(dipic)}_2^{2-}$, $\mathrm{Pb(dipic)}_2^{2-}$, and $\mathrm{Cd(dipic)}_2^{2-}$ are equal to 5.4×10^{-6}, 5.2×10^{-6}, and 7.1×10^{-6} $\mathrm{cm}^2\ \mathrm{s}^{-1}$, respectively.
2. Pyridine-2,6-dicarboxylic acid forms more stable complexes with metal ions than does pyridine-2,5-dicarboxylic acid. The higher complexing ability of pyridine-2,6-dicarboxylate compared to that of pyridine-2,5-dicarboxylate can be attributed to different denticity of ligands, i.e., pyridine-2,6-dicarboxylate acts as tridentate ligand, whereas pyridine-2,5-dicarboxylate is didentate.
3. The complexing ability of 4-hydroxypyridine-2,6-dicarboxylic acid is lower in acid solutions and higher in alkaline solutions than that of pyridine-2,6-dicarboxylic acid. This difference is attributed to the different actions of substituent, i.e., OH-group, depending on solution pH. In acid solutions the OH-group acts as an electron acceptor and decreases the electron donation available to the nitrogen atom in 4-hydroxypyridine-2,6-dicarboxylic acid, whereas in alkaline solutions the OH-group is deprotonated, and the

deprotonated O^- group acts as electron donor and increases the coordination ability of the ligand.

4. Triple-deprotonated anion of 4-hydroxypyridine-2,6-dicarboxylic acid ($chel^{3-}$) forms a stable diligand complex with Cu(II), the stability constant logarithm being 21.5 ± 0.2. The diffusion coefficient for $Cuchel_2^{4-}$ complex is 3.9×10^{-6} cm^2 s^{-1}.

REFERENCES

[1] Brzyska, W.; Ozga, W. Pol J Chem 1997, 71, 436.
[2] Ferraro, S.; Passamonti, P.; Bartocci, V.; Pucciarelli, F. J Chem Soc, Faraday Trans 1997, 93, 289.
[3] Davis, B. D.; Dulbecco, R.; Eisen, H. N.; Ginsberg, H. S.; Wood, W. B. Microbiology, 2nd ed.; Harper & Row: Hagerstown, 1973.
[4] Crans, D. C.; Yang, L.; Jakusch, T.; Kiss, T. Inorg Chem 2000, 39, 4409.
[5] D'Ascenzo, G.; Marino, A.; Sabbatini, M.; Bica, T. Thermochim Acta 1978, 25, 325.
[6] Nathan, L. C.; Zapien, D. C.; Mooring, A. M.; Doyle, C. A.; Brown, J. A. Polyhedron 1989, 8, 745.
[7] Sileo, E. E.; Blesa, M. A.; Rigotti, G.; Rivero, B. E.; Castello, E. E. Polyhedron 1996, 15, 4531.
[8] Mojumdar, S. C.; Hudecova, D.; Melnik, M. Pol J Chem 1999, 73, 759.
[9] Laine, P.; Gourdon, A.; Launay, J.-P. Inorg Chem 1995, 34, 5129.
[10] Williams, N. H.; Yandell, J. K. Aust J Chem 1983, 36, 2377.
[11] Ellis, R. M.; Quilligan, J. D.; Williams, N. H.; Yandell, J. K. Aust J Chem 1989, 42, 1.
[12] Hseu, J. F.; Chen, J. J.; Chuang, C. C.; Wei, H. H.; Cheng, M. C.; Wang, Y.; Yao, Y. D. Inorg Chim Acta 1991, 184, 1.
[13] Richter, S. A.; Tsang, P. K. S.; Sawyer, D. T. Inorg Chem 1988, 27, 1814.
[14] Yamaguchi, K.; Sawyer, D. T. Inorg Chem 1985, 24, 971.
[15] Limburg, J.; Brudvig, G. W.; Crabtree, R. H. J Am Chem Soc 1997, 119, 2761.
[16] Smith, G.; Reddy, A. N.; Byriel, K. A.; Kennard, C. H. L. Polyhedron 1994, 13, 2425.
[17] Ali, M.; Saha, S. K.; Banerjee, P. Bull Chem Soc Jpn 1991, 64, 2497.
[18] Chessa, G.; Marangoni, G.; Pitteri, B.; Bertolasi, V.; Gilli, G.; Ferretti, V. Inorg Chim Acta 1991, 185, 201.
[19] Bertram, H.; Wieghardt, K. Inorg Chem 1979, 18, 1799.
[20] Chetterjee, B. J Inorg Nucl Chem 1981, 43, 2553.
[21] Wasylina, L.; Kucharska, E.; Weglinski, Z.; Puszko, A. Khim Geterotsikl Soedin 1999, 210.
[22] Serjeant, E. P.; Dempsey, B. Ionisation Constants of Organic Acids in Aqueous Solutions; Pergamon: Oxford, 1979.
[23] Ferraro, S.; Passamonti, P.; Bartocci, V.; Pucciarelli, F. J Chem Soc, Faraday Trans 1997, 93, 289.
[24] Tichane, R. M.; Bennett, W. E. J Am Chem Soc 1957, 79, 1293.
[25] Anderegg, A. Helv Chim Acta 1960, 43, 414.
[26] Grenthe, I. J Am Chem Soc 1961, 83, 360.
[27] Faucherre, J.; Petitfaux, C.; Charlier, B. Bull Soc Chim Fr 1967, 1091.
[28] Grenthe, I.; Hanson, E. Acta Chem Scand 1969, 23, 611.
[29] Petitfaux, C.; Fournaise, R. Bull Soc Chim Fr 1972, 914.
[30] Suzuki, K.; Yamasaki, K. Naturwissenschaften 1957, 44, 396.
[31] Campanella, L.; De Angelis, G.; Napoli, A. Bull Soc Chim Belg 1972, 81, 489.
[32] Soares, H. M. V. M.; Almeida, A. A. N.; Castro, M. P. O.; Pinho, S. C.; Vasconcelos, M. T. S. D. Analyst 1998, 123, 1377.
[33] Evtimova, B.; Scharff, J.-P.; Paris, M. R. Bull Soc Chim Fr 1969, 81.
[34] Anderegg, G. Helv Chim Acta 1960, 43, 1530.
[35] Timberlake, C. F. J Chem Soc 1964, 1229.
[36] Kragten, J.; Decnop-Weever, L. G. Talanta 1980, 27, 685.
[37] Crans, D. C.; Yang, L.; Jakusch, T.; Kiss, T. Inorg Chem 2000, 39, 4409.
[38] Anderegg, G. Helv Chim Acta 1963, 46, 1011.
[39] Blasius, E.; Brozio, B. Ber Bunsen-Ges Phys Chem 1964, 68, 52.
[40] Crans, D. C. J Inorg Biochem 2000, 80, 123.
[41] Yang, L. Q.; Crans, D. C.; Miller, S. M.; la Cour, A.; Anderson, O. P.; Kaszynski, P. M.; Godzala, M. E.; Austin, L. D.; Willsky, G. R. Inorg Chem 2002, 41, 4859.
[42] Norkus, E.; Stalnionienė, I.; Crans, D. C. Cheminė technologija (Kaunas) 2002, N4, 9.
[43] Norkus, E.; Vaškelis, A. Chemija (Vilnius) 1995, N4, 19.
[44] Norkus, E.; Vaškelis, A.; Žakaitė, I. Talanta 1995, 43, 465.
[45] Norkus, E.; Pauliukaitė, R. Pol J Chem 2000, 74, 1231.
[46] Norkus, E.; Grincienė, G.; Vaitkus, R. Carbohydr Res 2002, 337, 1657.
[47] Matsuda, H.; Ayabe, Y. Z Elektrochem 1959, 63, 1164.
[48] Heyrovsky, J.; Kuta, J. Principles of Polarography; Academic Press: New York, 1966.