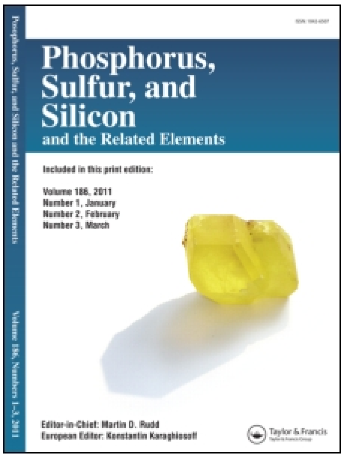

Phosphorus, Sulfur, and Silicon and the Related Elements
Publication details, including instructions for authors and subscription information:
http://www.informaworld.com/smpp/title~content=t713618290

Metal Complexes of Ethane and Propane Frame-Substituted Oligophosphonic and Oligophosphonocarboxylic Acids

Cs. Bogdán[a]; G. Péczely[b]; G. Hägele[c]
[a] Institute of Chemistry, Eötvös Loránd University, Budapest, Hungary [b] Department of Inorganic Chemistry, University of Pécs, Hungary [c] Department of Inorganic and Structural Chemistry, Heinrich Heine University, Düsseldorf, Germany

To cite this Article Bogdán, Cs. , Péczely, G. and Hägele, G.(2007) 'Metal Complexes of Ethane and Propane Frame-Substituted Oligophosphonic and Oligophosphonocarboxylic Acids', Phosphorus, Sulfur, and Silicon and the Related Elements, 182: 10, 2337 — 2350

To link to this Article: DOI: 10.1080/10426500701441465

URL: http://dx.doi.org/10.1080/10426500701441465

Phosphorus, Sulfur, and Silicon, 182:2337–2350, 2007

ISSN: 1042-6507 print / 1563-5325 online
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Metal Complexes of Ethane and Propane Frame-Substituted Oligophosphonic and Oligophosphonocarboxylic Acids

Cs. Bogdán[*]
Institute of Chemistry, Eötvös Loránd University, Budapest, Hungary

G. Péczely
Department of Inorganic Chemistry, University of Pécs, Hungary

G. Hägele
Department of Inorganic and Structural Chemistry, Heinrich Heine University, Düsseldorf, Germany

*The stability constants of the complexes of one oligophosphonic acid (ethane-1,1,2-triphosphonic acid **1**), five phosphonocarboxylic acids (ethane-1,2-diphosphono-1-carboxylic acid **2**, propane-1,1-diphosphono-2,3-dicarboxylic acid **3**, ethane-1,2-diphosphono-1,2-dicarboxylic acid **4**, ethane-1-phosphono-1,2-dicarboxylic acid **5**, and propane-1-phosphono-1,2,3-tricarboxylic acid **6**), and ethane-1,1,2-tricarboxylic acid **7** with Cu^{2+}, Ni^{2+}, Zn^{2+}, Cd^{2+} and Ca^{2+} were determined by pH-metric titrations. The six ligands of type H_NL ($N = 4 - 6$) form mononuclear MLH_n ($n = 0, 1, \ldots, N–1$) complexes with chelate structure. The calcium complexes seem to be less stable than the transition metal complexes.*

Keywords Oligophosphonates; phosphonocarboxylates; protonated metal complexes; stability constant

INTRODUCTION

Organophosphonate and phosphonocarboxylate compounds are widely used in a broad variety of applications. Their ability to prevent

Received April 26, 2007; accepted April 29, 2007.

[*] Currently at KFKI Research Institute for Particle and Nuclear Physics, Budapest, Hungary.

Cs. Bogdán is indebted to Prof. Ferenc Gaizer (Prof. Emeritus of the Institute of Chemistry, Eötvös Loránd University, Budapest, Hungary), whose help, hints and stimulating suggestions contributed to the current work.

Address correspondence to Cs. Bogdán; KFKI Research Institute for Particle and Nuclear Physics of the Hungarian Academy of Sciences; Konkoly Thege Miklós út 29-33, Budapest H-1121, Hungary. E-mail: bogdan@rmki.kfki.hu

precipitation of calcium and magnesium salts at substoichiometric concentrations renders them effective in water treatment for scale inhibition.[1,2] Phosphonocarboxylates are used extensively in detergents as detergent builders and/or sequestering agents.[3] These materials are also applied as corrosion inhibitors in industrial cleaning and in peroxy bleach stabilization. The aqueous bath for corrosion-resistant steel coating contains organophosphonic acid, ester, salt and/or polymer.[4]

Organophosphonates are also applied in flame-resistant polymers,[5] as stabilizers or masking agents in photographic processing[6] and in actinide separation methods.[7] In the early 1990s, these compounds have been identified as promising candidates to create "structurally tailored" and microporous materials,[8,9] for catalysis[10] and for the electrochemical remediation of polluted soils.[11]

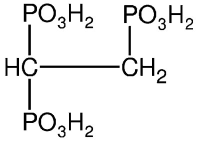
The potent biological activity of carboxyalkylphosphonates and aminoalkylphosphonates makes them useful as components of microfertilizers and pesticides in agriculture.[12]

Phosphonates play also a significant role in medicine and pharmacology. Arylphosphonates inhibit protein-tyrosine and serine/threonine phosphatases.[13] Phosphonocarboxylates show antiviral properties,[14] inhibit the intestinal and renal phosphate transport[15] as well as the sodium-dependent phosphate reabsorption.[16–18] Bisphosphonates inhibit bone reabsorption and are used as complexing agents for ^{99m}Tc for bone imaging.[19] Propane-1,1-diphosphono-2,3-dicarboxylic acid was found to play a major role in this field, since it stabilizes the radiopharmaceutical solutions.[20]

Development of technical applications of organophosphonates and phosphonocarboxylates requires reliable protonation and complex stability constants, that permits equilibrium modeling and prediction of the underlying molecular processes. Since most oligophosphonates resist biodegradation,[1] chemical speciation calculations and prediction of their fate in aqueous environment is of primary importance. Compared to phosphinates,[21] aminoalkylphosphonates[21,22] and alkyl phosphonocarboxylates,[21–23] considerably less attention has been devoted so far to the metal-binding properties of short-chain and oligophosphonocarboxylates. To the best of our knowledge, only in one publication[24] data on the Ca^{2+} complexes of propane-1,1-diphosphono-2,3-dicarboxylic acid (**3**) are reported: CaL, CaLH Ca_2LH and $CaLH_2$ complexes were substantiated in the pH > 4 interval at $T_{Ca}/T_L = 0.5$.

We have reported the protonation constants of 2 oligophosphonic and 5 phosphonocarboxylic acids: the ethane or propane frame-substituted polyacids **1–7** (see Table I for structures).[25] These ligands, containing several PO_3^{2-} and COO^- groups in vicinal and/or geminal positions are expected to be powerful complex-forming agents with a wide range of possible coordination modes.[24,26] pH-Metric titrations were performed

TABLE I Structures of the Polyacids 1–7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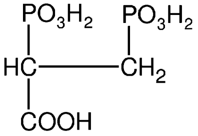
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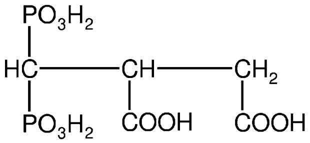
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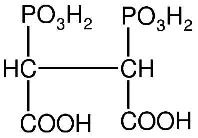
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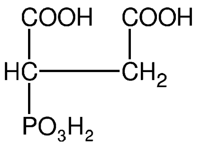
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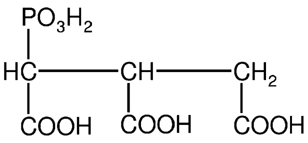
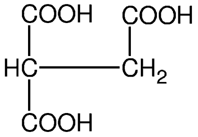

6


7

in order to determine the compositions and stability constants of their metal ion complexes. Transition metals of biochemical importance (Cu^{2+}, Zn^{2+}) or toxicity (Ni^{2+}, Cd^{2+}) were selected for our studies and Ca^{2+} was also investigated in view of possible practical (anti-scaling) applications.

EXPERIMENTAL

Materials and Syntheses

The polyacids were prepared in our laboratories by a variety of known methods, summarised briefly below. The effective concentration of all ligands was determined by potentiometric titrations against standardized NaOH and no protolytically active impurities could be detected. The $< 100\%$ purity data below result from the known observation, that hydrolysis of the alkyl esters of oligophosphonocarboxylic acids by aqueous HCl leads to highly hygroscopic parents acids.

Sodium diethyl phosphite reacted with vinylidene chloride in tetrahydrofurane solution at 100–105°C to form ethane-1,1,2-trisphosphonic acid hexaethyl ester in good yield, which was subsequently converted by hydrolysis with concentrated HCl to ethane-1,1,2-trisphosphonic acid (**1**, purity 93.6%, which corresponds to a monohydrate).[27] This procedure was termed as *Anomalous Michaelis-Becker* reaction, described in the patent literature[28] and briefly mentioned in two further publications.[29] For the sake of completeness, we want to point out, that a more complex route to compound **1** was reported by Saunders and Simpson.[30]

Sodium diethyl phosphite reacted with 2,3-dichloropropionic acid ethyl ester at reflux temperature in a Michaelis-Becker type reaction[31] to give 2,3-diphosphonopropionic acid pentaethyl ester (yield: 25.2%). It is interesting to note, that this reaction was first studied by Kamai and Kukhtin as early as 1957.[32] We found, that 2,3-diphosphonopropionic acid pentaethyl ester can also be prepared by Pudovik-type addition of two moles of $HP(O)(OEt)_2$ to H–C≡C–COOEt (see below). The pentaester was hydrolyzed by boiling in water for 5 h to give 2,3-diphosphonopropionic acid (**2**, purity: 97.4%).

Some phosphonocarboxylic acid esters were prepared by addition of diethyl phosphite to α-unsaturated carboxylic acid derivatives under alkaline (sodium ethylate) catalysis.[33] This synthetic route was pioneered by Pudovik.[34] Addition to acetylenedicarboxylic acid gave ethane-1,2-diphosphono-1,2-dicarboxylic acid hexa-ethyl ester in 10.3% yield, which was hydrolysed by conc. HCl to give ethane-1,2-diphosphono-1,2-dicarboxylic acid **4** in 98.4% yield.

Pudovik-type addition to diethyl maleate gave ethane-1-phosphono-1,2-dicarboxylic acid tetraethyl ester in 28% yield, the hydrolysis of which under acidic conditions led to ethane-1-phosphono-1,2-dicarboxylic acid **5** with a purity of 92.1%.

The propane-frame substituted phosphonocarboxylic acids **3** and **4** were prepared similarly.[33] In particular, the hexaethyl ester of propane-1,1-diphosphono-2,3-dicarboxylic acid **3** was synthesized by Michaelis-type addition of methylenediphosphonic acid tetraethyl ester to maleic acid diethyl ester in 38.4% yield. It was then hydrolysed to give propane 1,1-diphosphono-2,3-dicarboxylic acid **3** with a purity of 96.4%. Phosphonoacetic acid diethyl ester was reacted with maleic acid dibutylester to form the pentaester of propane-1-phosphono-1,2,3-tricarboxylic acid (yield: 40.3%). Its hydrolysis (in boiling water for 5 h), yielded propane-1-phosphono-1,2,3-tricarboxylic acid **6** with a purity of 97.6%. Both Pudovik-type reactions were conducted in the presence of sodium methylate at 100–110°C.

For the analytical studies presented below compound **4** was used as a genuine mixture of *meso* and *rac* forms, while **6** contained the *threo* and *erythro* forms. No attempts were made to separate the individual stereoisomers.

All other chemicals were obtained from Merck and were of analytical grade; bidistilled, carbonate-free water was used throughout this study.

Potentiometric Titrations

The pH-metric measurements were performed at 25 ± 0.1°C and a constant total ionic strength of 0.5 $\text{mol}\cdot\text{dm}^{-3}$, adjusted by $NaClO_4$. The titration apparatus included a glass electrode, a separate reference Ag/AgCl electrode with the inner KCl reference electrolyte replaced by 0.5 $\text{mol}\cdot\text{dm}^{-3}$ $NaClO_4$, and an OP-204 digital pH-meter with a precision of 0.1 mV (all from Radelkis, Hungary). 3–5 $\text{mmol}\cdot\text{dm}^{-3}$ polyacid, 3–10 $\text{mmol}\cdot\text{dm}^{-3}$ (for Ca^{2+}, 15–50 $\text{mmol}\cdot\text{dm}^{-3}$) of the corresponding metal perchlorate and 0.01 $\text{mol}\cdot\text{dm}^{-3}$ $HClO_4$ were titrated under a nitrogen atmosphere (99.9995% N_2 from Linde, washed with NaOH and $NaClO_4$ solutions) with a 0.1 $\text{mol}\cdot\text{dm}^{-3}$ NaOH titrant containing 0.4 $\text{mol}\cdot\text{dm}^{-3}$ $NaClO_4$ (for further details, see ref. [35]).

The titration datasets were evaluated with the least-squares MAXIPOT-F computer program.[36] This program refines the selected parameters by minimizing the sum of squared residuals in electromotive forces (emf), $[\sum(\Delta E)^2]$. In the course of the refinement, the optimum values for the decadic logarithm of the concentration stability constants, $\beta_{\text{qpr}} = [M_qL_pH_r][M]^{-q}[L]^{-p}[H]^{-r}$ along with their error estimates were calculated (ionic charges are throughout omitted). The previously determined cumulative protonation constants $\beta_{01i} = [H_iL][L]^{-1}[H]^{-i}$, listed in the rows H^+ and the columns of the corresponding *r* value in Table II, were held constant during the iterations. To account for differences in ligand donor group basicities during comparison of metal-binding affinities, basicity-adjusted stability constants[22] were calculated as [log β_{110} – $\beta_{01\text{n}}$], where *n* is the number of donor groups characterized by protonation constants. Similarly, the conditional stepwise stability constants corresponding to the reaction:

$$M + H_iL = MLH_i \qquad K_i = [MLH_i][M]^{-1}[H_iL]^{-1} \tag{1}$$

provide more reliable means to estimate the relative metal-binding ability of the protonated ligands than the log β values.[22]

RESULTS AND DISCUSSION

The pH-metric measurements were planned in the pH range 1.8–11. In certain systems, we could indeed carry out measurements up to pH∼11,

TABLE II **Brutto Stability Constants (log β_{pqr} with Standard Deviations of Last Digits as Lower Indices), Basicity-Adjusted (log $\beta_{110}^{\mathrm{adj}}$) and Conditional Stability Constants (log K_r) for the Metal Complexes $M_qL_pH_r$ (Both *in Italics*, for Definitions, see Experimental) and Cumulative Protonation Constants[25] log β_{01r} of the Ligands. All Equilibrium Constants at 25°C and $I = 0.5$ M**

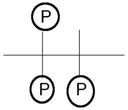
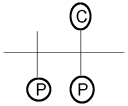
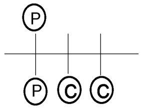
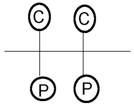
Compound, schematic formula[a]	Cation	T_M/T_L: pH range (NMP[b])	log β_{115} log K_5 (H$^+$: log β_{015})	log β_{114} log K_4 (H$^+$: log β_{014})	log β_{113} log K_3 (H$^+$: log β_{013})	log β_{112} log K_2 (H$^+$: log β_{012})	log β_{111} log K_1 (H$^+$: log β_{011})	log β_{110} log $\beta_{110}^{\mathrm{adj}}$	log β_{11-1}
1 	Cu^{2+}	1: 1.8–3.2 (55)	—	31.55_{64}	29.72_{41}	25.84_{59}	19.79	—	—
		2: 1.8–3.2 (52)		*2.54*	*3.76*	*5.88*	*8.66*		
	Ni^{2+}	1: 1.8–8.2 (138)	—	32.26_{8}	29.51_{6}	24.65_{10}	18.84_{9}	10.32_{20}	—
		2: 1.8–4.8 (58)		*3.25*	*3.55*	*4.69*	*7.71*	*−20.48*	
	Zn^{2+}	1: 1.8–3.7 (52)	—	32.34_{16}	29.87_{15}	~ 25	~22	—	—
		2: 1.8–3.7 (52)		*3.33*	*3.91*	*~5.0*	*~10.9*		
	Cd^{2+}	1: 1.8–3.5 (49)	—	31.37_{14}	29.39_{8}	24.40_{52}	—	—	—
		2: 1.8–3.0 (50)		*2.36*	*3.43*	*4.44*			
	Ca^{2+}	5: 1.8–5.4 (109)	—	30.52_{14}	28.02_{13}	22.33_{98}	16.62	—	—
		10: 1.8–4.7 (52)		*1.51*	*2.06*	*2.37*	*5.49*		
	H$^+$	—	(30.8)	(29.01)	(25.96)	(19.96)	(11.13)	—	—
2 	Cu^{2+}	1: 1.9–3.1 (39)	—	31.04_{15}	29.71_{11}	25.91_{12}	—	—	—
				6.27	*7.33*	*8.39*			
	Ni^{2+}	1: 2.0–6.1 (75)	—	31.91_{5}	29.54_{4}	24.96_{4}	19.58_{4}	12.65_{8}	~3.6
				7.14	*7.16*	*7.44*	*9.26*	*−12.12*	
	Zn^{2+}	1: 2.0–3.7 (41)	—	32.14_{3}	30.04_{3}	25.72_{8}	—	—	—
				7.37	*7.66*	*8.20*			
	Cd^{2+}	1: 2.0–2.9 (32)	—	31.22_{29}	29.04_{34}	—	—	—	—
				6.45	*6.66*				
	Ca^{2+}	5: 2.0–5.5 (55)	—	30.86_{38}	28.36_{41}	23.59_{57}	—	—	—
				6.09	*5.98*	*6.07*			
	H$^+$	—	—	(24.77)	(22.38)	(17.52)	(10.32)	—	—

TABLE II Brutto Stability Constants (log β_{pqr} with Standard Deviations of Last Digits as Lower Indices), Basicity-Adjusted (log $\beta_{110}^{\mathrm{adj}}$) and Conditional Stability Constants (log K_r) for the Metal Complexes $M_qL_pH_r$ (Both *in Italics*, for Definitions, see Experimental) and Cumulative Protonation Constants[25] log β_{01r} of the Ligands. All Equilibrium Constants at 25°C and $I = 0.5$ M *(Continued)*

Compound, schematic formula^a	Cation	T_M/T_L: pH range (NMP^b)	log β_{115} log K_5 (H$^+$: log β_{015})	log β_{114} log K_4 (H$^+$: log β_{014})	log β_{113} log K_3 (H$^+$: log β_{013})	log β_{112} log K_2 (H$^+$: log β_{012})	log β_{111} log K_1 (H$^+$: log β_{011})	log β_{110} log $\beta_{110}^{\mathrm{adj}}$	log β_{11-1}
3 	Cu^{2+}	1: 2.0–8.0 (43)	—	32.5_{7}	29.00_{8}	24.75_{11}	19.30_{25}	13.27_{13}	—
		2: 2.0–7.9 (43)		*3.37*	*4.10*	*5.39*	*7.86*	*−18.14*	
	Ni^{2+}	1: 2.0–10.5 (40)	—	32.01_{28}	28.91_{32}	24.09_{31}	18.43_{32}	11.19_{29}	—
		2: 2.0–10.5 (42)		*2.88*	*4.01*	*4.73*	*6.99*	*−20.22*	
	Zn^{2+}	1: 2.1–7.1 (37)	—	32.30_{15}	28.79_{13}	24.13_{29}	19.19_{24}	13.00_{35}	—
		2: 2.0–5.6 (43)		*3.17*	*3.89*	*4.77*	*7.75*	*−18.41*	
	Cd^{2+}	1: 2.0–4.6 (34)	—	32.74_{19}	29.98_{8}	25.19_{77}	20.64	—	—
				3.61	*5.08*	*5.83*	*9.20*		
	H$^+$	—	(31.41)	(29.13)	(24.9)	(19.36)	(11.44)	—	—
4 	Cu^{2+}	1: 1.9–5.1 (112)	—	33.76_{12}	29.36_{16}	24.52_{22}	18.62	—	—
		2: 1.9–4.4 (45)		*2.83*	*3.01*	*4.91*	*6.74*		
	Ni^{2+}	1: 1.9–4.9 (107)	—	33.66_{6}	29.37_{11}	24.28_{12}	18.44_{168}	—	—
		2: 1.9–4.9 (52)		*2.73*	*3.02*	*4.67*	*6.56*		
	Zn^{2+}	1: 1.9–6.7 (61)	—	33.94_{8}	29.95_{9}	24.90_{9}	18.19_{20}	—	3.88
		2: 1.9–6.1 (60)		*3.01*	*3.60*	*5.29*	*6.31*		
	Cd^{2+}	1: 1.9–5.3 (69)	—	33.39_{7}	29.21_{8}	23.91_{8}	—	—	—
		2: 1.8–6.3 (72)		*2.46*	*2.86*	*4.30*			
	Ca^{2+}	5: 1.9–9.9 (65)	—	32.32_{11}	28.10_{10}	21.74_{5}	14.31_{7}	—	—
		10: 1.9–10.3(68)		*1.39*	*1.75*	*2.13*	*2.43*		
	H$^+$	—	(34.34)	(30.93)	(26.35)	(19.61)	(11.88)	—	—

(Continued on the next page)

TABLE II Brutto Stability Constants (log β_{pqr} with Standard Deviations of Last Digits as Lower Indices), Basicity-Adjusted (log β_{110}^{adj}) and Conditional Stability Constants (log K_r) for the Metal Complexes $M_qL_pH_r$ (Both *in Italics*, for Definitions, see Experimental) and Cumulative Protonation Constants[25] log β_{01r} of the Ligands. All Equilibrium Constants at 25°C and $I = 0.5$ M *(Continued)*

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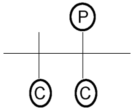
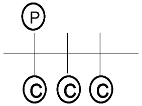
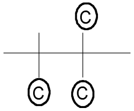
5 	Cu^{2+}	1: 2.3–5.6 (56)	—	—	28.24_6	25.25_5	21.17_4	16.45_3	10.61
		2: 2.1–4.2 (49)			*7.96*	*9.54*	*12.13*	*−6.35*	
	Ni^{2+}	1: 2.2–6.8 (69)	—	—	30.52_3	26.65_7	21.79_{13}	15.81_{35}	3.71
		2: 2.3–6.6 (61)			*10.24*	*10.94*	*12.75*	*−6.99*	
	Zn^{2+}	1: 2.3–4.7 (55)	—	—	29.58_8	25.68_7	20.79_{14}	15.31_{14}	7.46
		2: 2.0–4.7 (65)			*9.30*	*9.97*	*11.75*	*−7.49*	
	Cd^{2+}	1: 2.3–4.9 (38)	—	—	29.07_{21}	25.25_{18}	20.42_{27}	14.86_{27}	—
		2: 2.2–4.4 (41)			*8.79*	*9.54*	*11.38*	*−7.94*	
	Ca^{2+}	5: 2.2–8.0 (70)	–	–	27.72_1	23.73_1	18.41_2	11.33_2	1.65
					7.44	*8.02*	*9.37*	*−11.47*	
	H^+	—	—	(22.80)	(20.28)	(15.71)	(9.04)	—	—
6 	Cu^{2+}	1: 1.9–6.0 (56)	—	33.15_3	30.08_3	26.32_5	21.77_7	16.18_{10}	—
		2: 1.9–6.1 (60)		*10.28*	*10.70*	*11.54*	*13.22*	*−8.30*	
	Ni^{2+}	1: 2.2–8.5 (58)	—	32.49_5	29.15_5	24.82_9	19.48_{11}	13.12_{16}	2.49
		2: 1.9–8.3 (60)		*9.62*	*9.77*	*10.04*	*10.93*	*−11.36*	
	Zn^{2+}	1: 1.9–7.0 (62)	—	33.90_6	30.57_7	26.25_9	21.05_9	15.03_{11}	—
		2: 1.8–6.5 (56)		*11.03*	*11.19*	*11.47*	*12.50*	*−9.45*	
	Cd^{2+}	1: 1.9–5.3 (48)	—	33.05_7	29.92_5	25.56_7	20.47_7	14.31	—
				10.18	*10.54*	*10.78*	*11.92*	*−10.17*	
	H^+	—	(24.48)	(22.87)	(19.38)	(14.78)	(8.55)		—

TABLE II Brutto Stability Constants (log β_{pqr} with Standard Deviations of Last Digits as Lower Indices), Basicity-Adjusted (log β_{110}^{adj}) and Conditional Stability Constants (log K_r) for the Metal Complexes $M_qL_pH_r$ (Both *in Italics*, for Definitions, see Experimental) and Cumulative Protonation Constants[25] log β_{01r} of the Ligands. All Equilibrium Constants at 25°C and $I = 0.5$ M *(Continued)*

Compound, schematic formula[a]	Cation	T_M/T_L: pH range (NMP[b])	log β_{115} log K_5 (H$^+$: log β_{015})	log β_{114} log K_4 (H$^+$: log β_{014})	log β_{113} log K_3 (H$^+$: log β_{013})	log β_{112} log K_2 (H$^+$: log β_{012})	log β_{111} log K_1 (H$^+$: log β_{011})	log β_{110} log β_{110}^{adj}	log β_{11-1}
7 	Cu^{2+}	1: 1.9–5.0 (94)	—	—	—	12.30_5	9.47_2	4.78_4	—
		2: 1.9–4.8 (49)				*2.38*	*3.76*	*−7.54*	
		3: 1.9–4.7 (45)							
	Ni^{2+}	1: 2.0–4.5 (92)	—	—	—	—	5.13	2.94_3	—
		2: 2.0–4.8 (46)					*−0.58*	*−9.38*	
		3: 2.0–4.8 (48)							
	Zn^{2+}	1: 2.0–6.0 (96)	—	—	—	—	4.97	2.14_{10}	—
		2: 2.0–5.5 (48)					*−0.74*	*−10.18*	
		3: 2.0–5.4 (96)							
	Cd^{2+}	1: 2.2–5.6 (96)	—	—	—	—	—	2.36_{10}	—
		2: 2.0–5.6 (47)						*−9.96*	
		3: 2.0–5.6 (47)							
	Ca^{2+}	5:1 2.0–11.0 (93)	—	—	—	—	—	1.16_{21}	—
		10: 1.9–11.0 (92)						*−11.16*	
	H^+	—	—	—	(12.32)	(9.92)	(5.71)	—	—

[a] –P and –C denote $–PO_3H_2$ and –COOH groups, respectively; [b] number of measured points per titration curve.

but in others the emf became unstable even at pH~4. In these latter cases, the measurements were interrupted as the solutions became opalescent or a precipitate was formed. We could not recognize any clear-cut trend in connection with this phenomenon, but it seems that compounds containing vicinal carboxylate groups can be titrated to a higher pH range. On the other hand, the systems containing Ni^{2+} bear a less pronounced tendency to hydrolysis as compared to the systems with Cu^{2+}, Zn^{2+} or Cd^{2+}.

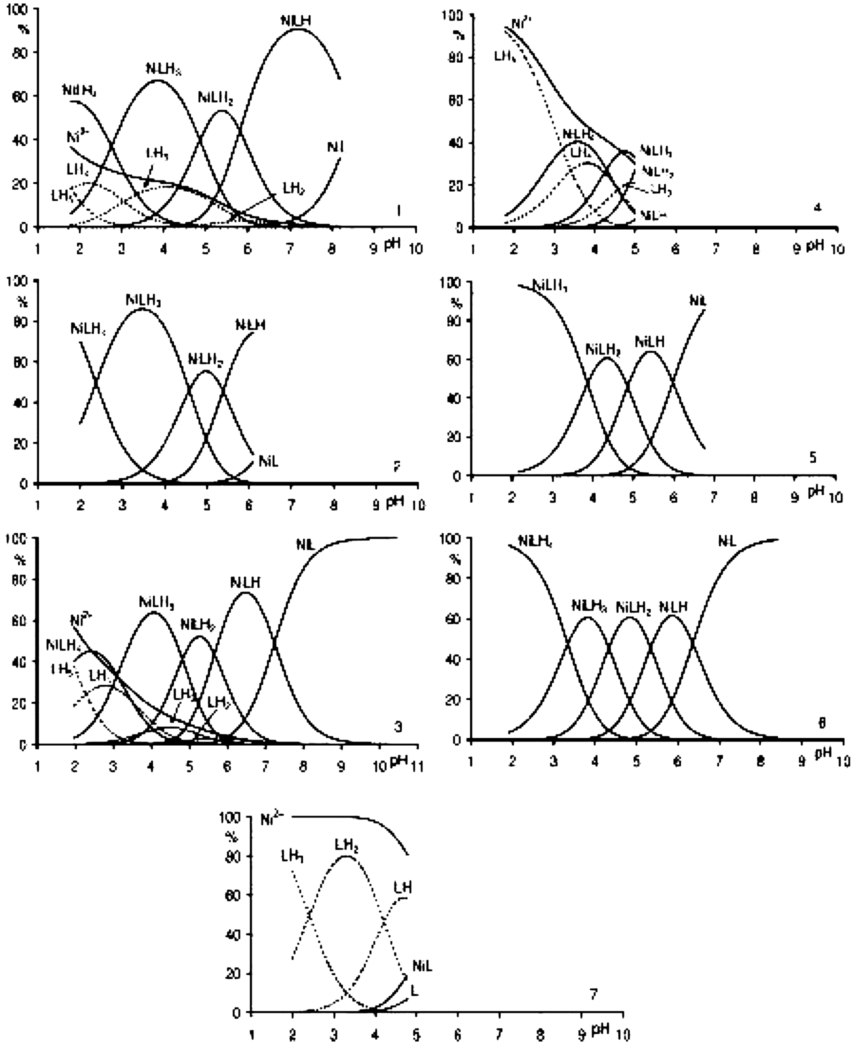
The log β values together with their errors obtained by joint computer evaluation of the parallel measurements are contained in Table II. The distribution curves for the Ni^{2+} complexes, calculated *via* the determined stability constants, are depicted as typical examples in Figure 1.

Although the ligands studied bear multiple functional groups allowing the binding of more than one metal ion and although the titrations were also performed at metal : ligand ratios higher than 1:1, the computer evaluation did not support evidence for the existence of polynuclear complexes.

ML complexes are typically formed at pH > 6, but complexes of adequate solubility were formed only with ligands holding more carboxylate than phosphonate groups. The double negative charge and the more basic character of the phosphonate group allows the formation of protonated complexes MLH_n ($n = 1$–4). As shown in Figure 1 for the Ni^{2+} complexes, the MLH_n species often become predominant in the pH range corresponding to the individual protonation step leading to LH_n, established previously[25] for the (binary) ligand-proton system. The solubility of phosphonate complexes presumably decreases upon charge neutralization concomitant with protonation. Several association processes, including the formation of colloids, can take place even at millimolar concentrations typical for potentiometric measurements,[26] which strongly affects the reliability of the obtained stability constants.

For metal ions with a greater tendency to hydrolysis (e.g. Cu^{2+} and Zn^{2+}), the formation of monohydroxo complexes could also be substantiated by the computer evaluation (see the log β_{11-1} values in Table II). In the Cd^{2+}-containing systems, precipitation occured even in the acidic region and no hydrolysed species could be detected in solution.

Considering the log β_{110} constants, the stability of the tricarboxylate **7** complexes is by far the lowest. This seems to be a primary consequence of the weaker basicity of the carboxylate group as compared to the phosphonate moiety. If the metal-binding affinity is corrected by the sum of the protonation constants (cf. the corresponding log β^{adj}_{110} values of ligands **7** and **6** in Table II), the loss in affinity greatly decreases (for Zn^{2+} or Cd^{2+}) or even reverses (for Ni^{2+} or Cu^{2+}).



SCHEME 1 Percentage distribution of Ni(II) in its different complexes and that of the total ligand in the Ni-free species as a function of pH. $T_L = T_{Ni(II)} = 5{\cdot}10^{-3}$ mol·dm^{-3}.

The stabilities of the *d*-metal complexes of the phosphonate compounds studied here vary in a wide range. The sequence of stability *with different metal ions* does not demonstrate a definite trend. Generally, the Ni^{2+}complexes exhibit a slightly lower stability than the other 3 metal ions.

Comparison of the relative complex stability *of different ligands* on the basis of uncomplexed metal ion concentration in the pH range up to 4–5 (plots not shown here) also reveals that the most stable complexes are formed with ligand **6**. Here the value of p_M ($p_M = -\log [M^{2+}]$) varies between 6 and 7. The next strongest complex-forming compounds are **2** and **5**, where the p_M values are in the range of 4.5–5.5. The weakest complexes are formed with ligands **4**, **1**, and **3**, where the concentration of the uncomplexed metal ion varies between 20% and 40% in the pH range < 5. These trends are corroborated by the magnitude of the conditional and basicity-adjusted stability constants. Here, the trisphosphonate **1** forms relatively weak complexes, as expected. The stability of the complexes decreases with the number of phosphonate moieties, however, due to the stability decreasing effects of the larger spatial requirement and the higher charge of the phosphonate group. This is also reflected in the shift of the pH at the formation maximum of the corresponding NiL species to higher values upon PO_3^{2-}/CO_2^- substitution (see Figure 1).

The levels of free metal ion concentration in the region of pH 2–5 suggest the following sequences of increasing degree of metal ion complexation:

$$\mathrm{Ni^{2+}complexes:} \quad \mathbf{7} \sim \mathbf{4} < \mathbf{1} \sim \mathbf{3} \ll \mathbf{2} < \mathbf{5} \ll \mathbf{6} \tag{2}$$

$$\mathrm{Cu^{2+}complexes:} \quad \mathbf{7} \sim \mathbf{4} \sim \mathbf{1} < \mathbf{3} \ll \mathbf{2} < \mathbf{5} \ll \mathbf{6} \tag{3}$$

$$\mathrm{Zn^{2+}complexes:} \quad \mathbf{7} \sim \mathbf{4} < \mathbf{3} \sim \mathbf{1} \ll \mathbf{2} \ll \mathbf{5} \ll \mathbf{6} \tag{4}$$

$$\mathrm{Cd^{2+}complexes:} \quad \mathbf{7} \sim \mathbf{4} < \mathbf{1} < \mathbf{3} \ll \mathbf{2} \ll \mathbf{5} << \mathbf{6} \tag{5}$$

The above order is in good accordance with the metal-binding ability of the *protonated ligands* estimated by the relative magnitudes of conditional stepwise stability constants (see Table I). In the series of compounds studied here, the structural motif for outstanding metal complex stability was proven to be the carboxylate and phosphonate groups in geminal position. The $\log K_i > 4$ values for these ligands indicate the formation of at least one chelate ring.[22] Compound **4** contains already two pairs of geminal COO^-/PO_3^{2-} moieties, but here the sterical demand of the charged donor groups leads to diminished metal-binding affinity.

Ca^{2+} forms complexes of significant lower stability as expected. Reliable formation constants could be evaluated only from titrations at $T_{Ca}/T_L \geq 5$. The pCa values range from 1.5 to 1.7 and the following sequence of metal-binding affinity can be established:

$$\mathrm{Ca^{2+}complexes:} \quad \mathbf{7} < \mathbf{4} < \mathbf{1} < \mathbf{5} = \mathbf{2} \tag{6}$$

CONCLUSION

The investigated oligophosphonate/carboxylate ligands were shown to form only mononuclear complexes even at higher metal : ligand ratios in aqueous solution (the analysis of precipitates observed at higher pHs was not attempted). The protonated complexes MLH_i prevail in the same pH range as LH_i. The high conditional stability constants found for several protonated complexes point to chelated structures. The increase in log β stability constants with increasing number of phosphonate groups shows their higher ability to form complexes in a biased fashion: log β values should be purified from effects of different basicity. Since the greater charge density of the phosphonate(s) is overcompensated by their higher sterical bulk and greater electrostatic repulsion, the basicity-adjusted or the conditional stability constants indicate that the gain in affinity to metals diminishes or vanishes, as reported previously in a similar comparison of aminocarboxylates/phosphonates.[22] The most stable complexes were observed for ligands holding PO_3^{2-} and CO_2^- groups in geminal position.

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