167. The Stability of Divalent 3d- and Trivalent 4f-Metal Complexes with Diazapolyoxa Macrocyclic Ligands

by **Giorgio Anderegg**

Laboratorium für Anorganische Chemie, ETHZ, Universitätsstrasse 6, CH-8092 Zürich

(18. V. 81)

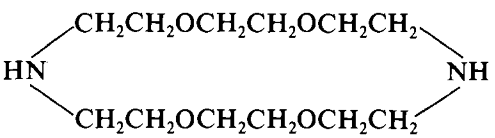
Summary

The complex formation of some divalent 3d- and trivalent 4f-metal ions with four macrocyclic ligands has been investigated in methanol/water 95:5 by use of alkalimetric titrations. In contrast to alkali and alkaline-earth cations the obtained stability constants are quite low. The reasons of this effect are discussed.

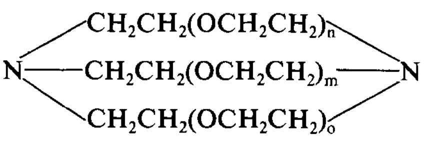
In recent years macrocyclic ligands have aroused considerable interest and the diazapolyoxa macrocycles probably represent one of the more interesting series of synthetic substances of this type. They have been obtained by *Lehn et al.* [1] [2] and form very stable complexes with different alkali and alkaline-earth cations. Their stability constants are of comparable magnitude with those obtained by the most powerful known non-cyclic multidentate ligands [3–6].

Some years ago we undertook an investigation on the stability of the complexes between 3d- and 4f-metal ions and these ligands. Since these complexes are not quite stable in aqueous solution we used methanol/water 95:5 *(v/v)* as solvent [4]. Indeed, in this medium the stability constants of the complexes are some orders of magnitude higher than in water [2]. In the meantime three publications have appeared on the formation of the 3d- and 4f-metal complexes but in aqueous solution [7] [8] and of the Cu^{2+} and Zn^{2+} complexes in absolute methanol [9].

This paper reports the study of one monocyclic and three bicyclic diamine ligands, *i. e.* [2.2], [2.1.1], [2.2.1] and [2.2.2] respectively, using *Lehns* symbolism [2].



[2.2] **1**



[2.1.1] **2**

[n.m.o] : [2.2.1] **3**

[2.2.2] **4**

0018-019X/81/6/1790-06$01.00/0

Experimental. – Ligands **2**, **3** and **4** *(Merck)* were used without further purification. The ligand *1* was synthesized following the method of *Lehn et al.* [1]. The inert salt, tetramethylammonium chloride *(Fluka, puriss.)* was recrystallized from methanol/acetone and dried under vacuum for 48 h at room temp. before use. Tetramethylammonium hydroxide was prepared as already described [10]. Methanol *(Fluka, puriss.)* was flushed with N_2 before use. The solvent used was a mixture of methanol/water 95:5 obtained by taking 50 ml of water and adding methanol up to a total volume of 1000 ml [4]. Metal chlorides and cadmium nitrate *(Merck* or *Fluka, p. a.)* were used.

For the determination of the protonation and the stability constants (K_1, K_2 and K_{ML}), solutions of the protonated ligands (0.002 M)

$$H^+ + L \rightleftharpoons HL^+ \qquad K_1 = [HL]/([H]\,[L]) \tag{1}$$

$$H^+ + HL^+ \rightleftharpoons H_2L^{2+} \qquad K_2 = [H_2L]/[([HL]\,[H]) \tag{2}$$

$$M^{z+} + L \rightleftharpoons ML^{z+} \qquad K_{ML} = [ML]/([M]\,[L]) \tag{3}$$

alone and in presence of the investigated metal ion (M^{z+}, 0.002M) were titrated with strong base. The pH of the solution, defined in the concentration scale (pH = –log [H^+]), was measured with the following cell:

	Test solution	Reference electrode
Glass electrode	$I = 0.1$ ($(CH_3)_4NCl$)	0.1 M $(CH_3)_4NCl$, Hg_2Cl_2(s), Hg(l)
	MeOH/H_2O 95:5	MeOH/H_2O 95:5

Because of the use of solutions of very similar compositions for the reference electrode and for the test solution, diffusion phenomena can be neglected. The calibration of the cell was done with HCl-solutions at $I = 0.1$ ($(CH_2)_4NCl$) by titration with 0.1 M tetramethylammonium hydroxide. From the results of this titration in basic solution it was possible to obtain the ionic product [H] [OH] = $10^{-15.00}$ valid for the conditions used: $I = 0.1$ and 25°C. Care should be taken in order to maintain unaltered the ratio MeOH/H_2O in the measured solution: a change of the ratio MeOH/H_2O causes a strong pH-change of the solution, *e. g.* for a 95:5 solution addition of 1 vol% H_2O causes a change in the potential of 5 mV, which corresponds to 0.09 pH units. As the complexes of the 4f-cations are formed slowly a batch procedure was used. A series of solutions with the same quantity of protonated ligand and cation but different quantities of strong base were measured till a constant pH-value was obtained. To obtain equilibrium in the case discussed (La^{3+}) one month was necessary. Using the same tedious procedure, the formation of hydrolytic solid products was observed for the bicyclic ligands and the following cations: Pr^{3+}, Yb^{3+}, Gd^{3+}, Sm^{3+} and Th^{4+}. With [2.2.1] and [2.1.1] hydrolysis occurs also for La^{3+}.

The protonation and stability constants were calculated using, in all cases, appropriate programs on a CDC (6500–6400) computer by the method described elsewhere [10–12].

The error in the logarithm of the constants is estimated to be approximately 0.1 for the reasons given above.

Results. – The stability constants obtained, together with some literature values are reported in the *Table*. The values for other cations, in some cases as a result of new measurements, have been included for comparison. In general it was difficult to obtain accurate values for the constants of the equilibria involved. This is due to the very limited pH range in which complex formation occurs without competing processes, such as the formation of weak protonated species and hydrolysis involving hydroxo species. The situation is not improved even when the more powerful ligand [2.2.2] is used. Some of the 3d- and 4f-cations showed a similar behaviour also in aqueous solution [7] [8]. For complexes of the 4f trivalent cations with the bicyclic

Table. *Constants of the equilibria* 1, 2 *and* 3 *at* I = *0.1* (($CH_3)_4NCl$) *and 25°C and corresponding literature values for MeOH/H_2O/95:5 and for aqueous solutions*

Ligand	**1**			**2**				**3**				**4**				
	95/5	H_2O		95/5	H_2O			95/5	H_2O			95/5	H_2O			
		a)	b)		c)	d)	b)		c)	b)	d)		c)	a)	b)	d)
H^+ log K_1	9.28	9.20	9.08	11.59	11.00	10.64	11.32	10.97	10.42	11.02	10.53	10.05	9.85	9.71	10.00	9.60
log K_2	7.97	8.02	7.94	7.12	6.56	7.35	8.14	7.31	6.60	7.74	7.50	7.15	6.64	7.31	7.53	7.28
Li^+	1.07			7.93	7.58	5.5		4.46	4.18		2.5	~ 1				
Na^+				6.53	6.08	3.2		9.35	8.84		5.4	7.4	7.21	4.1		
Ca^{2+}	4.04	1.8		4.5	4.34	2.5						7.5	7.60	4.0		
Sr^{2+}	5.60	2.8														
Co^{2+}	≤3.5		≤2.5					5.92		5.40		≤ 4			≤ 2.5	
Ni^{2+}	≤2.5		≤2.5				≤ 4.5			≤ 4.28		≤ 4.5			≤ 3.5	
Cu^{2+}	8.77		6.18	≤ 7.3			7.78	8.71		7.56		≤ 6			6.81	
Zn^{2+}	≤2.5						≤ 5.3			5.41		≤ 4			≤ 2.5	
Cd^{2+}	7.18e)	5.25e)														
Ce^{2+}												8.4				
La^{3+}	6.18											9.4		f)		

a) $I = 0.1((CH_3)_4NCl)$ and 25°C[6].
b) $I = 0.1((C_2H_5)_4NClO_4)$ and 25°C[7].
c) $I = 0.01$ and 25°C[4]
d) $I = 0.05$ and 25°C[4]
e) $(CH_3)_4NNO_3$ instead of $(CH_3)_4NCl$ was used:
f) In 0.25 M $(CH_3)_4NCl$ and 25°C, log $K_{ML} = 6.45$[8].
Other values are given for some 4f-cations as well as with **2** and **3**.

ligands, all values which can be obtained with reasonable accuracy are given. The reluctance of the cryptands to form such complexes has been reported in the literature [13–17]. Indeed their salts are, in general, only obtained working with anhydrous solvents. Because of the strong basicity of the ligands, these salts in presence of water form the protonated ligand H_2L^{2+} with precipitation of metal hydroxide. For this reason, the stability constants of the 4f cations of *Burn & Baes (Table)*[8] for aqueous solutions are unexpected. However, calculations show that at the given pH-values of the equilibrated solutions (<6.4), less than 10% of the metal ion lies as complex ML^{3+}. Probably, under these conditions, metal hydrolysis can be avoided but the very limited range of measurement can be dangerous for the interpretation of the results. In comparison, in our measurements of Ce^{3+} and La^{3+} with **4** up to 75% of the metal was present as 1:1 complex.

A further complication is the low rate of complex formation observed by alkalimetric titration: after addition of a small quantity of tetramethylammonium hydroxide, which corresponds to the neutralisation of 10% of the diamine, several days are necessary to obtain a constant pH-value. In other cases some precipitation of metal hydroxide occurs. On the other hand, the complexes once formed, also show great inertness: no dissociation of LaL^{3+} is observed if an acidified solution at pH = 3 was left one day at 60° C. Under these conditions, on the basis of the measured stability constant, the complete disruption of the complex is expected.

Discussion. – The macrobicyclic ligands **2**, **3** and **4** form with alkali and alkaline-earth cations (1:1)-complexes of unexpected high stability even in aqueous solution [1–6]. As found here and elsewhere [7], this seems not to be the case for the divalent 3d- and the trivalent 4f-cations. Indeed, with these metal ion complexes organic solvents are more appropriate and sometimes necessary for the equilibrium measurements and for the synthesis. This arises because, in these media, the stability constants are more than two orders of magnitude higher with respect to those in water, and side reactions are generally less frequent.

The destabilizing effect of the ligands investigated on complex formation with copper(II) *in water* appears in unequivocal manner from the graphical plot of the values of the stability constants of the (1:1)-complexes with these and other diamines [7] [18] *vs.* the product of the protonation constants $K_1 \cdot K_2$ of the ligands *(Figure)*. A straigth line can be drawn between the points obtained, giving approximately the dependence of the basicity of the ligands on the stability of their (1:1)-Cu(II)-complexes. For a diazapolyoxa ligand having similar protonation constants such as ethylenediamine (en [18]) a stability constant of $10^{6.8}$ is expected, *i. e.* a value which is four orders of magnitude lower than that for copper ethylenediamine. The plot of *Figure* shows that, for copper(II), the bicyclic ligands are not particulary preferred with respect to the monocyclic ligand [2.2]. Only for the copper(II) complex with [2.1] the destabilization is lower. For the other divalent 3d-cations, because for the lower values of their stability constants with ethylenediamine, the same destabilization would lead to systems in which hydrolytic processes would prevail over complex formation.

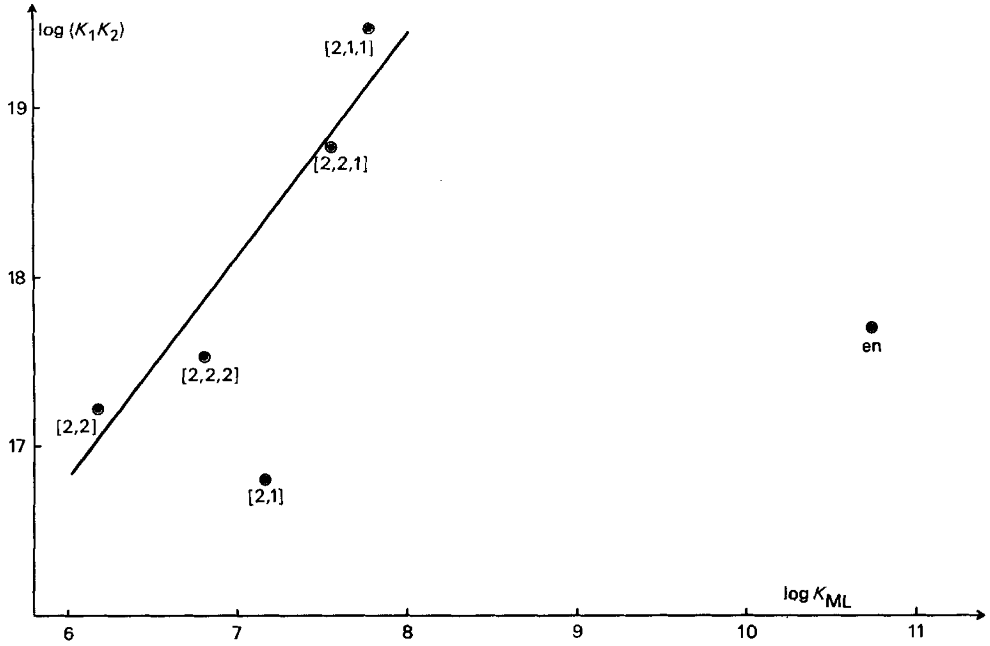
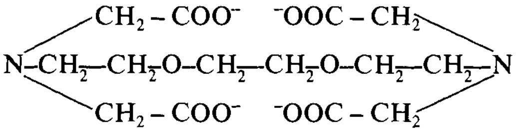


Figure. *The stability constants of the (1:1)-Cu(II)-complexes of diamines* vs. *their overall protonation constants*

The formation of weak complexes in the case discussed can be also explained calculating for the ligands **1** to **4** the apparent conditional $K' = [ML/([M] [L'])$ at $pH = 7$ [19]. The calculations give K' values between $10^{2.8}$ and $10^{3.3}$ which corresponds for an excess of ligand of $5 \cdot 10^{-3}$M to a concentration of the free copper from 0.3 to $0.1 \cdot [Cu]$. If also the hydrolysis of the copper ions is considered, *i. e.* the formation of $Cu_2(OH)_2^{2+}$ ($[Cu_2(OH)_2][H]^2/[Cu]^2 = 10^{-10.6}$[20]), it appears that the amount of this species could reach values between 60 and 6% of $[Cu]_t$. Summing-up, the complexation of these ligands takes place, for millimolar solutions, in a pH range in which also hydrolysis of the Cu(II)-ion occurs.

Some difficulties in the coordination of Ca^{2+} (r = 1.06 Å) by [2.2.2] are also apparent from X ray data of the bromide complex salt. Indeed, already for this cation, the ligand [2.2.2] 'has not enough flexibility to adjust to the small size of the alkaline-earth cations' [21]. It is therefore expected that these difficulties will increase in the case of the much smaller divalent cations with r = 0.8 to 0.7 Å. Furthermore, these cations are more strongly stabilized, with respect to the alkaline-earth cations, by hydration and show, as aqua ions, weak tendency to bind the macrocyclic ligands **1** to **4**. Probably a similar situation also accounts for the slight tendency of trivalent 4f-cations to coordinate the macrocyclic ligands **2** to **4**. In conclusion, the selective and efficient coordination of the ligands **2** to **4** with alkali and alkaline-earth cations is mainly due to the cationic hydration terms (ΔH_h and ΔS_h) which make the above ligands competitive with respect to water. However, already for the small Mg^{2+}-ion, the stability constant is too low. The enthalpy and the entropy of hydration decreases strongly for highly charged (4f trivalent) and for small (3d divalent) cations in such a way that the uncharged ligands **2** to **4** are not able to substitute the coordinated water molecules even if this reaction were to cause an appreciable entropy increase [6]. Within the lanthanide cations, only the first two members of the series, La(III) and Ce(III), form quite stable complexes with the ligand **4** but already for Pr(III) hydrolytic products are observed. It is therefore conceivable, that already for Ce(III) the solvent can partly compete as ligand, making the effective coordination with the bicyclic ligand more difficult, in such a way that the difference in the constants K_{ML} of La(III) and Ce(III) can be justified.

In contrast to the ether O-atom, the negatively charged carboxyl O-atom is more competitive for the investigated cations. Leaving in [2.2.2] only one dioxa bridge unaltered and substituting four CH_2-O-CH_2 groups with four carboxylate groups the flexible EGTA is obtained. It is also an octadentate ligand and forms quite stable complexes with the 3d divalent and 4f trivalent cations [21].



EGTA

The author wishes to express his gratitude to Mr. *B. Rüttimann* for his help in the experimental work.

REFERENCES

[1] *B. Dietrich, J. M. Lehn & J. P. Sauvage,* Tetrahedron Lett. *1969,* 2885.

[2] *B. Dietrich, J. M. Lehn & J. P. Sauvage,* Tetrahedron *29,* 1647 (1973).

[3] *J. M. Lehn,* Structure and Bonding *16,* 1 (1973).

[4] *J. M. Lehn & J. P. Sauvage,* J. Am. Chem. Soc. *97,* 6700 (1975).

[5] *E. Kauffmann, J. M. Lehn & J. P. Sauvage,* Helv. Chim. Acta *59,* 1099 (1976).

[6] *G. Anderegg,* Helv. Chim. Acta *58,* 1218 (1975).

[7] *F. Arnaud-Neuf, B. Spiess & M. J. Schwing-Weill,* Helv. Chim. Acta *60,* 2633 (1977).

[8] *J. M. Burns & C. F. Baes jr.,* Inorg. Chem. *20,* 616 (1981).

[9] *B. Spiess, F. Arnaud-Neuf & M. J. Schwing-Weill,* Helv. Chim. Acta *92,* 1531 (1979).

[10] *G. Anderegg, A. Ekström, L. F. Lindoy & R. J. Smith,* J. Am. Chem. Soc. *102,* 2670 (1980).

[11] *G. Anderegg,* Helv. Chim. Acta *44,* 1673 (1961).

[12] *H. Stünzi & G. Anderegg,* Helv. Chim. Acta *59,* 1621 (1976).

[13] *O. A. Gansow, A. R. Kausar, K. B. Triplett, M. J. Weaver & E. L. Yee,* J. Am. Chem. Soc. *99,* 7087 (1977).

[14] *O. A. Gansow, D. Pruett, R. Kausar & K. B. Triplett,* Proc. XIX ICCC Prague 1978, 47a.

[15] *O. A. Gansow, D. J. Pruett & K. B. Triplett,* J. Am. Chem. Soc. *101,* 4408 (1979).

[16] *M. Ciampolini, P. Dapporto & N. Nardi,* J. Chem. Soc., Dalton *1979,* 974.

[17] *F. A. Hart, M. B. Hursthouse, K. M. A. Malik & S. Moorhause,* J. Chem. Soc., Chem. Commun. *1978,* 549.

[18] *J. Bjerrum & E. J. Neilson,* Acta Chem. Scand. *2,* 307 (1948).

[19] *A. Ringbom,* 'Complexation in Analytical Chemistry', Interscience Publ., 1963.

[20] *R. E. Mesmer & C. F. Baes, jr.* 'The Hydrolysis of Cation', Oka Ridge National Laboratory 1974.

[21] *B. Metz, D. Moras & R. Weiss,* Acta Crystall. Ser. B *29,* 1377 (1973).

[22] *G. Anderegg,* Helv. Chim. Acta *47,* 1001 (1964).