ARTICLE

OBC
www.rsc.org/obc

Novel chiral N_4S_2- and N_6S_3-donor macrocyclic ligands: synthesis, protonation constants, metal-ion binding and asymmetric catalysis in the Henry reaction

Jian Gao and A. E. Martell *
Department of Chemistry, Texas A & M University, College Station, TX 77842-3012, USA

Received 19th May 2003, Accepted 18th June 2003
First published as an Advance Article on the web 7th July 2003

New hydrophobic chiral macrocyclic ligands **L1**–**L3** with chiral diamino and thiophene moieties have been synthesized by the Schiff base condensation approach. Protonation constants of **L1** and **L2** were determined by potentiometry titration. Metal-ion binding experiments exhibited that **L1** and **L3** are pronounced in selective recognition, Ag^+, Cu^{2+} and Ca^{2+} ions among the surveyed metal ions (Cu^{2+}, Co^{2+}, Ni^{2+}, Zn^{2+}, Cd^{2+}, Pb^{2+}, Ag^+, Li^+, Na^+, K^+, and Ca^{2+}). **L1** was found to spectroscopically detect the presence of Cu^{2+} and Ca^{2+} to function as a multiple readout sensor. The detection limit for Ca^{2+} ions was found to be 9.8×10^{-5} M in CH_2Cl_2–MeOH solution. The trimeric chiral ligand **L3** has been shown to be an efficient auxiliary in a Zn(II)-mediated enantioselective Henry reaction.

1. Introduction

Currently, the development of novel ligands with specific functional groups for the efficient detection of metal ions is one of the most important areas in host–guest chemistry.[1,2] Polyaza macrocyclic ligands were found to be effective in binding transition metal ions[3,4] and continue to be the subject of intense research.[5,6] To establish new properties and the potential applications of these macrocyclic ligands, one avenue of special interest would be to create new hydrophobic macrocycles, for the purpose of selectively binding metal ions,[7,8] for the stabilization of reactive intermediates[9] and for the design of hydrophobic complex catalysts.[10] The combined structural features of **L1**–**L3** comprising chiral diamines and hydrophobic thiophene moieties would be expected to display a range of specific metal-binding and asymmetric catalytic properties in addition to those discovered in our former investigations.

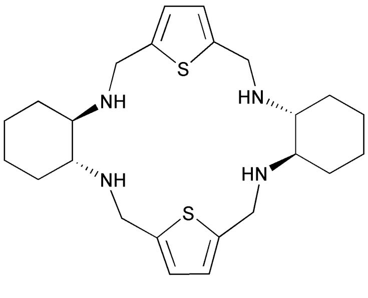
2. Results and discussion

2.1 Synthesis of the chiral macrocyclic ligands

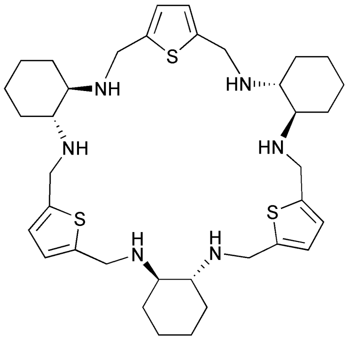
The preparation of the achiral ligand (**L1**) was first attempted by the Schiff base condensation of 2,5-thiophenedicarboxaldehyde with 1*R*,2*R*-diaminocyclohexane at room temperature and this led to a mixture containing the dimeric and trimeric Schiff base ligands, and the corresponding polyamino ligands obtained after reduction with $NaBH_4$ were identified by ESI-MS spectra (Figs 1A and 1B respectively). Switching to a Pb(II) ion templated approach, the resulting product was a mixture of the dimeric and trimeric macrocycles, with **L2** being the main product (~76%, Fig. 1C), indicating that a metal ion template is less useful in the synthesis of polyazathio crowns possessing more than approximately 24 atoms in their macrocyclic chain. Pure dimeric and trimeric ligands were successfully separated by chromatographic purification (CH_2Cl_2–MeOH = 2 : 8). Reaction of **L1** with four equiv. of bromomethylbenzene in the presence of K_2CO_3 afforded the substituted derivative **L3** in moderate yield (49%, Scheme 1).

2.2 Protonation constants of the ligands

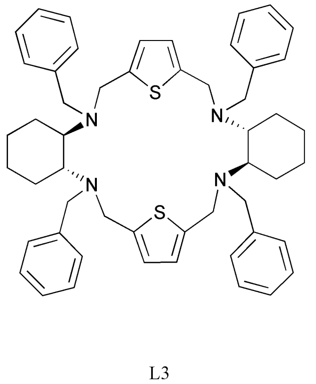
The potentiometric equilibrium curves for **L1**·4HBr and **L2**·6HBr are illustrated in Fig. 2. The pH profiles of **L1** reveal an inflection at *a* = 2 (*a* = moles of base added per mole of the ligand), indicating the first two acidic protons are ready to be neutralized. For **L2**, from *a* = 0 to 3 and *a* = 3 to 6, there are two buffer regions. The first buffered region corresponds to the completion of the neutralization of the three most acidic protons. The consequent buffered region at high *a* value corresponds to the dissociation of other protons from the macrocycle. The calculated protonation constants which are derived from the titration data are profiled in Table 1. For **L2**, the constants display the order p*K*1~ p*K*2 ~ p*K*3 > p*K*4 ~ p*K*5 ~ p*K*6. The overall log protonation constants, $\Sigma \log K_i^H$, for the chiral hexaza macrocycles turn out to be 38.50, which is much higher than the component 1*R*, 2*R*-diaminocyclohexane value of 16.32.[11] The $\Sigma \log K_i^H$ value for **L1** is 33.26, which is much higher than that of cyclen (22.39),[11] a tetraaza macrocyclic analogue. This indicated that the overall basicity of certain



L1



L2



L3

DOI: 10.1039/b305582e

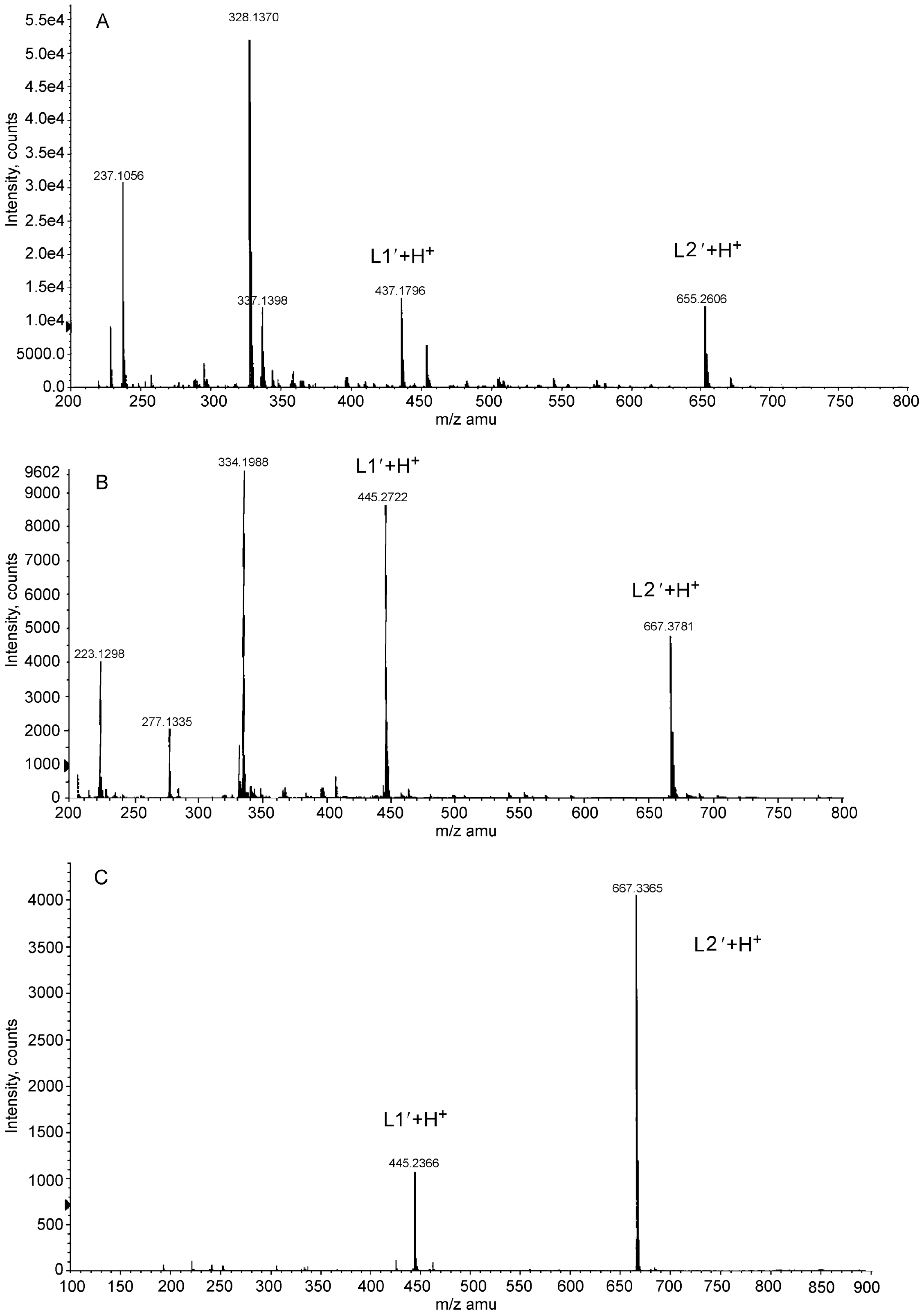


Fig. 1 ESI-MS spectra of **L1′**, **L2′** (A); **L1**, **L2** (B) and **L1**, **L2** obtained from template synthesis (C).

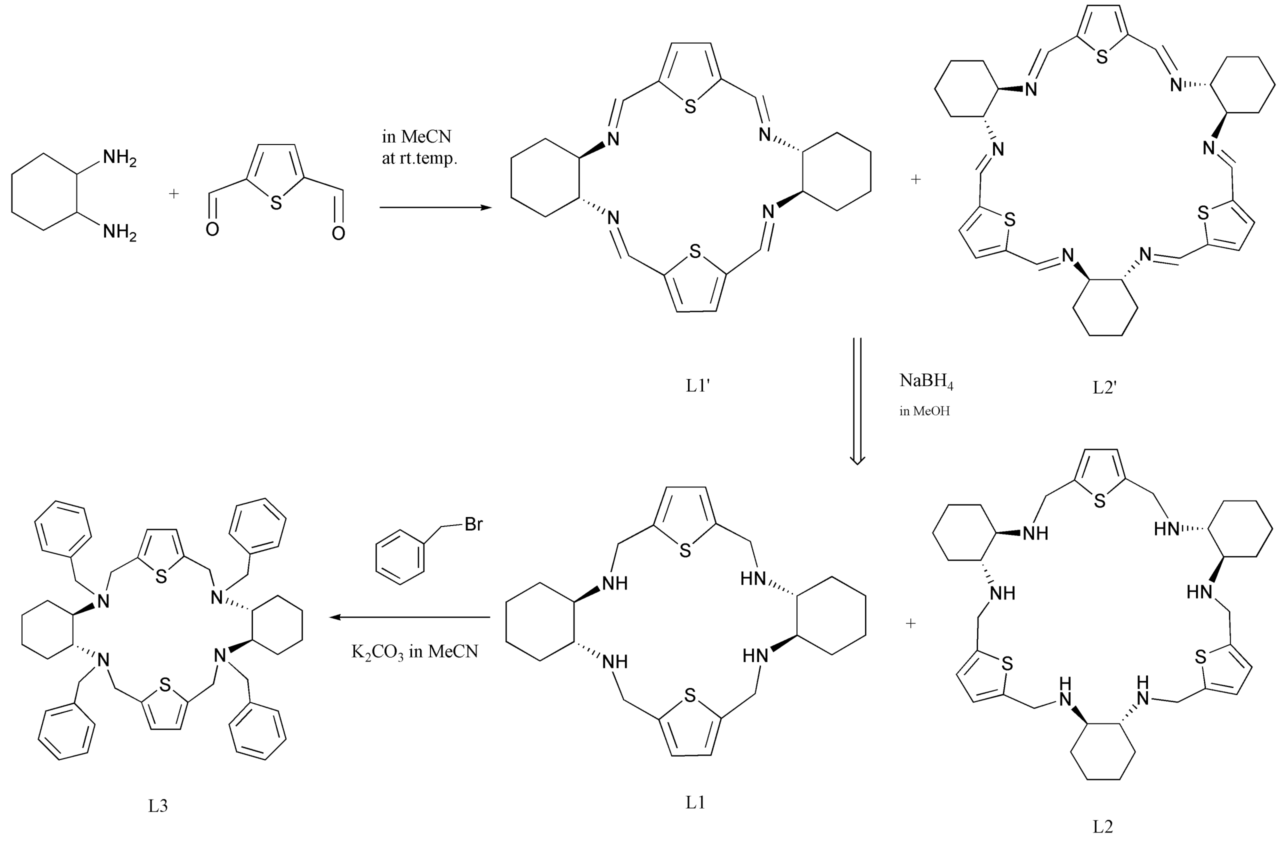
ligands is strongly influenced by the chirality and conformational properties. Incorporation of the thiophene moieties may also contribute to this high overall protonation constant. The species distribution diagram for **L1** (Fig. 3) clearly shows that the forth proton remains unneutralized even in strong basic solution. The species distribution curves for **L2** (Fig. 4) show that **L2**·6HBr is easily deprotonated with the triple deprotonated form prevailing at pH = 6.8 and the free ligand dominating at above pH 10.0.

2.3 Macrocyclic ligands binding with metal ions

Preliminary binding abilities of **L1** were investigated by the well-established potentiometric method.[12] The potentiometric equilibrium curves for **L1** and a 1 : 1 ratio of **L1** to metal ions (Cu^{2+}, Ni^{2+}, Co^{2+}, Zn^{2+}, Ag^{+} and Ca^{2+}) are shown in Fig. 5. A strong inflection at $a = 2$ was observed for Ca^{2+}, Co^{2+}, Ni^{2+} and Zn^{2+}. However, for Ag^{+} and Cu^{2+}, the inflection was observed at a = 3–4, indicating that their affinity for **L1** is stronger than

Table 1 Logarithms of the protonation constants of **L1** and **L2** ($\mu = 0.100$ M KCl, $T = 25.0$ °C, under argon)

Symbol	Equilibrium quotient	LogK^{H}_{i} of **L1**	Equilibrium quotient	LogK^{H}_{i} of **L2**
K^{H}_{6}			$[H_6L]/[H_5L][H]$	3.62
K^{H}_{5}			$[H_5L]/[H_4L][H]$	3.76
K^{H}_{4}	$[H_4L]/[H_3L][H]$	3.76	$[H_4L]/[H_3L][H]$	5.15
K^{H}_{3}	$[H_3L]/[H_2L][H]$	5.87	$[H_3L]/[H_2L][H]$	8.41
K^{H}_{2}	$[H_2L]/[HL][H]$	8.14	$[H_2L]/[HL][H]$	8.43
K^{H}_{1}	$[HL]/[L][H]$	15.49	$[HL]/[L][H]$	9.11
ΣK^{H}_{i}	$[H_4L]/[L][H]^4$	33.26	$[H_6L]/[L][H]^6$	38.50



Scheme 1

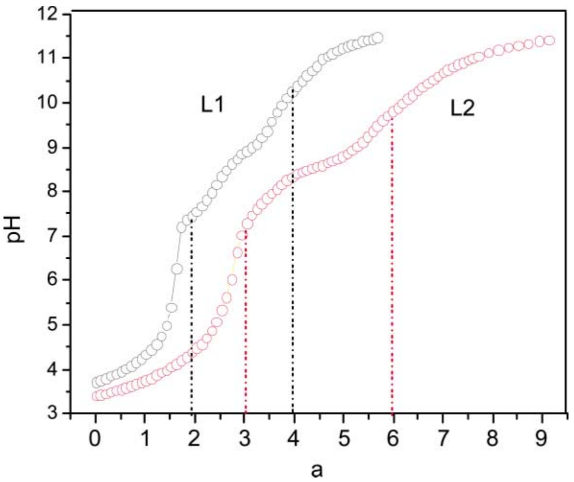


Fig. 2 Potentiometric equilibrium curves for **L1** and **L2** ($\mu = 0.100$ M KCl, $T = 25$ °C, a = moles of KOH added per mole of L, $T_L = 0.001$ M).

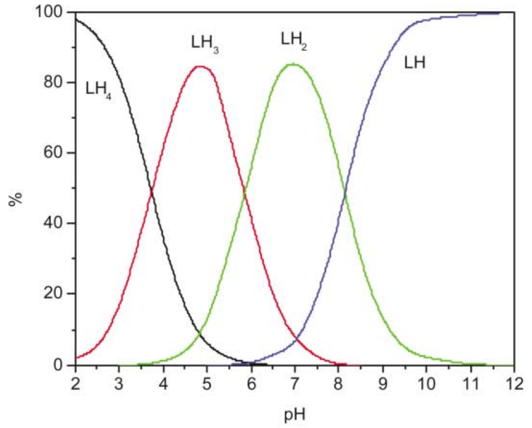


Fig. 3 Species distribution diagram showing the species formed as a function of pH when **L1** = 0.001 M ($\mu = 0.100$ M KCl, $T = 25$ °C).

the former metal ions. Spectrophotometric titration is a useful tool for the determination of the stability constants of metal complexes, especially for the solution structure of 1 : 1 complexes between ligand and metal ions.[13] For brevity, only the spectroscopic data for the 1 : 1 **L1**-Cu(II) system is reported in detail. Fig. 6 presents the changes in the UV-vis spectrum of **L1** upon the titration with Cu(II) solution. The binding constants that were determined from the experimental data are compiled in Table 2, showing the relative binding order for the surveyed metal ions as: Ni(II) < Co(II) < Zn(II) < Ca(II) < Cu(II) < Ag(I).

Metal ion solvent extraction experiments (water–chloroform) have been carried out using **L1** and **L3** as extractants. The procedure employed involved competitive metal extractions from an aqueous phase at pH = 9.0 containing equal concentrations

Table 2 Logarithms of the stability constants (in form of [**L1**M^{n+}]) of metal complexes (T = 25 °C, μ = 0.100 M KCl)

[L1M^{n+}]/[L1][M^{n+}]	Ni^{2+}	Co^{2+}	Zn^{2+}	Ca^{2+}	Cu^{2+}	Ag^{+}
Stability constants (Log$K$)	18.24	18.50	19.41	21.65	22.55	23.57

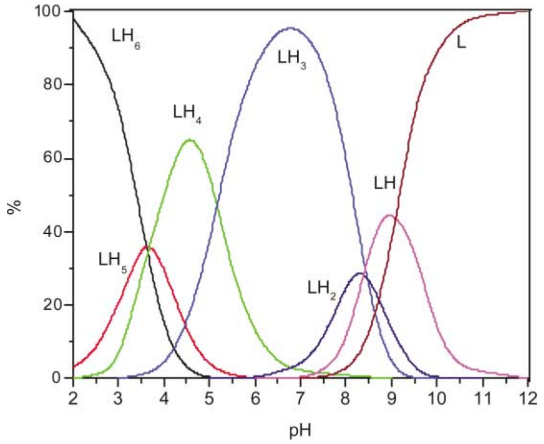


Fig. 4 Species distribution diagram showing the species formed as a function of pH when **L2** = 0.001 M (μ = 0.100 M KCl, T = 25 °C).

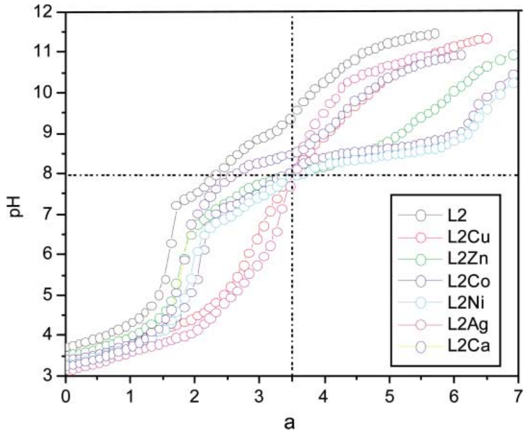


Fig. 5 Potentionmetric equilibrium curves for **L1** and 1 : 1 **L1**/M^{n+} systems (M^{n+} = Cu^{2+}, Zn^{2+}, Co^{2+}, Ni^{2+}, Ag^{+}, and Ca^{2+}) (μ = 0.100 M KCl, T = 25 °C, a = moles of KOH added per mole of L, T_L = T_{Mn+} = 0.05 M).

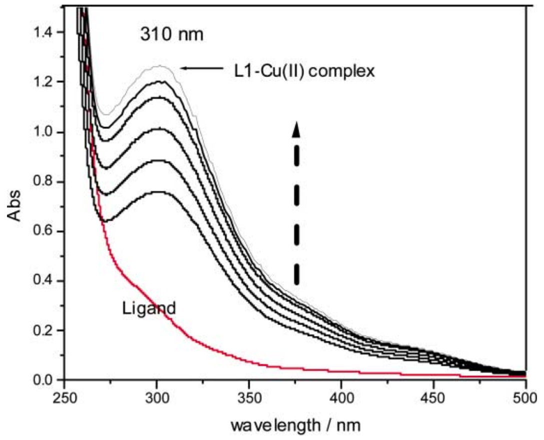


Fig. 6 Absorption variation of spectrophotometric titration of **L1** with Cu^{2+} in methanol solution.

of metal ions. The results are presented in Fig. 7. For ready comparison of the efficiencies of different macrocycles, the degree of extraction is presented as the percentage of metal ions extracted in each experiment. In the two experiments, the respective systems showed sole selectivity for Ag$^+$ in conformation of the expected affinity of this ion for a N_4S_2-donor set.

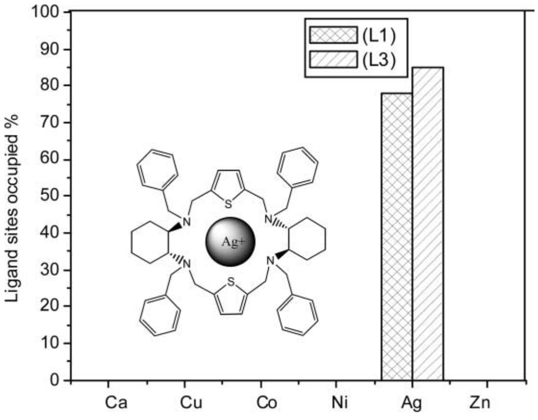


Fig. 7 Mixed metal ion extraction data (source phase contained six metal nitrates, each at 10^{-2} mol dm^{-3}).

While the substitution on the macrocyclic ring clearly influences the efficiency of the extraction, silver ion selectivity is maintained in each case. In the present study, **L3** was found to be a more efficient extractor of silver(I), which could be attributed to the great lipophilicity of this ligand. The hydrophobic binding pocket that formed after tetrakisbenzylation will readily host Ag$^+$ ions.

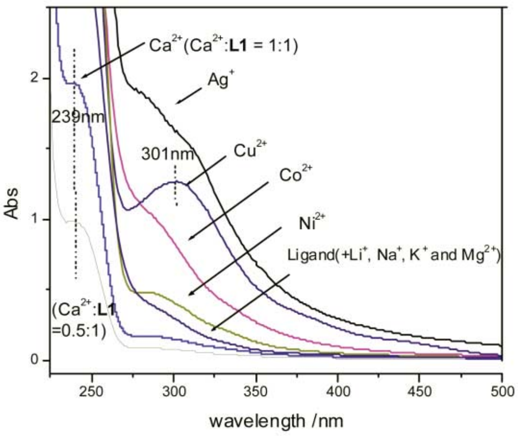


Fig. 8 UV-vis absorption spectra of **L1** (5.0 × 10^{-4} M) in CH_2Cl_2–MeOH (10 : 1) in the presence of metal ions ([M^{n+}] = 0.50 mM at 25 °C).

To evaluate the ability of **L1** to discriminate and signal specific metal ions in the presence of other metal cations, competitive binding experiments were performed. UV/vis spectra of **L1** (5.0 × 10^{-4} M) with stoichiometric mixtures of ions taken from the series of Ca^{2+}, Co^{2+}, Ni^{2+}, Cu^{2+} and Ag$^+$ were recorded in 10 : 1 CH_2Cl_2–MeOH and were compared to those of the corresponding 1 : 1 **L1**–M^{n+} solution. A mixture of 1 : 1 : 1 **L1**–Cu^{2+}–Co^{2+} and **L1**–Cu^{2+}–Ni^{2+} gave spectra identical to those of the 1 : 1 **L1**–Cu^{2+} spectrum (Fig. 7), showing that Co^{2+} and Ni^{2+} are displaced from **L1** by Cu^{2+}. A 1 : 1 : 1 : 1 : 1 : 1 mixture of **L1**–Ag$^+$–Cu^{2+}–Ni^{2+}–Zn^{2+}–Ca^{2+} was identical to that of a 1 : 1 **L1**–Ag$^+$ solution at the same [**L1**]. This strong selectivity for Ag$^+$ in the presence of Ca^{2+}, Cu^{2+}, Ni^{2+} and Zn^{2+} must reflect the well-documented affinity for soft donors[14,15] such as a thioether sulfur.

Preliminary ionophoric properties of **L1** for alkali and alkaline earth metal ions (Li$^+$, Na$^+$, K$^+$ and Ca^{2+}) of physiological interest[16] are also investigated spectroscopically. Upon interaction with Ca^{2+} ions, the colorless solution of **L1** was

Table 3 Catalytic nitroaldol reaction with CH_3NO_2 (catalyst concentration is 5 mol%)

Entry	Catalysts	Solvent	Yield (%)[a]	ee (%)[b]
1	1*R*,2*R*-DACH+Zn(II)	THF	19	21
2	**L1**+Zn(II)	THF	45	39
3	**L2**+Zn(II)	THF	54	51
4	**L2**+2Zn(II)	THF	63	57
5	**L2**+3Zn(II)	THF	68	75

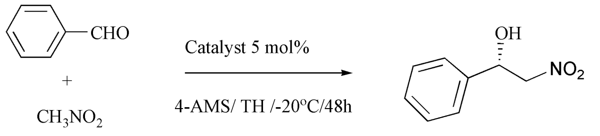
[a] All reaction s were run on a mmol scale, 20 mmol aldehyde in the presence of 100 mg 4 Å MS [b] Enantionmer excess was determined by Chiral HPLC (Chiralcel OD column).

transformed into a bright yellow one, which can be easily observed by the naked eye. As shown in Fig. 8, with the addition of equiv. of Ca^{2+} ions, a new absorption band at λmax = 239 nm was revealed. Other surveyed ions induced almost no change in the UV/vis spectrum.

The metal ion complexation studies thus revealed that **L1** functions as an UV/vis spectroscopic sensor for Cu^{2+} and Ca^{2+}, to function as a readout sensor. The detection limit for Ca^{2+} was estimated to be at a $[Ca^{2+}] = 1.02 \times 10^{-5}$ M at 235–245 nm. **L1** may potentially function as a chromogenic spectroscopic sensor. The study demonstrates the high selectivity of the present ligands for Ag^+ and confirms our proposal that the selective binding ability increases on introducing a hydrophobic moiety into a polyaza macrocyclic ligand.

2.4 Catalysis of an asymmetric Henry (nitroaldol) reaction

We have succeeded in using Zn(II) complexes of **L2** to catalyze the asymmetric nitroaldol reaction (Scheme 2). In the first step, metal ions (perchloride salts of Cu^{2+}, Co^{2+}, Zn^{2+}, Ni^{2+}, Ag^+, Ca^{2+}) were screened and the Zn^{2+} complex of **L1** was found to be the most promising in terms of product yield (Zn^{2+}, 44%; Cu^{2+}, 32%; Co^{2+}, 27%; Ni^{2+}, 11%; Ag^{2+}, 34%; Ca^{2+}, 25%). Consequently, Zn(II) complexes of **L1** and **L2** were tested in the enantioselective Henry reaction. The catalysts are prepared by treating **L1** or **L2** with one, two or three equiv. of diethylzinc (Scheme 3). The solution of the complex is then added directly to the reaction system containing nitromethane and the aldehyde substrate. Shibasaki and co-workers have reported a series of heterobimetallic catalysts that proved to be effective for asymmetric Henry reactions.[17] Trost recently reported a new type of asymmetric catalyst which involves a dinuclear zinc complex center with a chiral semi-azacrown ligand.[18] The trimeric chiral diamino macrocycle provides an unique opportunity to observe the cooperative mechanism, which is common to this type of reaction. The trinuclear complex catalyst system displayed substantial improvements in enantioselectivity relative to the mono, dinuclear analogues and the Zn(II) complex of chiral diamocyclohexane (Table 3), with kinetic behavior consistent with cooperative reactivity within the macrocyclic fromwork. The absolute configuration of the product was assigned by comparison to the literature.[18] A further investigation of the catalytic reaction is under the way.



Scheme 2

4. Experimental

Determination of the stability constants by spectrophotometric titrations

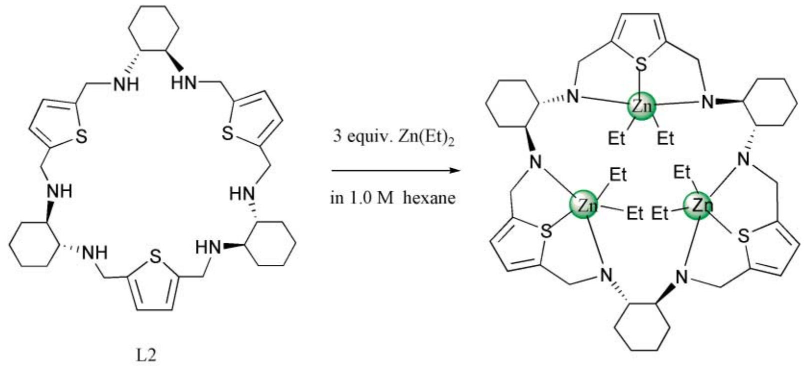
The stability constant *K*s which control the equilibrium between the free ligand **L1**, Cu^{2+} and the complex may be obtained from the variation of the absorbance intensity at 310 nm. It is easy to derive the following relationship involving the absorbance *A*o of the free ligand **L1** and *A* of the solution at a given wavelength: $A_o/(A_o - A) = [\varepsilon_L/(\varepsilon_L - \varepsilon_{ML})] \times (1/K\text{s}[M] + 1)$. ε_L and ε_{ML} are the molar absorption coefficients of **L1** and the complex.

Extraction experiments

The competitive metal extraction is from an aqueous phase into a chloroform phase. The aqueous phase was buffered at pH = 9.0 ± 0.1 with sodium phosphate buffer and contained equal concentrations (10^{-2} M) of Co^{2+}, Cu^{2+}, Ca^{2+}, Ag^+, Zn^{2+} and Ni^{2+} as their respective nitrate salt. The extractions were carried out in small sealed flasks in the absence of light to minimize the possibility of light-induced silver(I) decomposition. The flasks was agitated for one hour on a mechanical shaker after which time an aliquot of the organic phase was removed and evaporated to dryness. The residue was then taken up in nitric acid and, after appropriate dilution, was analyzed by atomic absorption spectrophotometry. Each experiment was performed in quadruplicate and the quoted values are the average from individual experiments.

Titration procedure

All of the metal stock solutions for potentiometric studies were reagent grade chloride salts prepared with doubly distilled water and standardized by EDTA. CO_2-free Dilute-it ampules of KOH were obtained from J. T. Baker Inc. KOH solutions (about 0.1 M) were prepared with doubly distilled water and were standardized. The extent of carbonate accumulation (<1.8 %) was checked periodically by titration with a standard HCl solution. A Corning 250 digital pH meter, fitted with Fisher full-range blue-glass and Fisher calomel reference electrodes were used for potentiometric titrations. A Metrohm



L2

L2[Zn(II)(Et$_2$)]$_3$

Scheme 3

(10 mL capacity) piston buret was used for precise delivery of standard KOH. The solution to be studied was contained in a 75 ml jacketed g1ass cell thermostated at 25.00 ± 0.05 °C by a circulating constant-temperature water bath.

Potentiometric determinations

All pH calibrations were performed with standardized HCl solutions to measure hydrogen ion concentrations directly (pH = $-\log [H^+]$). The ionic strength was adjusted to 0.100 M with KCl. Titrations of the ligand in the presence of metal ions in aqueous solution were conducted in the manner described by Martell and Motekaitis.[12] Cell solutions (in general, 50.00 ml) were purged with a purified argon stream. Standard base was introduced into the sample solutions with a Metrohm piston buret. Experimental runs were carried out by adding increments of standard base to a solution containing **L1**·4HBr or **L2**·6HBr plus other components such as KCl solution. The concentration of the sample solution was 1×10^{-3} M for **L1**·4HBr or **L2**·6HBr. The pH range for accurate measurements was considered to be 2–12. The pK_w for the aqueous system, defined as $-\log$ ([H][OH]) at the ionic strength employed was found to be 13.78. Protonation constants from the direct titrations were calculated from the potentiometric data with the program BEST.

The error in the constants are estimated as ±0.04 log units on the basis of the σ_{off} value, which measures the deviation of the experimental curve and the curve calculated from the equilibrium constants, being less than 0.01 pH unit in all potentiometric determinations. Species distribution diagrams were computed from the measured equilibrium constants with SPE and plotted with SPEPLOT.[12]

Synthesis of L1·4HBr

A solution of 2,5-thiophenedicarboxaldehyde (560.4 mg, 4 mmol) in 200 ml MeCN was added dropwise from a dropping funnel to a stirred solution of 1*R*,2*R*-diaminocyclohexane (4 mmol) in the presence of $Pb(SCN)_2$ (1.35 g, 4.17 mmol). The yellow solid formed was filtered off, washed with methanol and chloroform, and dried *in vacuo*. The precipitate was suspended in 100 ml of MeOH and solid $NaBH_4$ (2 g, 50 mmol) was added in small portion at 0 °C in an ice bath, over a period of 2 h. The suspension was magnetically stirred for 2 h more at room temperature and then gently heated to 50 °C to ensure the reaction was complete. The yellow-colored solution that resulted was filtered to remove any suspended material. The filtrate was diluted with water (300 ml), acidified (pH =2.0) with H_2SO_4 (8 M), and then was filtered off. The filtrate was treated with aqueous ammonia (30 ml). The solution that resulted was then extracted with chloroform (3 × 100 ml). The combined organic layer, after washing with water and dry Na_2SO_4, was rotary evaporated nearly to dryness. The crude product was separated as **L1** and **L2** by column chromatography (silica gel, CH_2Cl_2–MeOH ~ 2 : 8). **L1** was then dissolved in 20 ml of EtOH and 5 ml 48% HBr in 10 ml of EtOH was added solowly. The light yellow microcrystals were obtained as the tetrahydrobromide salt, **L1**·4HBr (yield, 15%). Mp > 298 °C; $[\alpha]_D^{25} = -110.4$ (*c* 1, CH_2Cl_2); ESI-MS, *m*/*z* 445.24 (M + H)$^+$; calcd for $C_{24}H_{36}N_4S_2$: 444.24; ^1H NMR (300 MHz, D_2O) δ 7.16(s, 4H, thiophene), 4.32–4.56(m, 8H, CH_2–NH), 2.27(s, 4H, CH–NH of cyclohexane), 1.73–1.33(m, 16H, cyclohexane); ^{13}C NMR (D_2O–CD_3OD), δ 134.85(CH–CHS of thiophene), 132.08(CH–CHS of thiophene), 57.49(CH_2–NH), 43.71 (CH–NH– of cyclohexane), 26.32 (cyclohexane), 22.21(cyclohexane); anal calc. for $C_{24}H_{36}N_4S_2$·4HBr C, 37.54; H, 5.25; N, 7.30. Found. C, 37.51; H, 5.22; N, 7.29%.

Synthesis of L2·6HBr

L2 was then dissolved in 20 ml of EtOH and 5 ml 48% HBr in 10 ml of EtOH was added solowly. The yellow microcrystals were obtained as the hexahydrobromide salt, **L2**·6HBr (yield, 46%). mp > 298 °C; $[\alpha]_D^{25} = -134.0$ (*c* 1, CH_2Cl_2); ESI-MS, *m*/*z*: 667.34 (M + H)$^+$; calcd for $C_{36}H_{54}N_6S_4$: 666.36; ^1H NMR, (D_2O) δ: 7.16(s, 2H, thiophene), 4.72–4.42(m, 4H, CH_2–NH), 2.41(s, 2H, CH–NH of cyclohexane), 1.97–1.22(m, 8H, cyclohexane); ^{13}C NMR (D_2O–CD_3OD) δ 134.84(CH–CHS of thiophene), 132.10(CH–CHS of thiophene), 57.45(CH_2–NH), 43.71 (CH–NH– of cyclohexane), 26.28 (cyclohexane), 22.20(cyclohexane); anal calc. for $C_{36}H_{54}N_6S_4$·6HBr, C, 37.54; H, 5.25; N, 7.30. Found. C, 37.45; H, 5.12; N, 7.24%.

Synthesis of L3

This compound was synthesized by heating 1.0 mmol of **L2** and 4.0 mmol of bromomethylbenzene in the presence of 4.0 mmol of K_2CO_3. The mixture was allowed to react for an additional 8 h at 60 °C to give a light-yellow organic compound which was then was purified by column chromatography. ^1H NMR. ($CDCl_3$) δ 7.23–7.20(m, 20H, beneze), 7.16–7.12(m, 4H, CH=CHS), 3.80(m, 8H, CH_2–NH), 3.84(m, 8H, CH_2–benzene), 2.84(m, 4H, CH–NH of cyclohexane), 1.74–1.39 (m, 16H, cyclohexane). ESI-MS *m*/*z*: 805.44 (M + H)$^+$;anal. calcd for $C_{52}H_{60}N_4S_{2:}$ 804.43.

Acknowledgements

This research program was supported by a grant A-259 from the Welch Foundation.

References

1 P. N. W. Baxter, *J. Org. Chem.*, 2001, **66**, 4170.
2 Y. H. Kim, N. R. Cha and S. K. Chang, *Tetrahedron Lett.*, 2002, **43**, 3883.
3 R. J. Motekaitis, S. Y. Sun, A. E. Martell and M. J. Welch, *Can. J. Chem.*, 1999, **77**, 614.
4 S. Y. Sun, D. Chen, A. E. Martell and M. J. Welch, *Inorg. Chim. Acta*, 2001, **324**, 180.
5 J. Gao, A. E. Martell and J. Reibenspies, *Inorg. Chim. Acta*, 2002, **329**, 122.
6 J. Gao, J. Reibenspies and A. E. Martell, *Inorg. Chim. Acta*, 2002, **335**, 125.
7 A. M. Groth, L. F. Lindoy and G. V. Meehan, *J. Chem. Soc., Perkin Trans. 1*, 1996, 1553.
8 P. C. Reisen and T. A. Kaden, *Helv. Chim. Acta*, 1995, 1325.
9 C. Wieser-Jeunesse, D. Matt and A. D. Cian, *Angew. Chem., Int. Ed.*, 1998, **37**, 2861.
10 B. Kersting, *Angew, Chem., Int. Ed.*, 2001, **40**, 3988.
11 A. E. Martell, R. M. Smith and R. J. Motekaitis, *NIST Critically Selected Stability Constants of Metal Complexes*, Gaithersburg, MD 20899, USA, 2001.
12 A. E Martell and R. J. Motekaitis, *Determination and Use of Stability Constants*, 2nd edn., VCH, New York, 1992.
13 J. Bourson and B. Valeur, *J. Phys. Chem.*, 1989, **93**, 3871.
14 R. G. Pearson, *Coord. Chem. Rev.*, 1990, **100**, 403.
15 R. G. Pearson, *J. Am Chem. Soc.*, 1963, **85**, 3533.
16 *Applied fluorescence in chemistry, biology and medicine*, eds W. Retting, B. Strehmel, S. Schrader and H. Seifert, Springer, Berlin, New York, 1999.
17 M. Shibasaki, H. Sasai and T. Arai, *Angew Chem., Int. Ed*, 1997, **36**, 1236.
18 B. M. Trost and V. S. C. Yeh, *Angew Chem., Int. Ed.*, 2002, **41**, 861.