(H_2O), 97878-24-5; $Fe^{III}T4MP(H_2O)$, 65774-47-2; $Fe^{III}T2MP(OH)$, 97889-69-5; $Fe^{III}T3MP(OH)$, 97878-25-6; $Fe^{III}T4MP(OH)$, 97889-58-2; $Fe^{III}T2MP(OH)_2$, 97878-26-7; $Fe^{III}T3MP(OH)_2$, 97889-59-3; $Fe^{III}T4MP(OH)_2$, 75908-34-8; $Fe^{II}T2MP(H_2O)$, 97878-27-8; $Fe^{II}T3MP(H_2O)$, 97878-28-9; $Fe^{II}T4MP(H_2O)$, 97878-29-0; $Fe^{II}T2MP(OH)(H_2O)$, 97878-30-3; $Fe^{II}T3MP(OH)(H_2O)$, 97878-31-4; $Fe^{II}T4MP(OH)(H_2O)$, 75908-36-0; $Fe^{III}T2MP(Cl)$, 97878-32-5; $Fe^{III}T3MP(Cl)$, 97878-33-6.

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Noncyclic Reference Ligands for Tetraaza Macrocycles. Synthesis and Thermodynamic Properties of a Series of α,ω-Di-N-methylated Tetraaza Ligands and Their Copper(II) Complexes

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Received August 1, 1984

The synthesis and characterization of 2,5,8,11-tetraazadodecane, 2,5,9,12-tetraazatridecane, 2,6,9,13-tetraazatetradecane, and 2,6,10,14-tetraazapentadecane are described. The protonation constants (pK_{1-4}) for these ligands have been determined together with the stepwise enthalpies of protonation (ΔH_{1-4}). ^{13}C NMR provides evidence for the sequence of protonation of the four nitrogen atoms. K_{ML} and ΔH° values for the formation of the copper(II) complexes have been determined, and it has been shown that the behavior of this series of di-N-methylated ligands toward both protonation and copper(II) complex formation parallels that of the simple tetraaza macrocyclic ligands much more closely than that of noncyclic reference ligands containing terminal primary amine groups. With regard to the macrocyclic effect, both free energy and enthalpy terms are essentially independent of the size of the ligands.

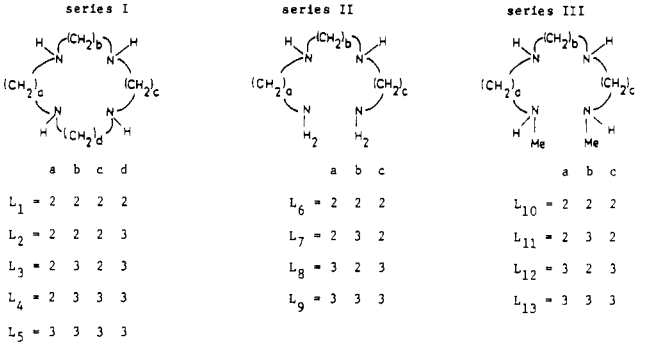
Introduction

Considerable interest has been shown in recent years in the thermodynamic origins of the macrocyclic effect, the term used to describe the enhanced stability of the metal complexes of a macrocyclic ligand over similar complexes of an analogous noncyclic ligand.[3] Early work ascribed the macrocyclic effect exclusively to either enthalpy[4] or entropy[5,6] terms, although more recently, by use of direct calorimetric values of enthalpy,[7,10] it has been shown, at least for tetraaza ligands, that the entropy term is always favorable whereas the enthalpy term varies in magnitude with the matching of size of the metal ion and the aperture in the macrocyclic ligand. In assessing the magnitude of the macrocyclic effect, represented by the metathetic reaction

$$[ML_{noncyclic}]^{n+}(aq) + L_{cyclic}(aq) \rightarrow [ML_{cyclic}]^{n+}(aq) + L_{noncyclic}(aq)$$

it is necessary, for any particular macrocyclic ligand, to choose a suitable noncyclic reference. Traditionally, when considering tetraaza macrocyclic ligands in series I (see Chart I), different authors have used a noncyclic ligand from series II and have tried to ensure that the sequencing of the chelate rings is the same. Thus the thermodynamic properties of complexes of L_1 have been compared to those of L_6[8] (all chelate rings being five-membered) whereas the properties of complexes of L_3 have been compared to those of L_7[4] (the alternating sequence of five- and six-membered rings remains the same). This choice of reference ligand has had the added advantage that the ligands themselves were readily available and the thermodynamic properties of their metal complexes had already been determined.

Chart I



A more critical appraisal, however, suggests that ligands from series II may not be the most appropriate choice for two reasons.[11] First, the macrocyclic ligands contain four secondary amine nitrogen donor atoms compared with two secondary and two primary amine nitrogens for the reference ligands. It has already been established for complexes with N-alkylated ethylenediamines that changing a primary amine nitrogen to a secondary by, e.g., N-methylation causes a significant change in both ΔG°[12] and ΔH°.[13] Second, the macrocyclic enthalpy term (ΔH_{mac}) is made up of three component terms, as shown in the thermochemical cycle of Scheme I. The second of these terms is the difference in solvation enthalpies of the two complexes and is virtually impossible to determine experimentally.[14] In an experimental investigation into the contribution that the solvation of the free ligands makes to the macrocyclic enthalpy, it is, therefore, very desirable to minimize any difference between the hydration enthalpies of the two complexes.[15] Enthalpies of hydration are dependent on both ionic

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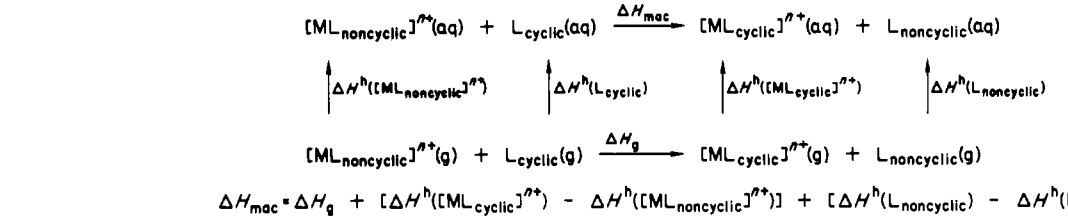
Scheme I

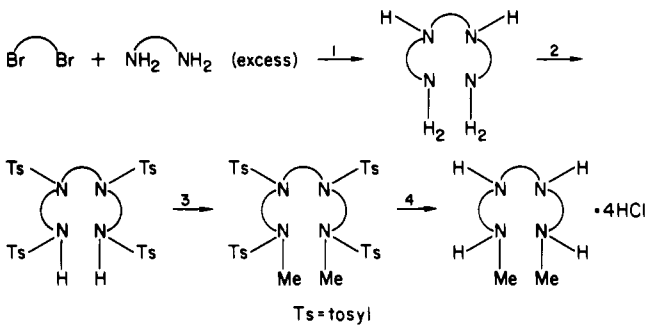


charge and ionic radius. In order, therefore, to keep the term relating to the difference in hydration enthalpies of the complexes as small as possible, it would appear to be very desirable to try to ensure that the radius of the two complexes should be very similar. For these two reasons, therefore, we are suggesting that the ligands in series III, which contain N-methylated terminal amine groups, would be more appropriate choices as reference ligands when considering the magnitude of the macrocyclic effect. L_3 would now be compared with L_{11}, both ligands having four secondary amine nitrogen donor atoms, with the added expectation that $[CuL_3]^{2+}$ and $[CuL_{11}]^{2+}$ should have more closely matched ionic radii than $[CuL_3]^{2+}$ and $[CuL_7]^{2+}$ due to the greater similarity in composition of the two ligands.

Preliminary work on L_{10} indicated that there was no significant difference between the thermodynamic properties of the complexes of L_6 and L_{10}.[11] This work completes this study by considering the three remaining ligands in series III, namely L_{11}, L_{12}, and L_{13}, and shows that the conclusions drawn on the basis of L_{10} alone are not general to other members of this series.

Experimental Section

Ligand Synthesis (L_{10}, L_{12}, and L_{13}). L_{10}, L_{12}, and L_{13} were synthesized by the following general procedure:



Ts = tosyl

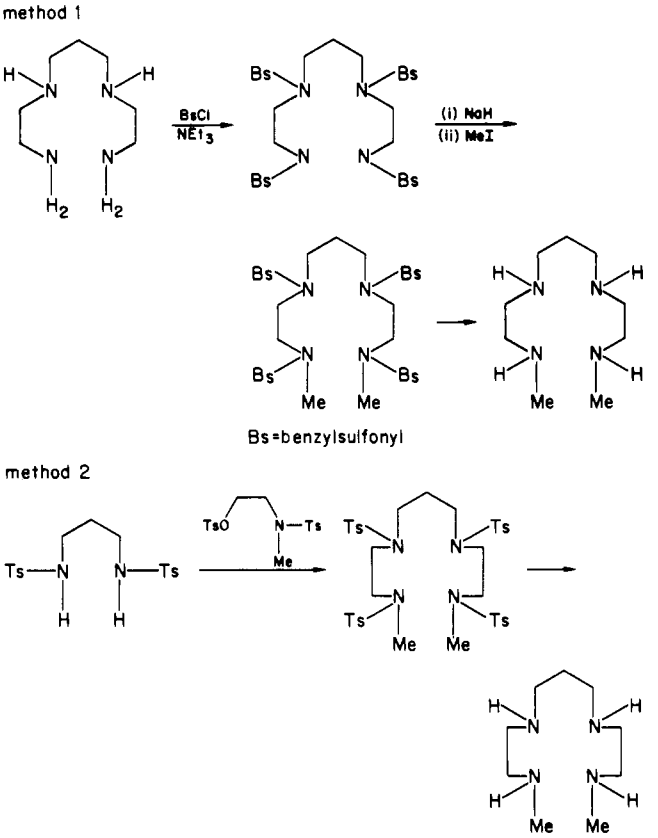
Step 1 was carried out under existing literature conditions,[16] and the resulting tetraamines were purified by fractional distillation under reduced pressure. A sample of each was converted to the corresponding tetrahydrochloride salt by addition of concentrated HCl to a solution of the amine in methanol. A satisfactory elemental analysis was obtained in each case on the tetrahydrochlorides.

Step 2 was accomplished by adding, dropwise, with stirring, a solution of *p*-toluenesulfonyl chloride (4 equiv) in diethyl ether to an aqueous solution of the tetraamine containing 4 equiv of NaOH. The resulting suspension was stirred for 2 h and filtered, and the solid product was washed with water followed by methanol and dried in vacuo. Infrared and ^1H NMR spectra were consistent with their proposed formulations, and the products were used without further purification. A sample of each was recrystallized from methanol and gave satisfactory elemental analysis.

Step 3. An excess of NaH (50% suspension) was added in small portions with stirring to a solution of the tetrakis(tosylamine) in dry DMF (20 g in 250 mL). After the mixture was stirred for 30 min in a flask protected from atmospheric moisture, excess NaH was removed by filtration. The filtrate was heated to 120 °C under anhydrous conditions and a 100% excess of CH_3I added over 30 min. The solution was kept at 100–120 °C for 8 h, during which time a deep brown color developed. The reaction mixture was cooled on ice. The volume of the filtrate was reduced by 50% under reduced pressure, and the residual solution was poured slowly, with vigorous stirring, onto 3 times its volume of ice/water. The resulting yellow precipitate was filtered, washed with water, and dried in vacuo. Infrared and ^1H NMR spectra were consistent with the expected products.

Step 4 was accomplished by heating the tetrakis(tosylamines) for varying lengths of time with sulfuric acid to temperatures not greater than 140 °c. The reaction mixture was cooled on ice and filtered, and the solids were washed with ethanol followed by diethyl ether. The washings and filtrate were combined and refrigerated to yield a further small batch of product. The resulting tetrakis(hydrogen sulfate) salts were converted into the corresponding tetrahydrochloride salts by dissolving the impure products in water and adding an excess of $BaCO_3$ in small portions, with stirring. The resulting suspension was warmed until no further CO_2 evolution was observed and filtered. Excess concentrated HCl was added to the filtrate. The solution was taken to dryness, and the resulting solids were recrystallized from aqueous ethanol. A second recrystallization from the same solvent yielded analytically pure tetrahydrochloride salts. The combined yields for steps 3 and 4 were over 85% in each case.

Synthesis of L_{11}:



Bs = benzylsulfonyl

Method 1. L_7 (synthesized as step 1 above) was treated with a 10% excess of benzylsulfonyl chloride in triethylamine. The resulting mixture, protected from atmosphere moisture, was stirred at room temperature for 24 h and then poured slowly, with vigorous stirring, onto 3 times its volume of ice/water. The precipitated solids were filtered out, washed with water followed by cold methanol, and recrystallized from methanol (yield 30%). TLC on silica with a range of $CHCl_3$/diethyl ether eluent mixtures indicated a single product, confirmed by infrared and elemental analysis. Methylation, deprotection, and subsequent workup procedures were identical with those used for L_{10}, L_{12}, and L_{13}. Yield: 89%.

Method 2. *N*-Methylethanolamine (0.2 mol) was dissovled in 100 mL of triethylamine. Finely ground tosyl chloride (0.42 mol) was added in small portions with stirring over 30 min. The resulting suspension, protected from atmospheric moisture, was stirred at room temperature for

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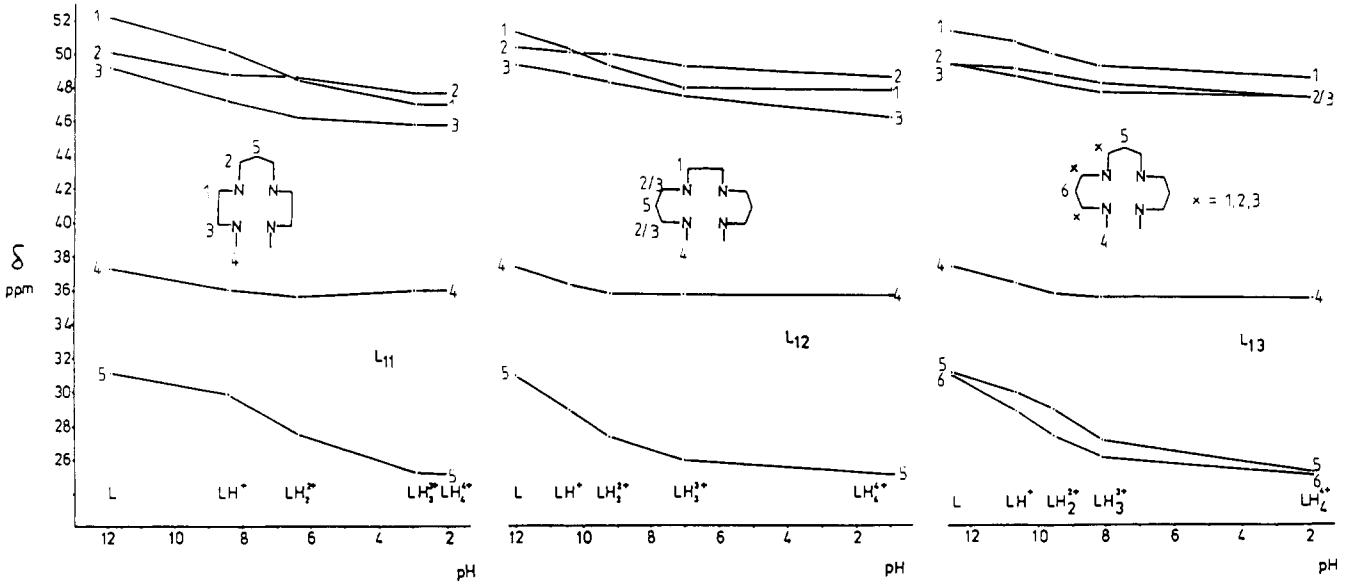

Figure 1. ^{13}C chemical shifts as a function of pH.

Table I

	% found			% calcd		
	C	H	N	C	H	N
L_{10}·4HCl ($C_8H_{26}N_4Cl_4$)	30.10	8.21	17.60	30.03	8.13	17.51
L_{11}·4HCl ($C_9H_{28}N_4Cl_4$)	32.38	8.70	16.70	32.34	8.45	16.77
L_{12}·4HCl ($C_{10}H_{30}N_4Cl_4$)	34.50	8.70	15.80	34.49	8.67	16.09
L_{13}·4HCl ($C_{11}H_{32}N_4Cl_4$)	36.50	8.90	15.00	36.48	8.91	15.47

6 h. The resulting pale brown solution was poured onto 500 mL of ice/water with stirring, the mixture filtered, and the precipitate washed with water. The product, *N,O*-ditosyl-*N*-methylethanolamine, was recrystallized first from acetone (including treatment with activated charcoal) and finally from ethanol, from which it was obtained as fine needles, which were dried in vacuo over P_2O_5. Satisfactory elemental analysis and infrared and ^1H NMR spectra were obtained. Yield: 48 g (62%).

N,N'-Ditosyl-1,3-propanediamine (0.05 mol) was dissolved in 200 mL of dry DMF; the solution was treated with NaH and filtered as described previously. A solution of 0.1 mol of *N,O*-ditosyl-*N*-methylethanolamine in 100 mL of dry DMF was added over 30 min with stirring. The resulting mixture, protected from atmospheric moisture, was heated to 110 °C with stirring, held at that temperature for 4 h, cooled on ice, and poured onto 750 mL of ice/water with stirring. The precipitate was filtered, washed with water, and dried at 70 °C overnight. No suitable solvent for recrystallization was found, as an oil often resulted. The product was used without further purification. Yield: 60%.

Detosylation and subsequent conversion to the tetrahydrochloride salt yielded a product indistinguishable from that of method 1. Yield: 89%.

Infrared, ^1H and ^{13}C NMR spectra, and elemental analysis were consistent with the expected product.

Elemental analyses are given in Table I.

^1H NMR spectra (Perkin-Elmer R32 NMR spectrometer, 90 MHz ^1H, D_2O solvent; chemical shift values in ppm relative to Me_4Si): L_{10}·4HCl, 7.35 (12 H, m), 7.60 (6 H, s); L_{11}·4HCl, 6.45 (2 × 4 H, s), 6.65 (4 H, t), 7.15 (6 H, s), 7.75 (2 H, qi); L_{12}·4HCl, 6.45 (2 × 2 H, s), 6.75 (2 × 4 H, t), 7.25 (6 H, s), 7.75 (4 H, qi); L_{13}·4HCl, 6.8 (12 H, m), 7.25 (6 H, s), 7.85 (6 H, m).

^{13}C NMR/pH Studies. The pH values required to give the optimum concentration of each of the protonated species, LH_n^{n+}, relative to the other protonated species in solution were calculated from the protonation constants for the ligands L_{11}, L_{12}, and L_{13}. ^{13}C spectra were recorded at each of the calculated pH values by using a Bruker WP80 FT NMR spectrometer at 20.13 MHz. H_2O/D_2O (4/1) was used as solvent and no correction for pH/pD differences was made. The ligand tetrahydrochloride was dissolved in the H_2O/D_2O with a slight excess of HCl present and the spectrum corresponding to LH_4^{4+} recorded. A concentrated solution in H_2O of AR NaOH was added to bring the solution to the next calculated pH, the spectrum recorded, and the process repeated until a pH greater than 11.5 was reached, at which no protonated species are present in any significant amounts. As a check on dilution effects, a similar sample was treated with H_2O instead of NaOH, and no change in the chemical shifts was observed even after 50% dilution of the original sample. The chemical shifts are shown as a function of pH in Figure 1.

Table II

		% found			% calcd		
		C	H	N	C	H	N
$[CuL_{10}](ClO_4)_2$	blue	22.50	5.01	12.45	22.00	5.07	12.83
	purple	22.32	4.99	12.48			
$[CuL_{11}](ClO_4)_2$		23.78	5.20	12.31	23.98	5.37	12.43
$[CuL_{12}](ClO_4)_2$	blue	25.60	5.61	11.43	25.84	5.64	12.05
	purple	22.54	5.61	12.03			
$[CuL_{13}](ClO_4)_2$		27.60	5.89	11.67	27.59	5.89	11.70

Table III

			vis		
			soln		solid
		IR $\bar{\nu}$(N–H), cm^{-1}	λ_{max}, nm	ϵ, mol^{-1} dm^3 cm^{-1}	λ_{max}, nm
$[CuL_{10}](ClO_4)_2$	blue	3210, 3180, 3140	580	200	
	purple	3240	578	215	560
$[CuL_{11}](ClO_4)_2$		3225, 3200	536	87	520
$[CuL_{12}](ClO_4)_2$	blue	3260, 3220, 3170	545	142	530
	purple	3260, 3130	545	186	580
$[CuL_{13}](ClO_4)_2$		3190, 3130, 3080	629	183	593

Synthesis of Copper(II) Complexes, $[CuL_n](ClO_4)_2$. (a) n = 10, 11, 13. The appropriate ligand tetrahydrochloride (1.5 g) was dissolved in 3 cm^3 of water, and the solution was passed through an anion-exchange column (Amberlite IR45, OH^- form) directly onto a solution of $Cu(ClO_4)_2·6H_2O$ (1 equiv) in water. The volume of the resulting solution was reduced until crystallization commenced. For L_{11} and L_{13} single products were recovered, which were recrystallized and dried in vacuo over P_2O_5. For L_{10}, both a blue and a purple form were isolated by fractional crystallization, although it was not possible to free the blue form totally from the purple.

(b) n = 12. L_{12}·4HCl (2.0 g) was dissolved in methanol (100 cm^3), and the solution was stirred over crushed NaOH pellets for 3 h. The resulting solution was filtered, and 1 equiv of $Cu(ClO_4)_2·6H_2O$ in hot methanol was added. This solution was filtered while hot; a blue precipitate forming as the filtrate cooled was removed by filtration. On concentration of the filtrate and washings, a purple product formed in approximately equal quantity. Both solids were washed with diethyl ether and dried in vacuo.

Elemental analyses are given in Table II.

The IR spectra differed most predominantly in the N–H stretching region, and this is shown in Table III for the copper complexes together with electronic spectral data in the visible region.

Protonation and Complex Formation Constants and Enthalpies of Protonation. All potentiometric and calorimetric measurements were carried out in 0.5 M KNO_3 as an inert ionic medium, using commercial KNO_3 (Merk Suprapur) without further purification. Standardized CO_2-free solutions of NaOH, used in the potentiometric and calorimetric

Table IV. Thermodynamic Parameters for Protonation of the Ligands in Series II and III[a]

	L_{10}[b]	L_6[c]	L_{11}[b]	L_7[c]	L_{12}[b]	L_8[c]	L_{13}[b]	L_9[c]
log K_1	10.75 (1)	9.95	10.304 (5)	10.25	10.841 (5)	10.665	10.946 (6)	10.45
log K_2	9.68 (1)	9.31	9.571 (5)	9.50	10.123 (5)	9.956	10.226 (5)	9.88
log K_3	7.59 (7)	6.86	7.179 (5)	7.28	8.443 (5)	8.536	8.894 (5)	8.54
log K_4	4.08 (8)	3.66	5.660 (7)	6.02	5.693 (5)	5.837	7.426 (5)	7.22
log β_4	32.10	29.78	32.71	33.05	35.10	34.99	37.49	36.03
$-\Delta H_1$	49.4 (4)	46.07	49.0 (4)	46.0	51.9 (4)	51.17	53.1 (4)	51.04
$-\Delta H_2$	41.2 (4)	47.15	42.7 (4)	47.3	48.1 (4)	53.51	47.7 (4)	52.17
$-\Delta H_3$	42.7 (4)	39.87	41.4 (4)	41.8	43.5 (4)	44.81	48.5 (4)	48.74
$-\Delta H_4$	30.9 (3)	28.58	42.2 (4)	38.5	41.4 (4)	40.88	44.4 (4)	45.52
$-\sum\Delta H$	164.2	161.7	175.3	173.6	184.9	190.4	193.7	197.5
ΔS_1	4.02 (12)	35.9	32.9 (12)	41.9	33.5 (12)	32.6	31.5 (12)	28.8
ΔS_2	47.3 (12)	20.0	39.9 (12)	23.2	32.5 (12)	10.9	35.9 (12)	14.1
ΔS_3	2.0 (12)	−2.4	−1.3 (12)	−0.9	15.8 (12)	13.0	7.7 (12)	0.0
ΔS_4	−25.5 (12)	−25.8	−33.2 (12)	−13.9	−29.9 (12)	−25.5	−6.7 (12)	−14.5
$\sum\Delta S$	64	28	38	50	52	31	68	28

[a] ΔH in kJ mol^{-1}; ΔS in J mol^{-1} K^{-1}. [b] This work. [c] See references contained in ref 12.

Table V. Thermodynamic Parameters for Cu(II) Complex Formation with Tetraaza Ligands at 298 K[a]

ligand	log K_{ML}	$-\Delta G°$	$-\Delta H°$	$T\Delta S°$	ref K_{ML}	ref ΔH
			Series I			
L_1	24.8	141.5	95.0	46.5	5, 6	8
L_2	29.1	166.0	107.1	58.9	28, 29	9
L_3	27.2	155.2	135.6	19.6	30	7
L_4	24.4	139.2	110.9	28.3	31	8
L_5	20.9	119.3	83.7	35.6	10	10
			Series II			
L_6	20.2	115.3	90.4	24.9	32	32
L_7	23.9	136.4	115.9	20.5	33	33, 34
L_8	21.8	124.5	108.4	16.1	19	19
L_9	17.3	98.8	81.6	17.2	27	27
			Series III			
L_{10}	20.9 (1)	119.3 (5)	88.3 (12)	31.0 (17)	11	11
L_{11}	21.89 (1)	124.9 (1)	108.4 (5)	16.6 (6)	this work	this work
L_{12}	18.50 (2)	105.6 (1)	93.1[b]	12.5[b]	this work	this work
L_{13}	14.62 (2)	83.4 (1)	62.3 (2)	21.5 (3)	this work	this work

other species: $[CuL_{12}H]^{3+}$, log β_{MLH} = 24.33 (1); $[CuL_{13}H]^{3+}$, log β_{MLH} = 20.97 (3); $[CuL_{13}H_2]^{4+}$, log β_{MLH} = 27.02 (5)

[a] $\Delta G°$, $\Delta H°$, and $T\Delta S°$ in kJ mol^{-1}. [b] Average of blue and purple.

titrations, were prepared according to standard procedures. The potentiometric titrations were carried out with a fully automatic apparatus as described in ref 17. The experimental values of the emf have not been corrected for the liquid-junction potential because this effect was found to be negligible in the pH range investigated. The computer program MINIQUAD[18] was used to process the data and calculate the basicity and complex stability constants.

The enthalpies of protonation were determined with an LKB calorimeter (Model 8700), by employing the continuous-titration technique. The apparatus, experimental procedures, and computer programs necessary to determine the stepwise enthalpy of protonation values have already been reported.[19,20] The enthalpy of ionization of water was determined by adding NaOH solution to a solution of HCl contained in the calorimetric vessel. The value found, 56.40 (5) kJ mol^{-1} was in good agreement with the accepted literature value.[21] The thermodynamic results for the protonation of L_{10}, L_{11}, L_{12}, and L_{13} are reported in Table IV, and the formation constants (log K_{CuL}) for the copper complexes, in Table V.

Enthalpies of Complex Formation. All calorimetric measurements were performed on an LKB Batch Microcalorimeter (10700-2), by using previously described procedures.[8] One molar NaOH was used to prevent protonation of the strongly basic ligands in solution. (Typically, pK_4 values are in the range 10.5–11.0.) The heat output was determined by comparison of the area under the reaction heat curve with that of a similar reaction of known heat output, in this case the fomation of $[CuL_7]^{2+}$. As a check on absolute accuracy, the enthalpy of formation

Scheme II

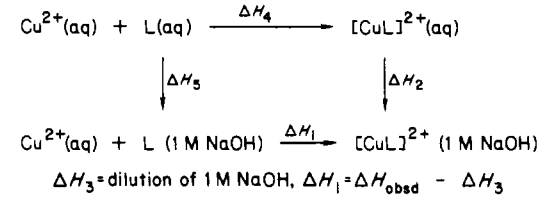

Table VI. Calorimetric Terms in the Thermochemical Cycle Used for Determining Enthalpies of Complex Formation[a]

ligand	$-\Delta H_1$	ΔH_2	ΔH_3	$-\Delta H_4$	lit.
L_7[b]					115.9[12]
L_8	108.0 (3)	0	2.1	108.0 (4)	108.4[12]
L_{10}	87.0 (1)	1.1 (1)	2.1	88.1 (2)	88.3[12]
L_{11}	105.9 (4)	2.5 (1)	2.1	108.4 (5)	
L_{12}	84.3 (8)	9.6 (1) (blue), 7.9 (1) (purple)	2.1	93.9 (9) (blue), 92.2 (9) (purple)	
L_{13}	52.7 (1)	9.6 (1)	2.1	62.3 (2)	

[a] ΔH in kJ mol^{-1}. [b] Used as calibrant.

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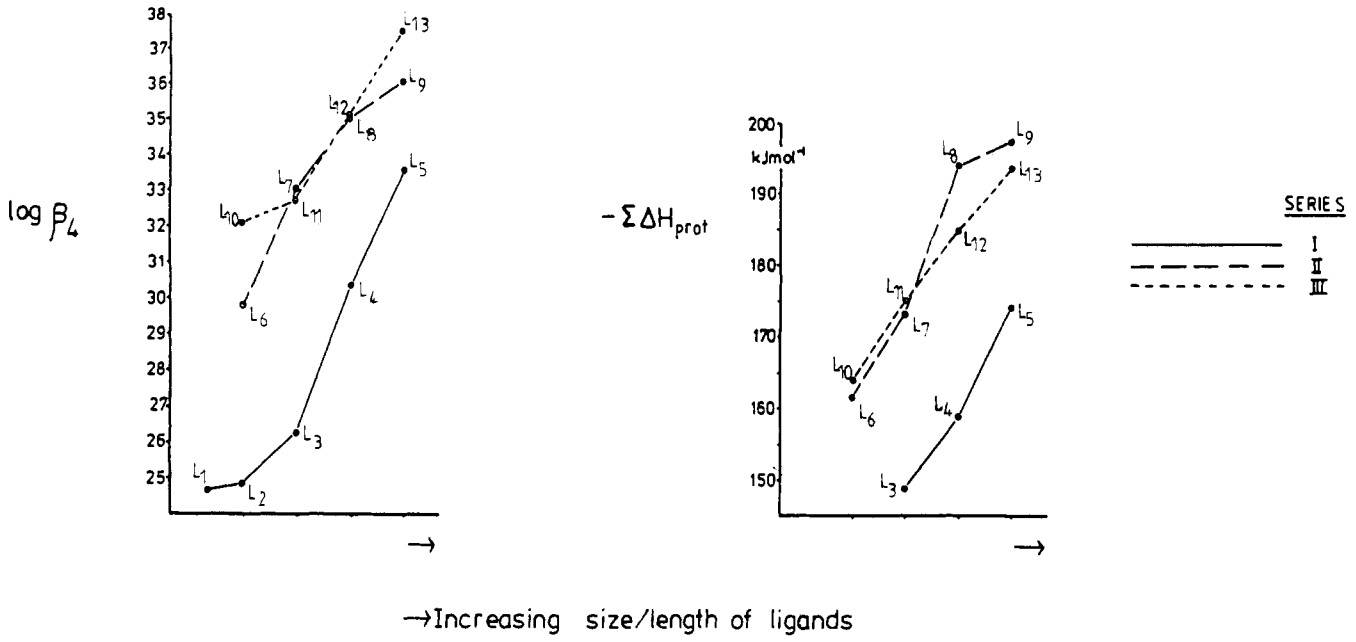

Figure 2. Overall protonation constants and enthalpies of protonation.

of $[CuL_8]^{2+}$ was determined. The results are in good agreement with existing literature values (measured −108.0 (4) kJ mol^{-1}; lit.[19] −108.4 kJ mol^{-1}). The results are related to the thermochemical cycle in Scheme II and presented in Table VI.

Discussion

1. Sequence of Protonation. In order to deduce the sequence of protonation, the ^{13}C NMR spectra of the ligands in series III were recorded as a function of pH. The chemical shifts are shown as a function of pH in Figure 1. It is well documented that the chemical shift of a carbon atom in a position β to the center of a reaction is often more sensitive to changes in the nature of that reaction center than one in an α or γ position; this phenomenon has been called the β-shift.[22] In the case of the polyamines under consideration, we would therefore expect that the major chemical shift changes observed on protonation would correspond to carbon atoms in a position β to the nitrogen concerned. For L_{13}, the only β-carbons are labeled 5 and 6, the resonances being assigned on the basis of the relative intensities and chemical shift values. If a terminal nitrogen becomes protonated, a shift in $\delta(6)$ should be observed, with little change in the other signals, whereas if a midchain nitrogen is protonated first, roughly equal changes in both $\delta(5)$ and $\delta(6)$ should be seen. The results indicate a more rapid change in $\delta(6)$ with decreasing pH, consistent with the first two protons becoming bound to the terminal nitrogens. The fact that $\delta(5)$ displays a shift on the addition of the first two protons is almost certainly due to proton exchange between the basic centers, in most cases a relatively rapid process on the NMR time scale. The *N*-methyl resonance, although less sensitive to change than that of the β-carbons, displays its largest change in chemical shift on the addition of the first two protons, changing by 1.5 ppm out of a total change of 1.7 ppm for the addition of all four protons, again indicating that the first two protons are, to a first approximation, located on the terminal nitrogens. A similar consideration of the L_{11} and L_{12} spectra indicates this same sequence of protonation. Similar studies have not been carried out on the macrocyclic ligands due to the similarity of the first two and the last two pK values for these ligands.

2. Protonation Thermodynamics. The observed thermodynamic parameters given in Table IV are presented in Figure 2 for the overall protonation of the ligands (log β_4 and $\sum\Delta H$). All three series of ligands display the same general trend in log β_4 as the carbon-bridge length increases. Comparable ligands in series II and III have similar log β_4 values, both substantially higher than that of the appropriate cyclic ligand in series I. This can be readily explained in terms of charge-repulsion effects. Series III appears to parallel series I more closely than does series II.

The observed differences between the ligands in series II and III can be compared with the existing data for *N*-methylethylenediamine (meen) and *N*,*N*′-dimethylethylenediamine (dmen). In going from a ligand in series II (e.g., L_6) to the corresponding ligand in series III (L_{10}), it is necessary to convert the two terminal primary amine groups into secondary (N-methylated) amine groups, the two nonterminal amine groups remaining secondary in both series of ligands. This is therefore equivalent to comparing the thermodynamics of protonation of 2 mol of meen with 2 mol of dmen.

The values of 2 log β_2 for *N*-methylethylenediamine (meen) and *N*,*N*′-dimethylethylenediamine (dmen) are 34.8 and 34.9, respectively.[12] The difference between these values (0.1) might therefore be expected to be comparable to the difference in log β_4 values between ligands of series II and series III. However, the differences between L_{10} and L_6 and between L_{13} and L_9 are significantly larger than this, whereas the differences between L_7 and L_{11} and between L_8 and L_{12} are close to this value. With repect to the enthalpies of protonation, series II shows a steady increase along the series, $\sum\Delta H$ increasing by approximately 10 kJ mol^{-1} for each additional methylene carbon. As was the case with the log β_4 values, this general increase can be explained in terms of charge repulsion and the increasing inductive effect of alkyl bridges.

In order to rationalize differences between ligands in series II and III, it is covenient to consider the metathetic reaction

$$L^{II}H_4^{4+} + L^{III} \rightarrow L^{II} + L^{III}H_4^{4+}$$

(L^{III} is a member of series III and L^{II} the corresponding member of series II.) The ΔH and $T\Delta S$ values for this reaction are as follows (all values in kJ mol^{-1}):

system	ΔH	$T\Delta S$	system	ΔH	$T\Delta S$
L_6/L_{10}	−2	+17	L_8/L_{12}	+5	+6
L_7/L_{11}	−1	−4	L_9/L_{13}	+4	+12

Estimated standard deviation for ΔH is 3 kJ mol^{-1}.

With the exception of the L_7/L_{11} pair, the metathetic reaction proceeds with a small positive change in entropy and negligible change in enthalpy.

The two most probable causes for this positive ΔS term would seem to be either (a) $L^{III}H_4^{4+}$ ion larger than the $L^{II}H_4^{4+}$ ion, the larger ion expected to be less heavily solvated, resulting in a release of bound solvent or (b) L^{III} more heavily solvated than L^{II}, again

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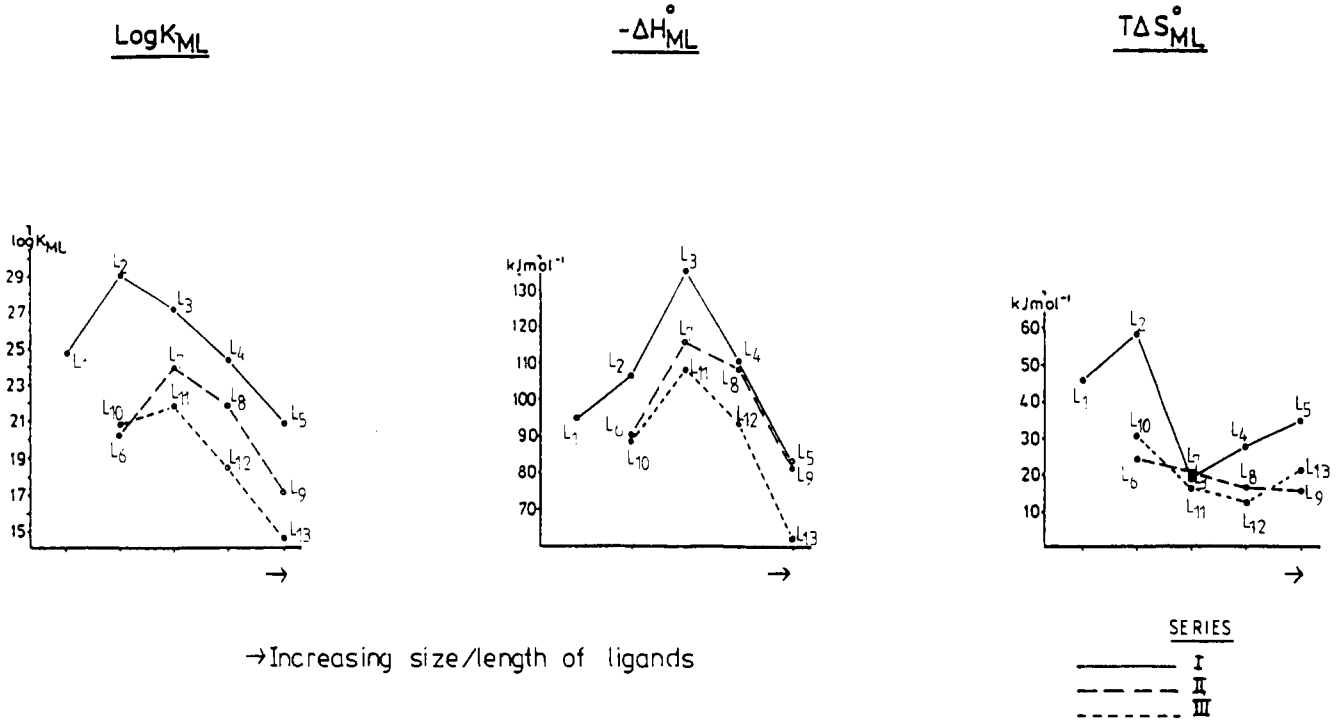

Figure 3. Thermodynamic parameters for Cu(II) complex formation.

resulting in the liberation of bound solvent, or most likely a combination of both. In any event, the solvation of the ligand or protonated ligand species is expected to be significant.

It has previously been demonstrated that the log K_{1-4} values for L_6 could be calculated with reasonable success if it was assumed that the primary nitrogens were protonated first. The method used to calculate these values was an empirical method, given by Clark and Perrin in a review of amine protonation.[23] The log K_{1-4} values were calculated for all members of both series II and series III and compared with the experimental values for the overall protonation. The excellent agreement between the calculated and observed values supports the NMR evidence that the initial protonation steps are at the terminal positions.

3. Complex Formation. When ligands in series III coordinate with metal ions, the presence of the two *N*-methyl groups gives rise to the possibility of stereoisomerism, which is not possible for ligands in series II. This can be seen in the isolation of differently colored copper complexes in the solid state for L_{10} and L_{12}. The thermodynamic data with which were are primarily concerned are obtained in strongly basic aqueous solution when we can assume that the most stable thermodynamic stereoisomer is formed. We have not, then, attempted to assign structures to the different solid stereoisomers, which would only be done reliably by X-ray crystallography. The enthalpy difference between the stereoisomers is not likely to be more than a few kJ mol^{-1},[24] and in the only case we investigated, that of $[CuL_{12}]^{2+}$, we found an enthalpy difference of only 1.7 kJ mol^{-1} between the blue and purple forms. The mean value has been used for this complex.

The thermodynamic properties for the formation of copper(II) complexes are given in Table V for ligand in series III (this work) as well as for the other ligands in series I and II (previous work). This is summarized in Figure 3, where it can be seen that for ligands in series III the complex formation enthalpies parallel those for ligands in series I, whereas for ligands in series II they converge on series I. L_{10} and L_6 were previously shown to have similar ΔH_{ML} values, and it was expected that this similarity between series II and III would continue. However, it is now clear that there are significant differences between the two systems for the larger members of each series. The observation that series III parallels series I, both for enthalpy and free energy terms, suggests that the factors causing the trends in series I also operate in series III but not in series II. The two main influences to be considered are almost certainly ring-size/steric effects and differences in ligand and complex solvation enthalpies.

It has been shown that L_3 has the correct "hole size" to accommodate the Cu^{2+} ion in a square-planar environment with optimum M–N interactions.[25] For the larger members of series I, the hole size becomes too big for optimal M–N bonding, and the ligand will adopt a less favorable configuration to accommodate the metal ion. The larger the ligand, the greater the degree of distortion required, and as a result, the enthalpy of complex formation will become less exothermic. Noncyclic ligands of series II and III should be more easily able to distort in order to accommodate the metal ion than the cyclic ligands in series I. However, from molecular models, it is evident that L_{13} and, to a lesser extent, L_{12} have an appreciable degree of steric hindrance associated with the methyl groups. With L_{13} any atttempt to give the maximum M–N interaction results in severe intraligand interactions, and the complex is therefore assumed to distort from its ideal geometry to relieve these steric effects. This distortion is almost certainly reflected in the enthalpy terms and, as such, models the distortion effects observed in series I. Such added effects are not observed in series II since there are no terminal methyl groups.

The presence of protonated complexes may well be a reflection of the adverse steric effects associated with coordination, the more strained systems being more readily protonated. Such monoprotonated species have been observed for Cu^{2+} complexes of L_4, L_5,[6,10] L_8,[26] and L_{12}, and both a mono- and a diprotonated species exist with L_{13}. No such species was observed for L_9,[27] and this

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seems surprising in the light of the other systems.

For the metathetic reaction

$$[CuL^{III}]^{2+}(aq) + L^{II}(aq) \rightarrow [CuL^{II}]^{2+}(aq) + L^{III}(aq)$$

the values of ΔH and $T\Delta S$ are compared as follows (values in kJ mol^{-1}; 25 °C) with those obtained from the reaction

$$[Cu(dmen)_2]^{2+}(aq) + 2meen(aq) \rightarrow [Cu(meen)_2]^{2+}(aq) + 2dmen(aq)$$

system	ΔH	$T\Delta S$	system	ΔH	$T\Delta S$
L_{10}/L_6	−2	+6	L_{13}/L_9	−20	+6
L_{11}/L_7	−8	−4	dmen/meen	−12	−2
L_{12}/L_8	−15	−4			

It is apparent that both the L_{11}/L_7 and L_{12}/L_8 systems follow the pattern of the dmen/meen system, whereas for the smaller and larger pairs the agreement is not so good, possibly reflecting the more severe distortions from square-plane geometry present in the complexes with ligands at the extremes of the series.

The purpose of this work was to prepare a series of noncyclic ligands that would serve as more appropriate reference ligands for the cyclic ligands in series I than the hitherto used ligands in series II. Considering the parallel behavior in both protonation and metal complex formation (Figure 3), it can be reasonably concluded that the new ligands in series III are more appropriate models with which to compare the macrocyclic ligands. However, the almost identical values that we previously observed for L_{10} and L_6 are not continued for the remainder of the series, where the difference between the thermodynamic parameters steadily increases with the length of the ligand skeleton.

The important consequence of considering the macrocyclic effect in relation to the reference ligands in series III rather than series II is that not only is the magnitude of the effect increased for both free energy and enthalpy terms but it also now remains approximately constant for different pairs of ligands along the series, as evident from the more or less parallel curves for log K and ΔH in Figure 3. Thus the magnitude of the macrocyclic effect appears to be independent of the size of the macrocyclic ligand provided the appropriate reference ligand is chosen. In further work we will be considering the effects of solvation on tetraaza ligands and their metal complexes, and we again expect, because of the greater similarity in both size and chemical constitution, that the ligands in series III will be more suitable models with which to compare the macrocyclic ligands in series I.

Acknowledgment. The authors wish to thank the Italian CNR, SERC, and NATO for financial support.

Registry No. L_6, 112-24-3; L_6·4HCl, 4961-40-4; L_6, tetratosyl derivative, 4961-40-4; L_7, 4741-99-5; L_7, tetratosyl derivative, 97807-55-1; L_7, tetrakis(benzylsulfonyl) derivative, 97807-53-9; L_8, 10563-26-5; L_8·4HCl, 13493-17-9; L_8, tetratosyl derivative, 74676-47-4; L_9, 4605-14-5; L_9·4HCl, 16632-23-8; L_9, tetratosyl derivative, 67341-38-2; L_{10}, 25077-85-4; L_{10}·4HCl, 27162-57-8; L_{10}, tetratosyl derivative, 97807-48-2; L_{11}, 6809-77-4; L_{11}·4HCl, 27162-58-9; L_{11}, tetrakis(benzylsulfonyl) derivative, 97807-54-0; L_{12}, 77500-19-7; L_{12}·4HCl, 97807-51-7; L_{12}, tetratosyl derivative, 97807-49-3; L_{13}, 97807-56-2; L_{13}·4HCl, 97807-52-8; L_{13}, tetratosyl derivative, 97807-50-6; $[CuL_{10}](ClO_4)_2$, 97825-42-8; $[CuL_{11}](ClO_4)_2$, 97825-44-0; $[CuL_{12}](ClO_4)_2$, 97825-46-2; $[CuL_{13}](ClO_4)_2$, 97825-48-4; $BrCH_2CH_2Br$, 106-93-4; $BrCH_2CH_2CH_2Br$, 109-64-8; $H_2NCH_2CH_2NH_2$, 107-15-3; $H_2NCH_2CH_2CH_2NH_2$, 109-76-2; $HOCH_2CH_2NHMe$, 109-83-1; $TsOCH_2CH_2N(Ts)Me$, 3559-06-6; $TsNHCH_2CH_2CH_2NHTs$, 53364-99-1.

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Reactions of Halopentaboranes and Other Haloboranes with Tributyltin Hydride

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Received March 26, 1985

Halopentaboranes(9) are converted to the parent pentaborane(9) in high yields with tributyltin hydride as the halogen reducing agent. Tributyltin hydride also reduces several other haloboranes and halometalloboranes. Deuterium-labeled pentaboranes(9) are produced from halopentaboranes(9) and tributyltin deuteride.

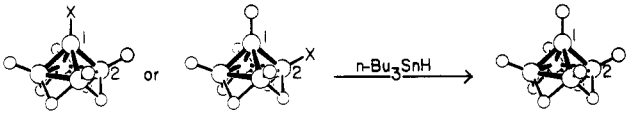
Introduction

Reactions of trialkyltin hydrides with a variety of organic functional groups have been known for nearly 40 years.[1] The first report of trialkyltin hydride reduction of boron–halogen bonds appeared in 1968.[2] In this study, chlorobis(dimethylamino)borane and 2-chloro-1,3,2-benzodioxaborole were reduced with *n*-Bu_3SnH to bis(dimethylamino)borane and 1,3,2-benzodioxaborole, respectively. Mechanistic investigations indicated a polar mechanism for these reactions. A second report of boron–halogen bond reduction by trialkyltin hydrides has recently appeared.[3] In this study, B_4Cl_4 was converted to B_4H_{10} with excess Me_3SnH.

Our interest in trialkyltin hydrides stems from a desire to replace halogens on boranes with hydrogen or deuterium atoms. We found that *n*-Bu_3SnH efficiently reduces halogens on halopentaborane(9) molecules, as well as on several other haloboranes. We report here the details of these hydrogen/halogen exchange studies along with preliminary investigations of deuterium/halogen exchange between *n*-Bu_3SnD and halopentaboranes(9) and pentaborane(9), respectively.

Results and Discussion

Halogen reduction of the several haloboranes examined by using trialkyltin hydrides or deuterides are listed in Table I. The halopentaborane(9) reductions with tributyltin hydride produce excellent yields of pentaborane(9) (eq 1), which are easily sepa-



(1)

rated from the low-volatility tin byproducts. Reaction conditions are remarkably mild when compared to the conditions frequently required for reactions of R_3SnH with alkyl or aryl halides.[1] The ease of reduction of the halopentaboranes(9) increases in the order Cl < Br < I. This trend is similar to that noted for alkyl and aryl halides.[1] A halogen position effect is also observed, in that the 2-halopentaboranes(9) produce slightly higher yields and react more rapidly than the corresponding 1-halopentaboranes(9). In view of the pseudoaromatic character associated with the chemistry of the H(1) position in pentaborane(9),[4] these observations may be construed as correlating with the organic halide analogues, in

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