

the interaction between the *l*-menthyl group and solvent molecules. In general tris(1,3-diketono)cobalt(III) and -chromium(III) are known to undergo slow racemization in solution,^{21,30} though it is recognized that cobalt(III) and chromium(III) complexes are generally inert. For example, racemization of optically active [Cr(acac)₃] in 1:1 hexane-benzene is 50% complete at 25 °C in 8 days.³⁰ So that the racemization of [Co(*l*-moba)₃] and [Cr(*l*-moba)₃] could be examined, CD spectra were measured after the solutions were left to stand for 1 month. However, no substantial change was observed for the CD spectra of these solutions. Further, [Mn(*l*-moba)₃] showed no CD spectral change after 10 days.

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All these facts suggest that the asymmetric field effected by the interligand, hydrophobic CH $\cdots\pi$ interaction preferentially stabilizes one stereoisomer (cis- Δ form) compared with other isomers.

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Registry No. Co(*l*-moba)₃, 80630-19-9; Cr(*l*-moba)₃, 80594-50-9; Mn(*l*-moba)₃, 80594-51-0; Fe(*l*-moba)₃, 80594-49-6; Ni(*l*-moba)₂(H₂O)₂, 80594-52-1; Co(*l*-moba)₂(H₂O)₂, 80594-53-2; Mn(*l*-moba)₂(H₂O)₂, 80630-20-2; Cu(*l*-moba)₂, 80594-54-3; H(*l*-moba), 80594-47-4; ethyl *l*-menthoxyacetate, 75410-31-0; K₃[Co(CO₃)₃], 15768-38-4.

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Thermodynamic Origin of the Macrocyclic Effect in Crown Ether Complexes of Na⁺, K⁺, and Ba²⁺

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As compared to the noncyclic analogue pentaglyme, 18-crown-6 complexes of Na⁺, K⁺, and Ba²⁺ have superior stabilities in methanol and methanol/water mixtures. This macrocyclic effect amounts to 3-4 orders of magnitude in the equilibrium formation constant and has been shown with use of calorimetric titrimetry to be primarily the result of more favorable enthalpic (ΔH) changes. The addition of water to the methanol solvent caused a decrease in the observed macrocyclic effects in all cases studied. These decreases were the result of less favorable enthalpic changes. As compared to K⁺, the macrocyclic effect for Na⁺ is smaller, and again the decrease was the result of unfavorable enthalpic changes. Some stability constants were also measured with use of pentaethylene glycol for comparison to the corresponding glyme.

Introduction

In earlier work describing the thermodynamics of cyclic polyether cation complexation reactions, Frensdorff¹ noted the higher stabilities of the complexes of cyclic polyethers over those of their linear glyme counterparts. He reported that the stability constants of complexes with the cyclic ligand 18-crown-6 (18C6) were several orders of magnitude larger than those with the corresponding open-chain compound pentaglyme (PG) (see Figure 1). Cabiness and Margerum² had also observed earlier that increased complex stability is found among transition-metal complexes of cyclic tetramine ligands over those of corresponding open-chain ligands. They suggested the term "macrocyclic effect" to describe this phenomenon which seemed to be a logical extension of the well-known chelate effect.

Just as some question arose over the thermodynamic origin of the chelate effect,³ so different conclusions have been drawn concerning the macrocyclic effect; i.e., some investigators have ascribed the effect to enthalpic (ΔH) factors while others have implicated entropic (ΔS) factors as the most important in contributing to the enhanced stabilities (ΔG) of complexes containing macrocyclic ligands. Margerum and co-workers^{2,4,5} found the enthalpy term to predominate in the observed macrocyclic effect when the formation of Ni(II) complexes of cyclam and those of 2,3,2-tet (Figure 1) are compared. Dei and Gori⁶ reached a similar conclusion using the Cu(II) complexes of these two ligands. However, both Paoletti and co-workers⁷ and Kodama and Kimura⁸ found that the entropy

term was primarily responsible for the macrocyclic effect using Cu(II) complexes of cyclen and 2,2,2-tet. Paoletti and co-workers later concluded that the relative magnitude of the enthalpy contribution is critically dependent on the match between cation and ligand cavity sizes for transition metals.⁹ In a study of Cu²⁺ complexes of ligands containing S, N, and O donor atoms, Arnaud-Neu et al. concluded that the macrocyclic effect has both enthalpy and entropy contributions.¹⁰

One of the objectives of this study was to determine whether the macrocyclic effect found in polyether complexes of Na⁺, K⁺, and Ba²⁺ was the result of either enthalpic or entropic factors or a combination of both factors. This was accomplished by comparing log *K*, ΔH , and $T\Delta S$ values for the reaction of these metal cations with cyclic 18C6 with corre-

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[†] Contribution No. 184.

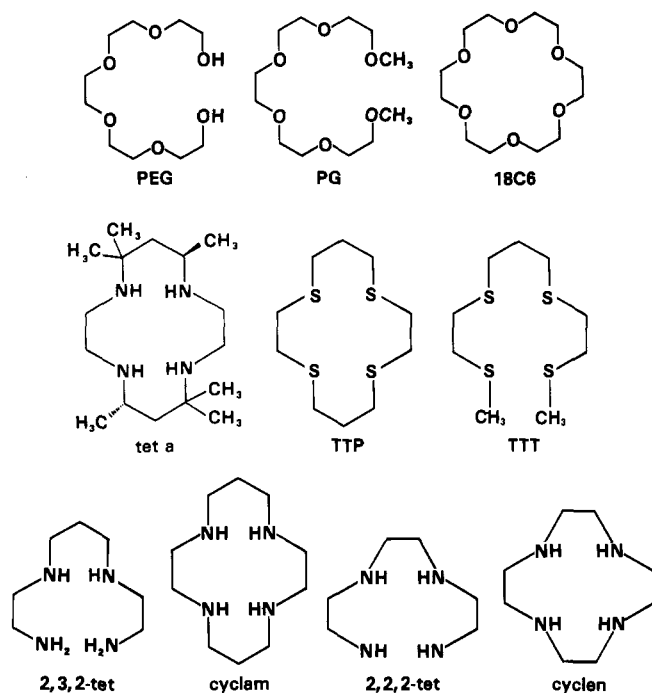


Figure 1. Structures of various cyclic and analogous noncyclic ligands.

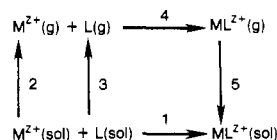


Figure 2. Born-Haber cycle for reaction of a ligand L with a cation M^{2+} [$\Delta X_1 = \Delta X_2 + \Delta X_3 + \Delta X_4 + \Delta X_5$ ($X = G, H, \text{ or } S$)].

sponding values for PG complexes and pentaethylene glycol (PEG) complexes. It was hoped that such experiments would help in understanding the nature of the polyether-metal cation interaction and give some information about the macrocyclic effect in general.

There are certain advantages in studying the complexation of polyether ligands with Na⁺, K⁺, and Ba²⁺ as opposed to the complexation of polyamine ligands by first-row transition metals: (i) the polyether ligands are weak, uncharged bases at neutral pH and formation of their complexes is not pH dependent; (ii) among the metal ions which form complexes with polyethers are alkali and alkaline earth cations which can be considered simple, solvated, charged spheres, unlike transition-metal ions which have specific stereochemical preferences; (iii) the rates of reaction between crown ethers and cations are usually quite rapid so that equilibrium measurements are readily obtained; (iv) stability constants are in the range (10^2 – 10^8 in methanol) that can be conveniently measured directly.

Another objective of our work was to come to some understanding of the origin of the macrocyclic effect in polyether complexes from a microscopic point of view. Hinz and Margerum⁴ proposed that ligand solvation enthalpy was primarily responsible for the macrocyclic effect in the system which they studied. However, this conclusion has been questioned by Paoletti and co-workers⁷ who, while ignoring lattice energies, showed that the dehydration enthalpies of 2,3,2-tet-4H₂O and tet a-2H₂O were nearly the same (26 kcal/mol). The ligand solvation hypothesis can be understood in terms of the inner-sphere ligand desolvation step (step 3) in the Born-Haber cycle shown in Figure 2. The reasoning was that the release of solvent molecules from the metal ion (step 2) and the ligand (step 3) results in positive ΔH and ΔS changes. If the desolvation of the metal is the same in the

reactions with both cyclic and noncyclic ligands, then the desolvation of the ligand is the important step when the reactions of Ni(II) in water with different ligands are compared. It was assumed that the cyclic ligand, being much more compact, is solvated by fewer solvent molecules and that the ligand desolvation enthalpies would be significantly smaller, thus resulting in greater stability for the complex containing the cyclic ligand. Hinz and Margerum also suggested that unfavorable configurational entropy changes for the noncyclic ligand played a smaller but important role in the macrocyclic effect in the Ni(II)/cyclam/2,3,2-tet system.⁴ Furthermore, Smith and Margerum also indicated that these configurational entropy effects were the primary cause for the smaller macrocyclic effect observed in the Ni(II)/TTP/TTT system which was studied in nitromethane.⁵ Here the differences in the thio ether-solvent interactions were presumed to be small enough to be ignored. Unfortunately, only ΔG ($\log K$) values are reported, and the conjectures of the authors are not easily substantiated. Although Margerum and co-workers dismiss the importance of configurational enthalpy changes of the ligands, other workers feel that they are important factors which contribute to the size of the macrocyclic effect.⁷⁻⁹

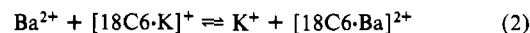
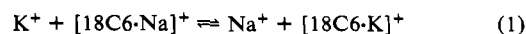
We now report comparative thermodynamic data for the reactions of Na⁺, K⁺, and Ba²⁺ with 18-crown-6, pentaglyme, and pentaethylene glycol in 100%, 99%, and 90% methanol.

Experimental Section

Materials. Reagent grade 18C6 (Parish Chemical) was used without further purification. Solutions of this ligand were standardized by calorimetric titration against standard KCl solutions in absolute methanol.¹¹ It was found from this standardization technique that the ligand was $99 \pm 1\%$ pure. The ligand PEG (Columbia Organic Chemicals) was carefully dried and vacuum distilled; PG was prepared by the method below. The following reagent grade salts were obtained: KCl and NaCl (Baker); Ba(ClO₄)₂ (Fisher). The salts NaCl, KCl, and Ba(ClO₄)₂ were dried at 120 °C for 24 h and stored in a desiccator over P₂O₅. The concentrations of the NaCl and KCl solutions were determined gravimetrically. Solutions of Ba(ClO₄)₂ were standardized with use of calorimetric titrations against 18C6 solutions which in turn had been standardized against KCl. The absolute methanol was Fisher reagent grade and was found by Karl Fischer titration to contain less than 0.04 wt % water.

Pentaglyme. Forty-six grams of small, freshly cut sodium chunks was dissolved in 70-mL dry 2-methoxyethanol. After the reaction ceased and the mixture was cooled to room temperature, 187 g of (ClCH₂CH₂OCH₂)₂ was added slowly with vigorous stirring over a period of 2 h. The mixture was stirred for an additional 2 h and then refluxed for 8 h. The yellow reaction mixture was filtered to remove the white solid, and the excess methoxyethanol was removed with use of a rotary evaporator. The residual liquid was vacuum distilled (1.0 torr), and two distinct fractions were obtained at 85 and 140 °C. The higher boiling fraction was carefully redistilled to give 85 g (32%) of the pure colorless liquid which boiled at 139–140 °C (1.0 torr). Proton NMR spectra at 60 MHz (CCl₄, 30 °C, TMS reference) showed two resonances, δ 3.18 (6 H, s) and 3.42 (20 H, br m). The mass spectrum showed a parent ion corresponding to a mass of 266. Anal. Calcd for C₁₂H₂₆O₆ C, 54.12; H, 9.84. Found: C, 53.86; H, 9.90.

Procedure. $\log K$, ΔH , and $T\Delta S$ values were determined as described earlier¹² by titration calorimetry using a Tronac Model 450 isoperibol calorimeter equipped with a 25-mL reaction vessel. $\log K$ values for reaction of 18C6 with Ba²⁺ and K⁺ were found to be greater than 5.5 (see eq 1 and 2). These stability constants were



accurately determined by a competitive technique described previ-

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Table I. $\log K$, ΔH , and $T\Delta S$ Values for Reactions of Na^+ , K^+ , and Ba^{2+} with Various Ligands in Water/Methanol Mixtures at 25 °C

cation	solvent ^a	ligand	$\log K$	ΔH^b	$T\Delta S^b$
Na^+	90	18C6	3.66 ± 0.02	-6.64 ± 0.07	-1.65
		99	4.33 ± 0.02	-8.11 ± 0.05	-2.20
	100	18C6	4.36 ± 0.02	-8.36 ± 0.37	-2.41
		PG	1.44 ± 0.05	-4.02 ± 0.28	-2.06
			1.0^c	-9.13^c	-7.8^c
K^+	90	18C6	5.35 ± 0.25	-11.77 ± 0.05	-4.47
		PG	1.95 ± 0.01	-7.00 ± 0.08	-4.34
		PEG	1.91 ± 0.01	-4.57 ± 0.02	-1.96
	99	18C6	6.05 ± 0.05	-13.21 ± 0.07	-4.96
		PG	2.27 ± 0.02	-8.15 ± 0.12	-5.05
		PEG	2.05 ± 0.03	-6.36 ± 0.10	-3.56
	100	18C6	6.06 ± 0.03	-13.41 ± 0.06	-5.14
		PG	2.1^c	-8.70^c	-5.8^c
		18C6	6.56 ± 0.09	-10.33 ± 0.11	-1.38
		PG	2.33 ± 0.03	-7.10 ± 0.16	-3.92
Ba^{2+}	90	18C6	6.56 ± 0.09	-10.33 ± 0.11	-1.38
		PG	2.33 ± 0.03	-7.10 ± 0.16	-3.92
		PEG	3.45 ± 0.01	-7.60 ± 0.01	-2.89
	99	18C6	7.03 ± 0.06	-10.38 ± 0.15	-0.79
		PG	2.51 ± 0.01	-5.65 ± 0.05	-2.23
		PEG	3.96 ± 0.08	-6.71 ± 0.12	-1.31
100	18C6	7.04 ± 0.08	-10.41 ± 0.06	-0.80	
	PG	2.3^c	-5.55^c	-2.4^c	

^a Given as wt % methanol in methanol/water mixture. ^b kcal/mol. ^c From Früh and Simon.¹⁴ ^d From Frensdorff.¹

ously.¹³ For reactions involving PG and PEG with the cations studied, a computer program allowed variation of the ligand concentration to give the best data fit. By this means, the purities of PG and PEG were found to be $98 \pm 2\%$ and $97 \pm 2\%$, respectively. Owing to the discrepancy between our data for Na^+ /PG in methanol and that of Früh and Simon,¹⁴ this Na^+ /PG reaction was also carried out at higher PG concentration (0.3 M) with use of freshly distilled PG whose purity was fixed and determined to be 99.1% by gas chromatography.

Results

$\log K$, ΔH , and $T\Delta S$ values for the reactions of Na^+ , K^+ , and/or Ba^{2+} with 18C6, PG, and PEG in 90 wt % methanol and 99 wt % methanol are given in Table I. In addition, some values are given for reactions in 100 wt % methanol as well as those values of other workers.^{1,14} At first we were not sure we could obtain reliable experimental data in absolute methanol (>99.9 wt %), so 99 wt % methanol was used; later experiments showed that we could indeed obtain reliable data in absolute methanol. Examination of the results for the same reaction in different solvents shows significant differences between results in 90% and 99% methanol and very little difference (see first three rows in Table I) between thermodynamic values for reactions in 99% and 100% methanol. For this reason, not all reactions were rerun in absolute methanol. We earlier noted good agreement between our $\log K$ results for the reactions of Na^+ and K^+ with 18C6 in methanol¹⁵ and those of Frensdorff.¹ Good agreement is also found between our values and those of Früh and Simon¹⁴ with one exception, the reaction of PG with Na^+ . Although our $\log K$ value agrees well with that of Frensdorff, we are unable to account for the discrepancy between our $\log K$ and ΔH values and those of Früh and Simon.

In order to obtain an idea of the relative magnitude of the ligand solvation energies, we titrated pure ligands into pure solvent. The values, given in Table II, were essentially constant over a wide dilution range, included in which were the dilution ranges used in the calorimetric complexation studies.

Table II. Heat Changes upon Titration of Several Liquid Ligands into CH_3OH and 90 wt % $\text{CH}_3\text{OH}/\text{H}_2\text{O}$ Solvents at 25 °C

ligand	ΔH , kcal/mol ^a	
	100% MeOH	90% MeOH
18C6	-0.22	-2.7
PG	0.37	-1.1
15C5	0.46	-1.0
glycols ^b	-0.11	-1.8

^a Uncertainties $\sim 2\%$. ^b 97% pentethylene glycol, 2% hexaethylene glycol, 1% tetraethylene glycol.

Discussion

Thermodynamic Origin of the Macrocyclic Effect. The data in Table I indicate that the macrocyclic effect in the five systems studied is primarily the result of favorable enthalpic factors, not entropic factors. The $\text{Na}/18\text{C6}/\text{PG}$ (100% MeOH), $\text{K}/18\text{C6}/\text{PG}$ (99% MeOH), and $\text{Ba}/18\text{C6}/\text{PEG}$ (99% MeOH) comparisons all show negligible entropy contributions ($\Delta(T\Delta S) \leq 0.5$ kcal/mol) to the macrocyclic effect. In the $\text{K}/18\text{C6}/\text{PEG}$ (99% MeOH) system, the data show that entropy changes actually work *against* ($\Delta(T\Delta S) = -1.4$ kcal/mol) complexation by the cyclic polyether; yet, the highly favorable enthalpy of complexation of K^+ by 18C6 overcomes these entropy changes and is large enough to cause a sizable macrocyclic effect. Even in the one system, $\text{Ba}/18\text{C6}/\text{PG}$ (99% MeOH), where entropy changes ($\Delta(T\Delta S)$) do contribute to the macrocyclic effect, 77% of this effect is the result of enthalpic factors. Thus, the expected entropy contribution to the macrocyclic effect resulting from less unfavorable configurational entropy changes in the cyclic ligand was not in fact observed, and if configurational entropy changes do play a role at all, they are masked by other thermodynamic factors.

It is worthy of note that the macrocyclic effect was observed for all systems studied and varied from 2.9 to 4.5 orders of magnitude in the equilibrium constant. Why is the macrocyclic effect for the $\text{K}/18\text{C6}/\text{PG}$ system 1 order of magnitude larger than that for the $\text{Na}/18\text{C6}/\text{PG}$ system? Structural data clearly show that of the alkali metal cations, K^+ best fits into the cavity of 18C6 although even K^+ is slightly too big.¹⁶ However, Na^+ is much smaller than the 18C6 cavity. In this latter case, the 18-membered ring must either (1) adopt an elliptical shape where four oxygen atoms are near and two are farther from the Na^+ ion or (2) adopt an envelope shape where one oxygen atom radically departs from the plane of the other five which collapse to form a smaller ring of oxygen atoms closer to the Na^+ or (3) remain in the pseudo- D_{3d} geometry with quite long $\text{O}\cdots\text{Na}^+$ interactions.¹⁷ On the basis of numbers and strengths of $\text{Na}^+\cdots\text{O}$ interactions (as judged from $\text{Na}^+\cdots\text{O}$ distances), the fit of the ion inside the cyclic polyether cavity is important; the sizes must match up for maximum ligand-metal bonding. This is another way of saying that conformational enthalpy plays an important role in cyclic polyether-metal ion binding. Transition metals impose even stricter requirements than Na^+ or K^+ on a cyclic ligand; not only do metal-donor atom distances have to be optimal for maximum bonding, but the donor atoms must be properly disposed in a geometric sense to assure optimal bonding between the metal d orbitals and lone pairs of electrons on the donor atoms. Indeed, one can envision the case where the conformational energy required to place the donor atoms of a cyclic ligand in their proper bonding locations is so great that the stability of the resulting guest-host complex is low enough to remove the macrocyclic effect or even cause an

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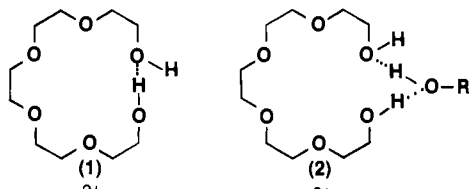
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inverse macrocyclic effect. This occurs in the Cu(II)/triza-9-crown-3/diethylenetriamine system in water where the stability constant for the cyclic ligand ($\log K = 15.5$ or 16.2)¹⁸ is about the same as that for the linear triamine ($\log K = 16.0$).¹⁹

Another example is the C₆H₅N₂⁺/15C5/tetraglyme system in acetone; tetraglyme binds to C₆H₅N₂⁺ at least 1 order of magnitude more strongly than does 15C5.¹⁶ Models show that the 15C5 cavity size is too small to accommodate the C₆H₅N₂⁺ ion and that unfavorable nonbonded H...H interactions are also present. On the other hand 18-crown-6 interacts more strongly with C₆H₅N₂⁺, and a significant macrocyclic effect is observed which is even larger for 21C7 which possesses the optimum size and steric requirements to bind diazonium salts.²⁰ Returning to K⁺/18C6/PG and Na⁺/18C6/PG, we observe that the enhanced macrocyclic effect in the potassium system is strictly the result of enthalpic factors which is in accord with the notion that unfavorable conformational energetics in the Na⁺/18C6 complex are responsible for the difference. We note that such comparisons are only meaningful if the nature of the final complexes with cyclic and noncyclic ligands are similar; that is, the number of solvent molecules in the primary coordination sphere (bonded to the metal) is the same in the two complexes and that all donor atoms of the noncyclic ligand are bonded to the metal in a pseudo cyclic fashion. Unfortunately we do not know the geometry or coordination number of pentaglyme cation complexes in alcoholic media and can only surmise that the primary coordination environments of 18C6 and PG complexes are similar.

A comparison of the K⁺/18C6/PG and Ba²⁺/18C6/PG systems shows that the macrocyclic effect is greater in each solvent for Ba²⁺ by approximately 0.8 log *K* units, but the enthalpy contribution to this effect is decreased in the case of Ba²⁺. It seems that the more strongly solvated bivalent cation (Ba²⁺ is about the same size as K⁺) is more difficult to desolvate which decreases the relative exothermicity of complexation. However, the increased number of liberated solvent molecules increases the entropy of complexation which more than offsets the less favorable enthalpy.

A comparison of PG complexation with that of PEG shows some interesting but not easily understandable trends. Toward Ba²⁺, PEG is a much stronger ligand than PG, but the reverse is true for K⁺. The choice of a reference noncyclic ligand is clearly important when one tries to measure the macrocyclic effect; the macrocyclic effect for Ba/18C6 is 4.5 log *K* units (99% CH₃OH) if PG is chosen as the reference ligand but only 3.1 log *K* units (99% CH₃OH) if PEG is chosen. The unique behavior of PEG most likely is related to (1) its ability to replace more solvent molecules from the primary coordination sphere of the metal upon complexation and (2) its ability to form hydrogen bonds with itself (1) or bound solvent molecules



(2). Further speculation is unwarranted until the structures of the PEG complexes can be ascertained.

Some results concerning the macrocyclic effect in polyether ligands have been reported by others. The thermodynamic properties of complexes of PG and dibenzo-18-crown-6 have

been compared.¹⁴ However, the presence of the benzene ring makes the validity of such comparisons questionable because such substituents are known to have large effects on log *K* and ΔH of complexation. In addition, data comparing the stability of Pb(18C6)²⁺ with Pb(tetraglyme)²⁺ and Pb(tetraglyme)₂²⁺ are available;²¹ again the comparison suffers greatly because tetraglyme has only five donor oxygen atoms, and the complexes are not strictly comparable.

Ligand Solvation and the Macrocyclic Effect. The heats of solution of 18C6, PG, 15C5, and a glycol mixture (predominately PEG) are listed in Table II. These values represent the difference in energies of self-solvation vs. solvation by the respective solvent. Although expected trends were observed, the changes between 90% and 100% CH₃OH of the numbers are small in all cases, so the ligand solvation term must be small in the thermodynamic cycle (Figure 2). Unfortunately, the heats of vaporization are not known for these ligands so the effects of these terms in the thermodynamic cycle cannot be evaluated. However, if PG or 18C6 did form inner-sphere solvates of the sort >O...H-OR, we would have expected much larger enthalpies of solution. PG and 18C6 cannot engage in self H bonding, although PEG could do so. These specific types of interaction are known to be weaker than >N...H-OR bonding but still important (ca. 4-5 kcal/mol). For example the MeO...H-OMe H-bonding energy in the gas phase is 4.1 kcal/mol,²² and *p*-FC₆H₄O-H...OEt₂ H-bonding energy is 5.6 kcal/mol in CCl₄.²³ Thus, we cannot find evidence for exothermic inner sphere H bonding (ligand solvation) between methanol and the polyethers, either cyclic or noncyclic. The enthalpy of solution of PG is not only about the same as that of cyclic 18C6 but also about the same as that of PEG which has internal hydroxyl groups.

The reactions of PG, PEG, and 18C6 with K⁺ and Ba²⁺ were also studied in 90 wt % methanol in order to learn the effect of the solvent change on the macrocyclic effect. These two cations were chosen because they are almost identical in size and form the strongest complexes with 18C6 of all cations in their respective groups in the periodic table.^{15,17} In 90% MeOH, the macrocyclic effect decreased by 0.3 log *K* units for both the K⁺/18C6/PG and the Ba²⁺/18C6/PG systems. In both cases the decrease in the macrocyclic effect was primarily the result of enthalpy changes. Similarly, for the K⁺/18C6/PEG system, the macrocyclic effect decreased by 0.5 log *K* units. However, for the unique Ba²⁺/18C6/PEG system, there was no change in the macrocyclic effect; both ligands were similarly affected by the solvent change. It was our intuitive notion that the addition of 10 wt % of H₂O to methanol would improve the ability of the solvent to form H bonding to the polyethers (if such bonding were present) because of the much smaller size of H₂O as compared to MeOH. If this were true and if the noncyclic polyether were more strongly solvated, then the macrocyclic effect should increase in the better solvating medium (90% MeOH). However, we observe the reverse behavior. Either (1) our assumptions about the improved solvating ability of 90% MeOH are incorrect or (2) the cyclic polyether is more strongly solvated than the linear one or (3) other solvent characteristics are far more important than specific solvent-polyether interactions. With regard to this last possibility we should mention that the dielectric constant of the solvent changes from 32.6 (100% MeOH) to 64.3 (90/10 wt % MeOH-H₂O).²⁴ Thus the well-known leveling effect of the more polar solvent in sec-

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ondary solvation spheres may reduce all equilibrium constants with the larger constants being affected more; one would expect an apparent decrease in the observed macrocyclic effect in this case when water is added to the methanol solvent. Although the change in solvents does not conclusively point to a specific microscopic cause, the changes in the observed macrocyclic effect are real and significant, and these experiments do show that the specific inner sphere solvation of the noncyclic polyether is *not* the dominant factor which causes the changes in the macrocyclic effect when water is added to the methanol solvent.

The small heats of dilution of the free ligands taken together with the decrease in the observed macrocyclic effect when water is added to methanol has led us to search for an explanation for the macrocyclic effect other than increased solvation of the noncyclic ligand. The macrocyclic effect for the systems we studied was caused by less favorable enthalpic factors for the noncyclic ligand. Changing from K^+ to Na^+ or changing from 100% MeOH to 90% MeOH both caused a decrease in the macrocyclic effect, and these decreases were *enthalpic* in origin. It seems to us that unfavorable conformational enthalpic factors associated with the coordination of the pentaglyme can explain most of the 4–6 kcal/mol in ΔH required to account for the macrocyclic effect in these polyether ligands.

In three systems reported by Margerum and co-workers, macrocyclic effects were measured to be about 10^4 (Cu^{2+} /tetramine in H_2O), 10^2 (Ni^{2+} /tetrathioether in CH_3NO_2), and 10^6 (Ni^{2+} /tetraamine in H_2O).^{2,4,5} For the last system, the ΔH and ΔS contributions to ΔG were determined, and it was found that the enthalpy term was the one which contributed to the magnitude of the macrocyclic effect. As noted earlier, they rationalized that the conformational enthalpy associated with the noncyclic ligand was much too small to contribute to the 14 kcal/mol difference which they observed between the binding of the cyclic and noncyclic ligands to Ni^{2+} . Hence, they attributed most of this 14 kcal/mol difference to the additional solvation (two inner-sphere H_2O molecules) of the

noncyclic tetraamine as compared to the cyclic tetraamine. However, we, like Paoletti and co-workers,^{7,9} emphasize the importance of conformational enthalpies of the ligands. During the synthesis of cyclic ligands, a certain amount of conformational strain is built into the macrocycle. The positions of the ligating donor atoms in the macrocycle are such that these atoms find themselves in either *more* or *less* favorable orientations for binding to a particular metal ion. One does not know a priori if the cyclization process will increase or decrease the enthalpies of ligand–metal binding.

Conclusion. Our experimental results show that the macrocyclic effect is operative in five linear/cyclic polyether combinations when methanol is used as a solvent and that this effect is the result of more favorable *enthalpy* factors. We have also shown that the magnitude of this effect is dependent on the “reference” linear polyether used and on how well the cation size and ligand cavity size are matched. We know that changing the solvent from methanol to methanol/water (90/10) and changing from K^+ to Na^+ causes a decrease in the macrocyclic effect which in each case is enthalpic in origin. Like other workers,^{2,4,5,8} our initial belief that the macrocyclic effect was caused, at least in part, by unfavorable conformational entropy changes in the linear ligand upon complexation must remain unsubstantiated in the case of polyethers in methanol. Our results do not conclusively point to any single microscopic source for the macrocyclic effect in polyethers, but they do indicate that unfavorable conformational enthalpy changes of the linear polyethers are important factors. Given the wide range of solvents, ligand sizes, numbers and types of donor atoms, metals, etc., which have been studied, we believe that a combination of factors at the molecular level is responsible for the presence or absence of the macrocyclic effect; in a specific case, one of these factors may predominate.

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Molybdenum(IV)–Oxo Complexes with Oxygen, Nitrogen, and Sulfur Ligands. Syntheses and Electrochemical Studies

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New Mo(IV)–oxo complexes with tri- and tetradentate oxygen, nitrogen, and sulfur ligands have been synthesized ($MoOLL'$, L = salicylaldehyde *o*-hydroxyanil, salicylaldehyde *o*-mercaptoanil, $L' = \alpha, \alpha$ -bipyridyl, *o*-phenanthroline, dmf; $MoOL$, $L = N, N'$ -bis(2-mercapto-2-methylpropyl)ethylenediamine) by phosphine oxo abstraction from Mo(VI)–dioxo complexes, ligand displacement from known Mo(IV)–oxo complexes, and reaction with $MoCl_4$ in methanol. Electrochemical properties have been determined by cyclic voltammetry and controlled-potential coulometry, and electronic and IR spectra were obtained. The complexes are oxidized to Mo(V)–oxo complexes and undergo one-electron reductions at a platinum electrode in DMF. The limitations of the synthetic methods and the implications of the electrochemical results for molybdenum enzymes are discussed.

Introduction

The molybdenum center of the molybdenum oxidases and nitrate reductase appears to cycle between the 6+, 5+, and 4+ states during catalysis of electron transfer between substrates and electron donors or acceptors.^{1,2} The most recent

EXAFS data for reduced xanthine dehydrogenase and reduced sulfite oxidase indicates the presence of a single terminal oxo ligand, two to three thiolate ligands, and possibly a thioether, oxygen, or nitrogen ligand in the Mo(IV) coordination sphere.³

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