Thermodynamics of Complexation of a Lower Rim Calix(4)arene Ketone Derivative and the Sodium Cation in N, N-Dimethylformamide/Acetonitrile Mixtures.

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Stability constants and derived standard Gibbs energies, enthalpies, and entropies of complexation of a lower rim calix(4)arene ketone derivative, namely 5,11,17,23-tetra *p-tert*-butyl-25,26,27,28-tetra(benzoyl)methoxycalix-(4)arene and the sodium cation in N, N-dimethylformamide/acetonitrile mixtures at 298.15 K are first reported. Stability constants were determined potentiometrically by the use of a sodium selective electrode with the silver/silver (I) reference electrode. A linear relationship is found between the stability constant and the solvent composition. The strength of complexation decreases from acetonitrile through the mixtures. The minimum stability for this system was found in N, N-dimethylformamide. The striking feature of these data are the small variations in enthalpy observed in the solvent mixtures relative to their pure solvents. In all cases, the complexation process in these solvent mixtures is enthalpically controlled. However, the higher stability found by increasing the content of acetonitrile in the mixture is accompanied by an increase of entropy. Standard enthalpies of solution of the reactants (NaClO$_4$ and ligand) and the product (metal-ion complex) are reported. These data are used to calculate the standard coordination enthalpies as a means of checking the reliability of the data. Transfer enthalpy data from acetonitrile are calculated in an attempt to establish the factors which contribute to the enthalpies of complexation in the mixed solvents relative to corresponding data in the pure solvents. The exothermic character of the transfer process involving the sodium complex electrolyte increases as the content of DMF in the mixture increases but $\Delta_t H°$ from acetonitrile to neat DMF became less negative. The ^{1}H NMR study carried out in CD$_3$CN, d$_7$ − DMF and mixed solvents shows that the interaction between these two solvents enhances the interaction of the molecule of acetonitrile through its methyl group with the hydrophobic cavity of the complex. This explains the gain in the enthalpic stability observed in the transfer of the sodium complex salt from acetonitrile to the mixtures.

Introduction

Thermodynamic studies of macrocyclic complexation reactions with metal cations not only provide relevant information on the binding process but also lead to a better understanding of the selective behavior of these ligands toward different metal cations. Although cation complexation studies involving macrocycles have been investigated in a wide variety of solvents,[1,2] very limited information is available in mixed solvents.[2−4] Most of these investigations have been carried out in water−methanol mixtures. As previously discussed,[5] the differences observed in the thermodynamics of complexation, $\Delta_c P^0$, in a given solvent, s$_1$ (reference solvent) relative to another (s$_2$) depend on the solvation changes, $\Delta_t P^0$ that the reactants (metal cation, M$^+$ and the ligand, L) and the product (metal-ion complex, M$^+$L) undergo in the two solvents. This is explicitly shown in eq 1 where $P^0 = G°, H°, S^0$

$$\Delta_c P^0 (s_2) - \Delta_c P^0 (s_1) = \Delta_t P^0 (M^+L) (s_1 \rightarrow s_2) -$$
$$\Delta_t P^0 (M^+) (s_1 \rightarrow s_2) - \Delta_t P^0 (L) (s_1 \rightarrow s_2) \quad (1)$$

We have extensively used this equation[5,6] to quantitatively assess

the medium effect on the complexation process involving macrocycles and metal cations in neat solvents.

There is information available in the literature concerning complexation reactions involving crown ethers and alkali-metal cations in N, N-dimethylformamide (DMF)−acetonitrile (MeCN) mixtures.[7,8] These studies have shown that a decrease of the amount of DMF in the mixture results in a decrease of the standard Gibbs energy and the enthalpy of complexation but an increase in entropy was observed. However, we are not aware of any thermodynamic studies of complexation involving calix-(4)arene derivatives and metal cations in mixed solvents. Unlike crown ethers,[9] characterized by the presence of a single hole, lower rim calix(4)arene derivatives[10−13] possess two cavities, one of which is situated between the aromatic rings, and therefore, this cavity is hydrophobic and the other is hydrophilic. Whereas the former is able to interact with neutral species, the latter is known to complex metal cations. The thermodynamics of calixarene chemistry reviewed by Danil de Namor and co-workers[14] in 1998 reflect that there is a lot to be done in this area and a great deal of data reported need to be revisited.
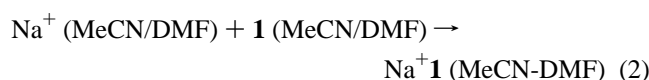
This paper reports the following:

(i) The thermodynamics of complexation of a *p-tert*-butylcalix-(4)arene derivative, namely, 5,11,17,23-tetra-*tert*-butyl-25,26,-27,28-tetra(benzoyl)methoxycalix(4)arene **1** and the sodium cation in acetonitrile/*N,N*-dimethylformamide (MeCN/DMF)

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Thermodynamics of Complexation

J. Phys. Chem. B, Vol. 105, No. 31, 2001 **7543**

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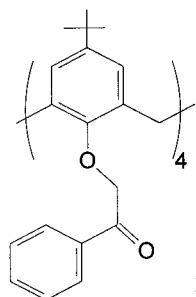
mixtures at 298.15 K. This process is represented by eq 2

$$Na^+ (MeCN/DMF) + \mathbf{1} (MeCN/DMF) \rightarrow$$
$$Na^+\mathbf{1} (MeCN-DMF) \quad (2)$$

For the formulation of eq 2, the composition of the complex and the speciations in solution are established from conductance measurements.

(ii) Standard enthalpies of solution of the reactants (NaClO₄ and ligand) and the product (metal-ion complex salt) at 298.15 K in three different composition mixtures were measured calorimetrically. The data were used to determine the standard coordination enthalpies, $\Delta_{coord}H°$ referred to the process where both the reactants and the product are in the solid (sol) state (eq 3)

$$NaClO_4 (sol) + \mathbf{1} (sol) \rightarrow Na\mathbf{1}ClO_4 (sol) \quad (3)$$
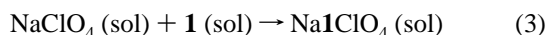
(iii) Transfer enthalpies from acetonitrile to the various mixtures are used to assess through eq 1, the contribution of the reactants and the product to the variations in complexation enthalpies resulting from the medium effect.

Experimental Section

Chemicals. 5,11,17,23-tetra-*tert*-butyl-25,26,27,28-tetra(benzoyl)methoxycalix(4)arene **1**, was synthesized at the Thermochemistry Laboratory, UniS, as described elsewhere.[15] The microanalysis was carried out at the University of Surrey. The found percentages of C (81.40) and H (7.17) are in agreement with the calculated values (C, 81.40; H, 7.19). Sodium perchlorate (99%) and silver perchlorate, both from Aldrich were dried under vacuum for several days before use. The former salt at 70 °C whereas for the latter P₄O₁₀ was used.

Tetra-*n*-butylammonium perchlorate (TBAP) (Fluka) was used without further purification.

Acetonitrile (HPLC grade, Fisher) and *N,N*-dimethylformamide (Fisher, AR) were purified as described elsewhere.[16]

The MeCN/DMF mixtures were made up by weighing the amounts of solvents that corresponded to the desired mole fractions.

The sodium complex of *p-tert*-butylcalix(4)arene tetraphenyl ketone **1** was prepared by extraction of NaClO₄ from its aqueous solution by a solution of the ligand in dichloromethane. The phases were separated, the organic layer dried with anhydrous sodium sulfate and the solvent evaporated. The white solid powder was dried at 70 °C at reduced pressure for several days before use. The microanalysis was carried out at the University of Surrey. The found percentages of C (72.98) and H (6.44) are in agreement with the calculated values for C (73.38) and H (6.48). ¹H NMR, *(*CDCl₃) (300 MHz), δ (ppm): 1.19 (s, 36H,), 3.47 (d, 4H, *J* = 12 Hz), 4.41 (d, 4H, *J* = 12 Hz), 5.35 (s, 8H), 7.20 (s, 8H), 7.43 (t, 8H), 7.62 (t, 4H,), 7.74 (d, 8H, *J* = 7 Hz).

Conductance Measurements. *Conductometric Titration of NaClO₄ with p-tert-Butylcalix(4)arene Tetraphenyl Ketone.* For these measurements, the Wayne-Kerr model 7330 Automatic LCR Meter Conductivity Bridge at a frequency of 1 kHz was used. The conductance cell was a Russel type glass bodied electrode with a cell constant (determined using 0.10 mol dm⁻³ aqueous KCl solution) of 1.081 ± 0.001 cm⁻¹. For these experiments, the vessel was filled with the sodium perchlorate solution in the appropriate solvent mixture (20 cm³), and the conductance of the solution was measured. Then, a known volume of solution of **1** in the same mixture was added stepwise into the vessel and the conductance measured after each addition. The temperature of the vessel was kept at 298.15 K using a thermostat.

Conductance Measurements of the Metal-Ion Complex Salts. For these experiments, the vessel was filled with the appropriate solvent mixture (20 cm³) and the conductance of the mixture was measured. The metal−ion complex salt solution in the same solvent mixture (∼1 × 10⁻³ mol dm⁻³) was added in steps into the vessel. After each addition, conductance readings at 298.15 K were recorded.

Potentiometric Titrations. The stability constants of **1** and the sodium cation were measured by potentiometry with a digital microprocessor pH/mV meter HANNA model H18417 using a sodium selective glass electrode (Russel) as an indicator electrode and a silver wire introduced in a solution of silver cations as the reference electrode. The working and reference vessels were connected with a salt bridge containing the supporting electrolyte (TBAP) and kept at a constant temperature (298.15 ± 0.05 K) using a thermostat bath. For the calibration experiment the vessel with the sodium selective electrode was filled with a solution of TBAP (10 cm³) in the appropriate solvent mixture and that containing the reference electrode was filled with AgClO₄ solution (10 cm³). The apparatus with the solutions was left to equilibrate and after about 30 min the solution of sodium perchlorate (1.5 × 10⁻³ mol dm³) was added stepwise (∼12 additions) into the indicator cell. Potential readings were taken after each addition. The Nernstian behavior of the electrode was evaluated from the slope of a plot of potential against −log [Na⁺].

For the determination of the stability constant, the ligand solutions (1.6 × 10⁻⁴ to 6.5 × 10⁻⁴ mol dm⁻³) were prepared in a TBAP solution (0.05 mol dm⁻³, 10 cm³) in the appropriate solvent mixture. In a typical experiment, the indicator electrode placed into a solution containing a known concentration of Na⁺ was titrated with a solution of **1** allowing the formation of the 1:1 sodium calixarenate complex. Potential readings were taken after each addition and the data were used for the calculation of the stability constant of the sodium calixarenate complex in MeCN/DMF mixtures. The experiments were repeated twice for each mixture.

Calorimetric Titrations. Enthalpies of complexation in the mixed solvents were determined by classical calorimetry using a Tronac 450 calorimeter. The vessel was filled with the ligand solution (50 cm³) in the appropriate mixture and a solution of sodium perchlorate in the same medium was added from the syringe. These experiments were repeated three times for each mixture. Blank experiments were carried out in all cases to account for heat of dilution effects resulting from the addition of NaClO₄ to the solvent contained in the calorimetric vessel. The enthalpy of protonation of THAM (Tris[hydroxy methyl]-aminomethane) in an aqueous solution of hydrochloric acid (0.1 mol dm⁻³), suggested as a standard thermochemical reaction,[17] was used to check the accuracy and reproducibility of the
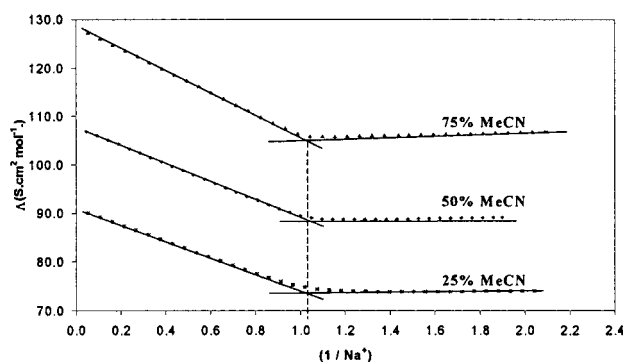
Figure 1. Conductometric curves for the titration of **1** with sodium (perchlorate as counter ion) in different acetonitrile/N,N-dimethylformamide mixtures at 298.15 K.

equipment. The value obtained (-47.57 kJ mol^{-1}) is in good agreement with that reported by Wilson and Smith[18] (-47.49 kJ mol^{-1}) at 298.15 K.

Solution Calorimetry. Enthalpies of solution were measured with the Tronac 450 calorimeter. The reliability of the solution calorimeter was checked by measuring the standard enthalpy of solution of THAM in an aqueous solution of 0.1 mol dm^{-3} of HCl. The value obtained ($\Delta_s H° = -29.54 \pm 0.3$ kJ mol^{-1}) is in good agreement with that of -29.71 kJ mol^{-1}given in the literature.[19]

For the determination of the standard enthalpy of solution of the ligand, salt, or complex, glass ampules were filled with an accurate amount of the compound and sealed. These were then placed in the calorimetric vessel containing the appropriate solvent mixture (50 cm^3). The system was placed in a thermostated bath until equilibrium was reached. The ampules were broken, and the resulting temperature changes were recorded. After each experiment an electrical calibration was performed. The total heat was corrected for the heat associated with the breaking of the empty glass ampules in the reaction vessel containing the solvent (50 cm^3). The heat of solution was calculated by subtracting the heat of the empty ampule from the total recorded heat. For the determination of the enthalpies of solution of the sodium complex, these experiments were carried out in the presence of a small excess of the ligand in the reaction vessel to ensure that no dissociation of the complex occurred during these measurements.

Results and Discussion

Conductometric Titrations. Plots of Λ_m (molar conductance) against the ligand/metal cation ratio for the titration of Na$^+$ (perchlorate as the counterion) with **1** in different acetonitrile/ N,N-dimethylformamide mixtures at 298.15 K are shown in Figure 1. In all cases, the break in the titration curve at a ligand/ metal cation (L/M$^+$) ratio of 1:1 indicates that one unit of **1** interacts with one metal cation in these solvents. In all cases, there is a decrease in the molar conductance as a result of complex formation. This is expected due to the size increase of the cation in moving from the free to the complex state. It can be seen from the titration curves (Figure 1) that the break of the titration curve becomes less pronounced, as the mixture composition becomes richer in N,N-dimethylformamide, thus indicating that the stability of complex formation becomes weaker. This effect is even more pronounced in the titration of this cation with **1** in MeCN and DMF as discussed elsewhere.[15] Having established the complex composition, we proceeded with
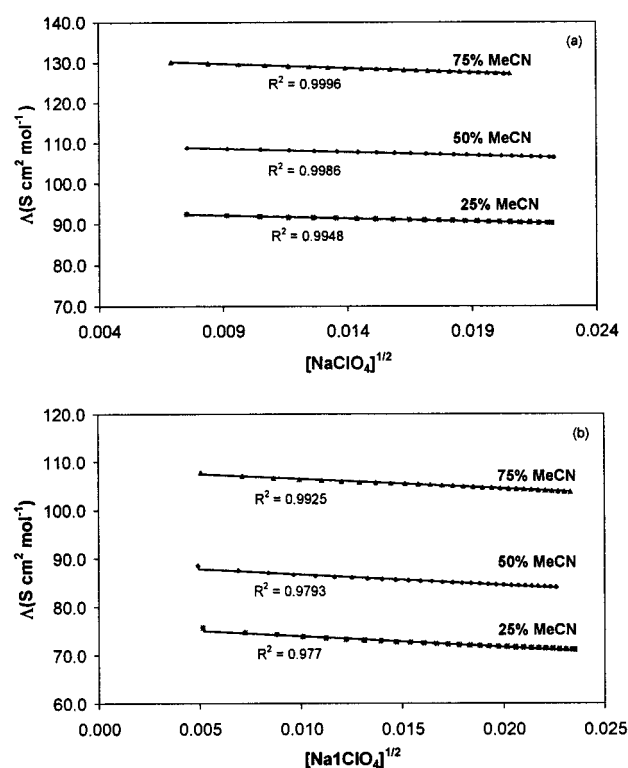
Figure 2. Molar conductances (S cm^2 mol^{-1}) of NaClO$_4$ (a) and Na**1**ClO$_4$ (b) as a function of the square root of molar concentrations in acetonitrile/N,N-dimethylformamide mixtures at 298.15 K.

TABLE 1: Limiting Molar Conductances, Λ^0_m and Ion-pair Formation Constants, K_a for NaClO$_4$ and Na1ClO$_4$ in Acetonitrile, N,N-dimethylformamide, and Their mixtures at 298.15 K

| | $\Lambda_m°/S$ cm^2 mol^{-1} | | K_a | |
solvent	NaClO$_4$	Na**1**ClO$_4$	NaClO$_4$	Na**1**ClO$_4$
MeCN	180.6[a]	145.3[b]	11	diss
75% MeCN	131.6 ± 0.4	108.5 ± 0.3	diss	diss
50% MeCN	110.2 ± 0.3	89.0 ± 0.4	diss	diss
25% MeCN	93.4 ± 0.3	76.2 ± 0.3	diss	diss
DMF	82.3[a]	73.4[b]	diss	diss

[a] Ref 20. [b] Ref 15, diss: fully dissociated

conductance measurements of the sodium complex salt in order to ensure that ions, rather than ion pairs are predominant in solution as demanded by eq 1 and this is now discussed.

*Conductance Measurements of the Sodium **1** Complex Salt in MeCN/DMF Mixtures at 298.15 K.*

Conductance measurements for sodium perchlorate and its metal−ion complex salt at different electrolyte concentrations were carried out in MeCN/DMF mixtures in order to assess that in these solvent systems, sodium perchlorate and the sodium-**1** complex salts are predominant in their ionic forms. In fact, plots of the molar conductance, Λ_m against the square root of the concentration, $c^{1/2}$, for NaClO$_4$ and Na**1**ClO$_4$ in three MeCN-DMF mixtures at 298.15 K (Figure 2) show that these salts behave as strong electrolytes in these solvent mixtures. Straight lines were obtained in all these cases, and the intercepts were used to estimate the values of the limiting molar conductance. These experiments were repeated three times for each mixture. The average values are shown in Table 1 together with the corresponding data in the pure solvents.[15] The standard deviations of the data (σ) were calculated from eq 4 and these are also included in Table 1. It can be easily noted that the

Thermodynamics of Complexation

J. Phys. Chem. B, Vol. 105, No. 31, 2001 **7545**

TABLE 2: Thermodynamic Parameters of Complexation of Sodium and 1 in Different *N*,*N*-dimethylformamide/Acetonitrile Mixtures at 298.15 K. Comparison with Data in the Pure Solvents

% MeCN	log K_s	$\Delta_c G^0$/ kJ mol^{-1}	$\Delta_c H^0$/ kJ mol^{-1}	$\Delta_c S^0$/ J K^{-1} mol^{-1}
100	8.89 ± 0.03^a	-50.74 ± 0.16^a	-76.1 ± 1.6^a	-85.1^a
75	7.45 ± 0.02	-42.52 ± 0.13	-62.4 ± 0.1	-66.6
50	6.52 ± 0.03	-37.21 ± 0.16	-64.1 ± 0.4	-90.1
25	5.66 ± 0.03	-32.30 ± 0.16	-65.6 ± 0.4	-111.6
0	3.43 ± 0.03^a	-19.55 ± 0.15^a	-49.3 ± 1.3^a	-99.9^a

^{a} Ref 15.

Λ_m^0 values decrease as the amount of DMF in the mixtures increases

$$\sigma = \left(\frac{\sum(x_i - \bar{x})^2}{n - 1}\right) \quad (4)$$

Thermodynamics of Complexation. Stability constants (log K_s) and derived standard Gibbs energies, $\Delta_c G°$, enthalpies, $\Delta_c H°$ and entropies, $\Delta_c S°$ of complexation of **1** and Na$^+$ in three different solvent mixtures at 298.15 K are shown in Table 2. Standard deviation of the data were calculated using eq 4. Also included in this Table are data reported for the complexation process involving this cation and this ligand in the pure solvents.[15] Stability constants were derived from data obtained by direct potentiometry using the sodium electrode. Typical potentiometric titration curves for the titration of sodium (perchlorate as the counterion) with **1** in 75% MeCN/25% DMF, 50% MeCN/50% DMF, and 25% MeCN/75% DMF are shown in Figure 3 where potential readings in mV are plotted against the volume of titrant added. In all cases a potential jump at the 1:1 L/M$^+$ ratio was observed, indicating that one metal cation interacts with one ligand unit as demonstrated from conductance measurements. Figure 3 also shows that the largest potential jump is observed in the mixture richest in acetonitrile (75% MeCN/25% DMF), and it decreases upon the addition of DMF. These results indicating a drop in the stability constant of this system as the DMF content of the mixture increases are in accord with the behavior observed in the conductometric titration curves discussed above.

The enthalpy data have been determined by titration calorimetry and the entropy values were calculated through the relationship

$$\Delta_c G° = \Delta_c H° - T\Delta_c S° \quad (5)$$

Quantitative data for log K_s confirms the semiquantitative assessment discussed above based on conductance and potentiometric titration curves in that the complex stability decreases with an increase of DMF in the mixture. This was also observed for complexation processes involving crown ethers in the same solvent mixtures.[7,8] In fact, a linear relationship exists between log K_s values and the solvent composition, as shown in Figure 4.

As far as the enthalpies are concerned, in all cases, the process is enthalpically controlled and entropically destabilized. However, the highest enthalpic stability is observed in neat acetonitrile and the lowest in neat *N*,*N*-dimethylformamide.[15] No appreciable complexation enthalpy changes are observed in the mixed solvents and within the experimental error, complexation enthalpies in these solvents are approximately the same. This behavior in terms of entropy clearly indicates the increase in stability (hence more favorable standard Gibbs energies) as the solvent composition becomes richer in acetonitrile. These
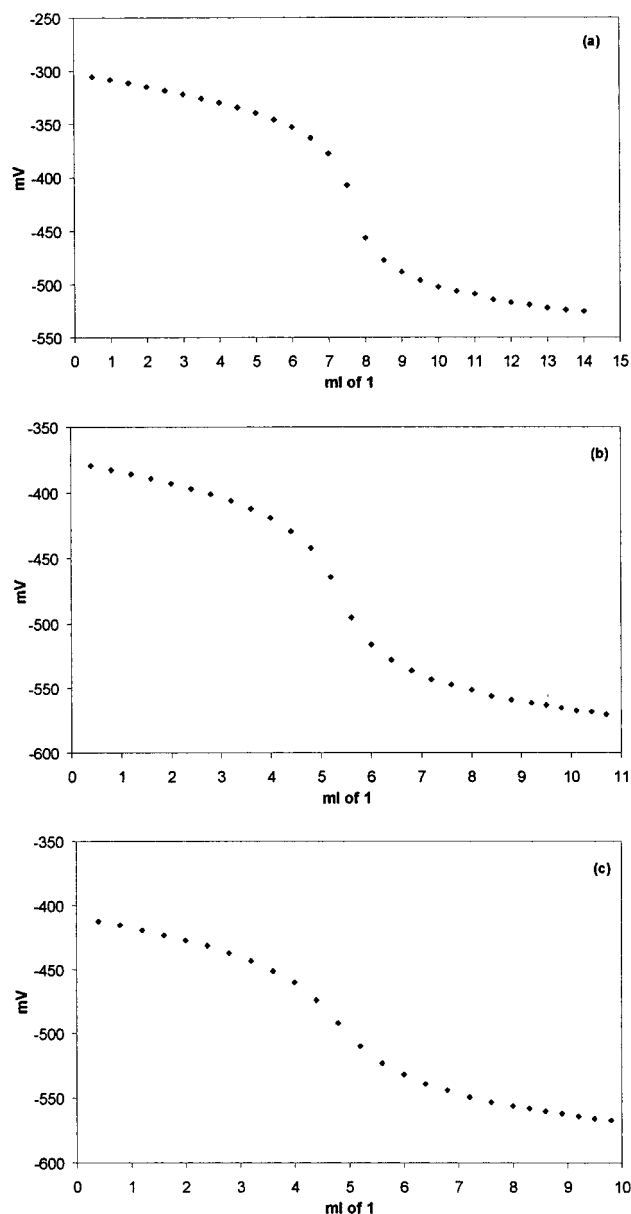
Figure 3. Potentiometric titration curves on the addition of **1** into the vessel containing sodium cations in 75% MeCN/25% DMF (a), 50% MeCN/50% DMF (b), 25% MeCN/75% DMF (c) at 298.15 K.
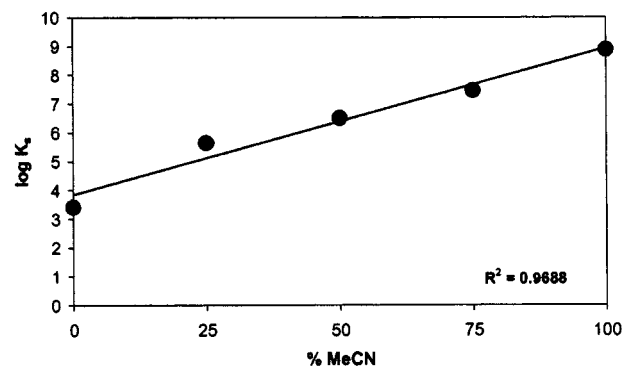
Figure 4. Stability constants (log K_s) of sodium and **1** as a function of the content of MeCN in the MeCN/DMF mixtures at 298.15 K.

changes may be related to the different structures of the solvent released when the cation enters the cavity of the ligand or the different conformational changes that the calix(4)arene undergoes upon complexation in these solvents. There is some

TABLE 3: Enthalpies of Solution of NaClO₄ in MeCN/DMF Mixtures at 298.15 K

87.5% MeCN/12.5% DMF		75% MeCN/25% DMF	
$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$	$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$
6.08×10^{-4}	-26.42	3.97×10^{-4}	-29.26
6.97×10^{-4}	-27.09	6.21×10^{-4}	-29.75
1.12×10^{-3}	-27.50	9.08×10^{-4}	-30.45
2.07×10^{-3}	-26.17	1.06×10^{-3}	-31.81
2.87×10^{-3}	-25.66	1.82×10^{-3}	-30.76
3.66×10^{-3}	-26.14	3.17×10^{-3}	-30.97
$\Delta_s H^\circ = -26.5 \pm 0.7$ kJ mol^{-1}		$\Delta_s H^\circ = -30.5 \pm 0.9$ kJ mol^{-1}	

50% MeCN/50% DMF		25% MeCN/75% DMF	
$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$	$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$
9.02×10^{-4}	-35.59	8.09×10^{-4}	-36.62
1.16×10^{-3}	-32.35	2.23×10^{-3}	-36.11
1.66×10^{-3}	-33.99	3.77×10^{-3}	-35.37
2.45×10^{-3}	-34.48	5.00×10^{-3}	-38.50
4.31×10^{-3}	-32.15	6.05×10^{-3}	-37.15
5.48×10^{-3}	-33.30	7.94×10^{-3}	-36.11
$\Delta_s H^\circ = -33.6 \pm 1.3$ kJ mol^{-1}		$\Delta_s H^\circ = -36.6 \pm 1.1$ kJ mol^{-1}	

TABLE 4: Standard Enthalpies of Solution of Sodium Perchlorate and Heats of Breaking of Empty Ampules in Several MeCN/DMF Mixtures at 298.15 K

NaClO₄		ampules
% MeCN	$\Delta_s H^\circ$ [kJ mol^{-1}]	Q^a_{ab}/J
0	-41.8^b	0.0105
25	-36.6	0.0410
50	-33.6	0.0714
75	-30.5	0.1065
87.5	-26.5	0.1250
100	-17.3^c	0.1506

[a] Heat of breaking of empty ampules in the appropriate solvent [b] Ref 22. [c] Ref 23.

evidence that *N,N*-dimethylformamide is more structured than acetonitrile.[20] However, to interpret the complexation data we proceeded with the determination of the solution thermochemical properties of the reactants (metal−ion salt and ligand) and the product (metal−ion complex salt) in order to calculate the transfer parameters among the various solvents.

Standard Enthalpies of Solution. The enthalpies of solution of NaClO₄ in different solvent mixtures as a function of the electrolyte concentration are reported in Table 3. Because no appreciable changes in $\Delta_s H$ are observed with the electrolyte concentration, the standard enthalpy of solution, $\Delta_s H^\circ$ was taken as the average value of the data reported in Table 3 at 298.15 K. It should be noted that the heats of breaking of empty ampules decreases linearly with the composition of the solvent mixtures from the largest in MeCN (high vapor pressure) to the smallest in DMF (low vapor pressure).[21]

Standard enthalpies of solution of NaClO₄ are negative in the mixtures as well as in the pure solvents indicating that in the dissolution of this compound, the solvation process predominates over the crystal lattice process. Standard enthalpies of solution are listed in Table 4 together with the data from the literature for NaClO₄ in the neat solvents.[22,23] Since the crystal lattice energy is the same, the enthalpy changes reflect the changes in solvation of NaClO₄ in the neat and mixed solvents. The results show that the enthalpic stability decreases from *N,N*-dimethylformamide (a protophilic dipolar aprotic solvent) to acetonitrile (a protophobic dipolar aprotic solvent). It is important to emphasize that these changes do not vary linearly with the composition of the solvent mixture.

As far as the ligand is concerned, in all the solvent mixtures the heats observed were those of breaking of the empty ampules.

TABLE 5: Enthalpies of Solution of Na1ClO₄ Complex in MeCN, DMF, and MeCN/DMF Mixtures at 298.15 K

MeCN		DMF	
$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$	$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$
1.42×10^{-4}	-48.05	2.21×10^{-4}	-39.58
3.39×10^{-4}	-46.60	3.27×10^{-4}	-39.90
5.54×10^{-4}	-47.93	4.34×10^{-4}	-38.75
5.70×10^{-4}	-47.94	5.11×10^{-4}	-39.35
6.98×10^{-4}	-47.67	8.05×10^{-4}	-39.72
9.93×10^{-4}	-47.03	1.31×10^{-3}	-39.29
1.16×10^{-3}	-47.20		
$\Delta_s H^0 = -47.5 \pm 0.5$ kJ mol^{-1}		$\Delta_s H^\circ = -39.4 \pm 0.4$ kJ mol^{-1}	

75% MeCN/25% DMF		50% MeCN/50% DMF	
$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$	$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$
4.35×10^{-4}	-46.55	1.94×10^{-4}	-51.18
5.44×10^{-4}	-47.81	4.59×10^{-4}	-52.48
6.76×10^{-4}	-48.35	5.59×10^{-4}	-52.80
9.78×10^{-4}	-47.92	7.20×10^{-4}	-51.01
1.29×10^{-3}	-48.20	1.16×10^{-3}	-51.36
$\Delta_s H^0 = -47.8 \pm 0.7$ kJ mol^{-1}		$\Delta_s H^\circ = -51.8 \pm 0.8$ kJ mol^{-1}	

25% MeCN/75% DMF	
$c/\text{mol dm}^{-3}$	$\Delta_s H/\text{kJ mol}^{-1}$
3.56×10^{-4}	-56.94
4.77×10^{-3}	-56.79
5.85×10^{-4}	-57.27
8.44×10^{-3}	-58.53
9.64×10^{-3}	-57.70
1.18×10^{-3}	-58.35
$\Delta_s H^0 = -57.4 \pm 0.7$ kJ mol^{-1}	

Therefore, it is reasonable to conclude that the enthalpy of solution of **1** in the mixtures is close to 0 kJ mol^{-1}. $\Delta_s H^\circ$ values for **1** in pure acetonitrile and N, N-dimethylformamide at 298.15 K have been previously reported.[15]

Table 5 shows the enthalpies of solution of the Na1ClO₄ complex in the pure solvents and in the mixtures at 298.15 K at different electrolyte concentrations. The average of these values are taken as the standard enthalpies of solution of the metal−ion complex salt in these solvents at 298.15 K because hardly any variations are observed in the $\Delta_s H$ values for different electrolyte concentrations. In the 25% DMF/75% MeCN mixture the interaction with acetonitrile predominates, since the presence of DMF does not change the $\Delta_s H^\circ$ from that in neat acetonitrile. As the composition of the mixture becomes richer in *N,N*-dimethylformamide (50% DMF/50% MeCN and 75% DMF/25% MeCN) the enthalpic stability increases. Because for the same compound, the crystal lattice enthalpy value is the same, this is indicative of an increase in the complex−solvent interactions. However, in the absence of acetonitrile, the enthalpic stability decreases again.

Before proceeding with the calculation of $\Delta_t H^\circ$ values from a reference solvent to another solvent, the reliability of the $\Delta_s H^\circ$ and $\Delta_c H^\circ$ values is checked by the calculation of the standard coordination enthalpy, $\Delta_{coord} H^\circ$, which is referred to the process in which the reactants and the product are in the solid (sol) state (eq 6)[24]

$$\text{NaClO}_4 \text{ (sol)} + \mathbf{1} \text{ (sol)} \rightarrow \text{Na}\mathbf{1}\text{ClO}_4 \text{ (sol)} \qquad (6)$$

As pointed out before,[24] for a given system, the $\Delta_{coord} H^\circ$ should be the same independently of the solvent from which these data are derived, and therefore, these calculations provide a suitable means to check the accuracy of the experimental data. Details given in Table 6 show the $\Delta_{coord} H^\circ$ for the Na1ClO₄ system derived from neat and mixed solvents. These quantities were

Thermodynamics of Complexation

J. Phys. Chem. B, Vol. 105, No. 31, 2001 **7547**

TABLE 6: Enthalpies of Coordination for Na1ClO$_4$ Calculated from Complexation and solution enthalpies of reactants and product in neat and mixed solvents at 298.15 K

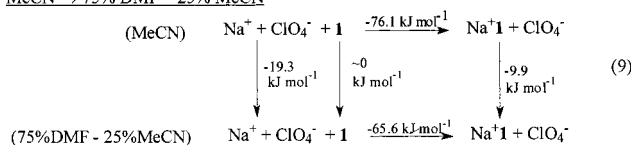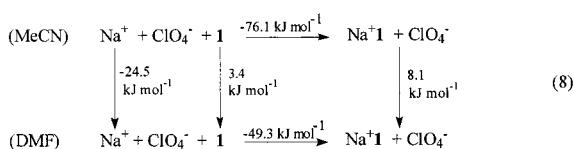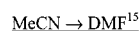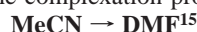|	$\Delta_c H°$ /kJ mol^{-1}	$\Delta_s H°$ /kJ mol^{-1}			$\Delta_{coord}H°$ /kJ mol^{-1}
solvent	Na$^+$ + **1** (solution)	NaClO$_4$	**1**	Na**1**ClO$_4$	NaClO$_4$(sol.) + **1** (sol.) → Na**1**ClO$_4$ (sol.)
MeCN	−76.1^a	−17.3^b	∼0	−47.5	−45.9
DMF	−49.3^a	−41.8^b	3.4	−39.4	−48.3
25% MeCN	−65.6	−36.6	∼0	−57.4	−44.8
50% MeCN	−64.1	−33.6	∼0	−51.8	−45.9
75% MeCN	−62.4	−30.5	∼0	−47.8	−45.1

^{a} From Table 2. ^{b} From Table 4.

calculated from the following equation

$$\Delta_{coord}H° = \Delta_s H° \text{ (NaClO}_4\text{) (s)} + \Delta_s H° \text{ (1) (s)} + \Delta_c H° \text{ (s)} - \Delta_s H° \text{ (Na1ClO}_4\text{) (s)} \quad (7)$$

Considering the number of individual processes involved in the derivation of $\Delta_{coord}H°$, excellent agreement is found between these sets of data. An average is taken as the standard enthalpy of coordination for this system in the solid state.

Having checked the reliability of the data we proceeded with the calculation of $\Delta_t H°$ values, taking acetonitrile as the reference solvent. $\Delta_t H°$ values for the free and complex salts and the ligand to the various solvents are calculated and together with the $\Delta_c H°$ values (Table 2) these are inserted in the thermodynamic cycle[14] in order to assess the medium effect on the complexation process (eq 1)

MeCN → DMF[15]

MeCN → DMF[15]

$$(8)$$

MeCN → 75% DMF/25% MeCN

MeCN → 75% DMF – 25% MeCN

$$(9)$$

MeCN → 50% DMF/50% MeCN

MeCN → 50% DMF – 50% MeCN

$$(10)$$

MeCN → 25% DMF − 75% MeCN

MeCN → 25% DMF – 75% MeCN

$$(11)$$
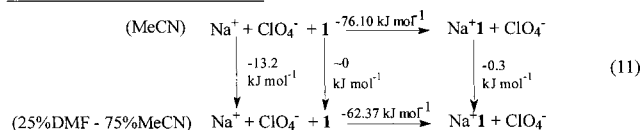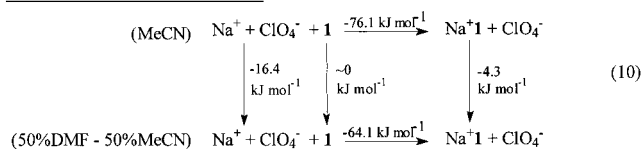
There are data available on the $\Delta_t H°$ value for the perchlorate anion based on the Ph$_4$AsPh$_4$B convention at 298.15 K from the acetonitrile to the solvent mixtures. The $\Delta_t H°$ value for this anion transferring from acetonitrile to *N,N*-dimethylformamide, based on this convention, is about 5 kJ mol^{-1}.[25] Therefore, the single-ion value from neat acetonitrile to the mixed solvents must be between 0 (MeCN → MeCN) and 5 kJ mol^{-1}. However, it should be emphasized that although the contribution of the counterion is included in the $\Delta_t H°$ values of the free and complex electrolytes, as far as $\Delta(\Delta_c H°)$ is concerned, the anion contribution is canceled out.

Equation 1 implies that the differences in enthalpic stability for complex formation between two solvents (s$_1$ and s$_2$) depend on the transfer enthalpies of the reactants (r) and the product (p). Thus, a process will be enthalpically more stable in s$_1$ than in s$_2$ if

Overall $\Delta_t H°$ (r) (s$_1$ → s$_2$) < 0 and $\Delta_t H°$ (p) (s$_1$ → s$_2$) ≥ 0

(i)

Overall $\Delta_t H°$ (r) (s$_1$ → s$_2$) < 0 and $\Delta_t H°$ (p) (s$_1$ → s$_2$) ≤ 0

(ii)

provided $\Delta_t H°$ (r) (s$_1$ → s$_2$) < $\Delta_t H°$ (p) (s$_1$ → s$_2$)

Overall $\Delta_t H°$ (r) (s$_1$ → s$_2$) ≥ 0 and $\Delta_t H°$ (p) (s$_1$ → s$_2$) > 0

(iii)

provided $\Delta_t H°$ (p) (s$_1$ → s$_2$) > $\Delta_t H°$ (r) (s$_1$ → s$_2$)

The data in the cycles show that whereas in the neat solvents the more negative enthalpy of complexation in MeCN relative to DMF falls within (i), this is not necessarily the case when the solvent mixtures are involved. Particularly striking is the reverse trend observed in the $\Delta_t H°$ values of the metal-ion complex salt from MeCN to 75%DMF/25%MeCN mixture relative to the corresponding data of transfer to DMF. It is the decrease in stability (in enthalpic terms) of the free metal cation salt combined with the enhancement in the $\Delta_t H°$ value for the metal-ion complex salt from MeCN to the mixture with respect to DMF, which leads to an increase in the exothermic character of the complexation process in the former relative to the latter solvent.

Unlike the enthalpic stability for NaClO$_4$, that of Na**1**ClO$_4$ is higher in MeCN than in DMF. The X-ray structure of the sodium complex perchlorate has shown that one molecule of acetonitrile interacts with the hydrophobic cavity of the complex ligand.[15] If the same occurs in solution, then this would lead to a higher enthalpic stability in MeCN than in DMF as it is observed. There is hardly any change in the $\Delta_t H°$ value for Na**1**ClO$_4$ on moving from the reference solvent to the mixture rich in MeCN (75%MeCN/25%DMF). However as the DMF content in the mixture increases (50%MeCN/50%DMF, 25%MeCN/75%DMF), the $\Delta_t H°$ values for Na**1**ClO$_4$ becomes more negative.

However, the absence of acetonitrile in the mixture reverses the $\Delta_t H°$ of the process because the $\Delta_t H°$ of Na**1**ClO$_4$ between the neat solvents clearly shows that this compound is enthalpically more stable in MeCN than in DMF. The $\Delta_t H°$ values to the mixtures (50%−50% and 25%−75%) are exothermic, and therefore, these differ significantly from that to DMF (8 kJ mol^{-1}, endothermic). These differences are appreciable and outside any experimental error. It appears that the presence of DMF in the mixture induces a higher interaction of the metal-ion complex salt with acetonitrile relative to neat MeCN.
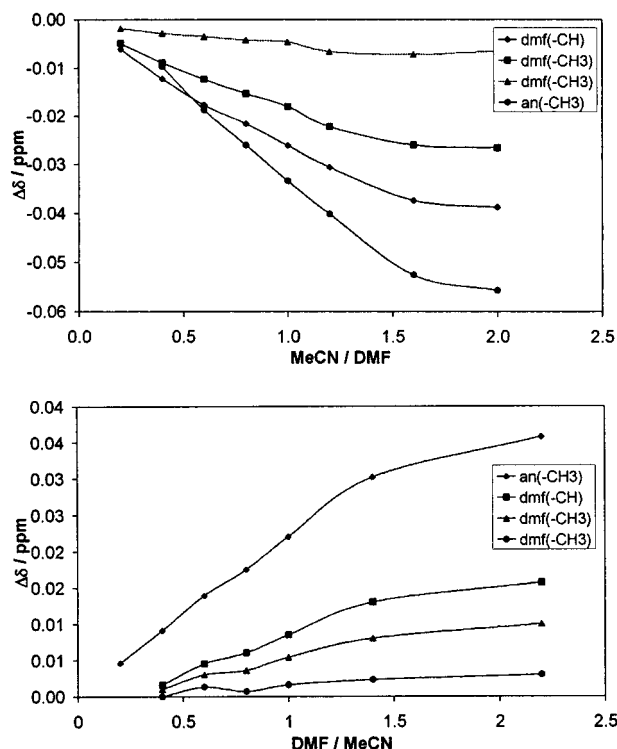
Figure 5. Proton chemical shift changes (ppm) as a function of the MeCN/DMF ratio at 298 K resulting from the titration of DMF with MeCN
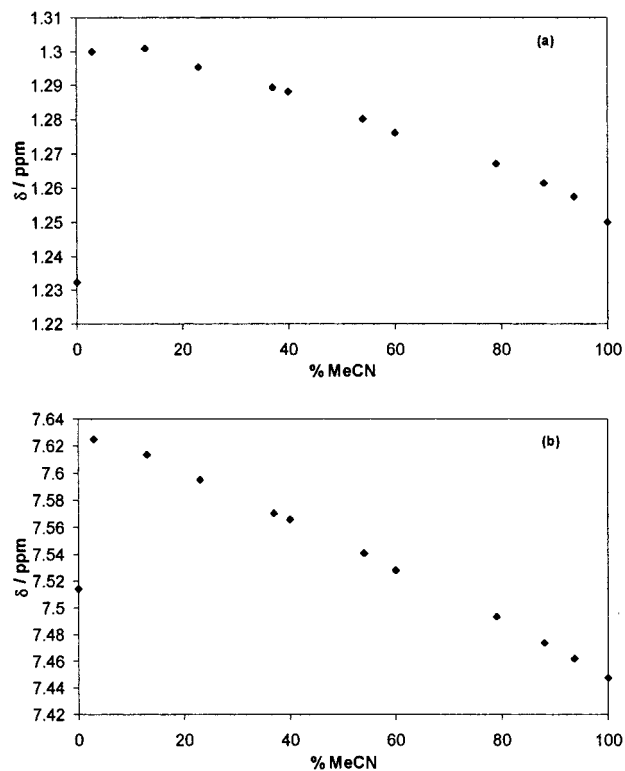

Figure 6. Chemical shifts for the *t*-butyl (a) and aromatic protons (b) at the upper rim of the sodium **1** complex (perchlorate as counter-ion) in different MeCN/DMF mixtures at 298 K.

To investigate the reverse trend observed in the $\Delta_t H°$ Na$^+$**1**ClO$_4^-$ from acetonitrile to the mixed solvents (exothermic) relative to that to DMF (endothermic), and particularly, the increase in the exothermic character of the transfer process involving the metal–ion complex as the content of DMF in the mixtures increases (see cycles, eqs 8–11) we carried out

(i) ^{1}H NMR titrations of MeCN (a mixture of MeCN and deuterated MeCN) with *N,N*-dimethylformamide and vice versa to gain information regarding the interaction between these two solvents.

(ii) ^{1}H NMR studies on the metal-ion complex salt in CD$_3$CN, in d$_7$ – DMF and in the mixtures.

The aim of these measurements was to assess the effect of the solvent mixture relative to the neat solvents on the aromatic and *tert*-butyl protons of the upper rim of the complex calixarenes. These are likely to be the most affected protons if indeed an enhancement of the interaction between the methyl group of MeCN and the hydrophobic cavity of the complex ligand occurs in the presence of DMF.

The results of item (i) are shown in Figure 5. In all cases, changes on the chemical shifts are exhibited by the protons of the methyl group of MeCN as well as by the proton attached to the carbonyl groups in DMF. The observed changes in the chemical shifts may be ascribable to an interaction of the dipole–dipole type between the two solvents, involving the nitrogen of MeCN and the carbonyl group of DMF. Indeed, it would be expected that as a result of the above interaction, the shielding of the methyl group of MeCN would decrease slightly because there will be a decrease in its electronic density. For the carbonyl group of DMF, there will be a slight increase in the sp^3 character of its carbon atom as a result of the interaction and consequently, a small increase in shielding should lead to a smaller chemical shift. This decrease in the electronic density of the methyl group of MeCN in the presence of DMF is likely to have implications on the strength of interaction of MeCN

(through its methyl group) with the hydrophobic cavity of the complex ligand as discussed below.

(iii) ^{1}H NMR measurements carried out with the Na**1**ClO$_4$ complex in different MeCN − DMF mixtures and in the neat solvents show that except for the aromatic and *t*-butyl protons at the upper rim, the chemical shift changes in the mixture fall within the values observed in MeCN and DMF.

For the aromatic and *tert*-butyl protons of the hydrophobic cavity a different behavior was observed. In pure DMF all protons (except those of the *tert*-butyl group) are more deshielded than in MeCN. However, in the mixtures, the aromatic protons at the upper rim became more deshielded upon the addition of DMF into the solution of Na**1**ClO$_4$ in MeCN to an extent that when the content of DMF in the mixture reaches 40%, these protons are more deshielded than in pure DMF. In the case of the *tert*-butyl protons, in all the mixtures these are more deshielded than in the neat solvents. These statements are corroborated by the data presented in Figure 6. In these figures, δ (ppm) values for the *tert*-butyl group (a) and the aromatic (b) are plotted against the compositions of the different solvent mixtures.

The ^{1}H NMR data for the complex were compared with those for the free ligand in neat DMF, MeCN and 60% DMF/40% MeCN mixtures as shown in Figure 7. In the latter, the shifts of the upper rim aromatic and *tert*-butyl protons fall between the values in the pure solvents and the same applies to the $\Delta_s H°$ for **1** in these solvents. (see Table 6)

From the ^{1}H NMR measurements discussed above, it is concluded that the aromatic and the *tert*-butyl protons at the upper rim of the calixarene complex are more affected in the mixtures than the other protons. X-ray diffraction studies of the Na**1**ClO$_4$ complex showed that one molecule of acetonitrile interacts through its methyl group with the hydrophobic cavity of the ligand,[15] and this is likely to occur in solution. Therefore,

Thermodynamics of Complexation

J. Phys. Chem. B, Vol. 105, No. 31, 2001 **7549**
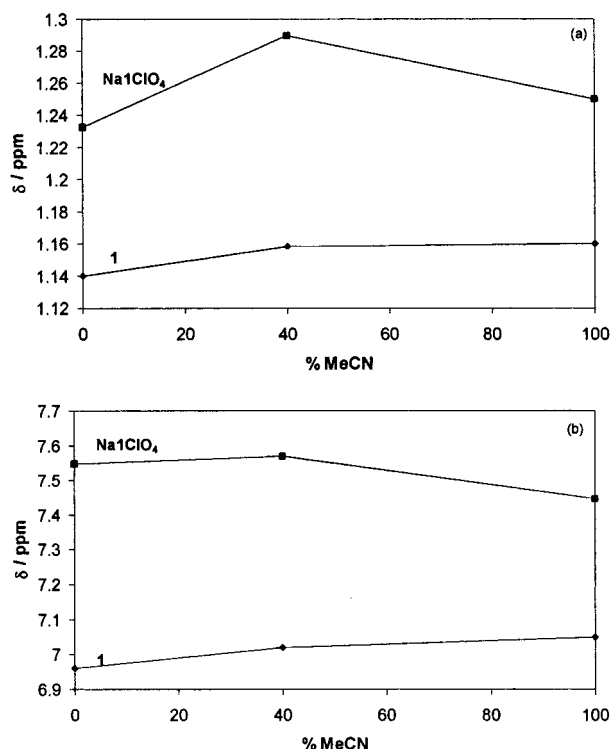
Figure 7. Chemical shifts for the *t*-butyl (a) and aromatic protons (b) at the upper rim of the ligand **1** and the Na1ClO4 complex in different MeCN/DMF mixtures at 298 K.

in the mixtures the interaction of MeCN and DMF decreases the electronic density of the methyl group in the former solvent in a way that its interaction with the hydrophobic cavity of the complex is enhanced.

This is in agreement with the higher enthalpic stability observed in the $\Delta_t H°$ values of the Na**1**ClO4 complex from MeCN to the mixtures relative to the neat solvents. As expected from these types of interactions, the ^{1}H NMR changes are relatively small. Equally, the enthalpy variations observed in the $\Delta_t H°$ values of the Na**1**ClO4 complex from MeCN to the mixtures are well within the range expected for a van der Waals type of interactions.

Final Remarks

The results reported in this paper encourage research on the thermodynamics of cation complexation process involving calixarene derivatives in mixed solvents.

However these studies require knowledge about solvent–solvent interactions in the mixture and their implications on the binding process. Again, the distinctive properties of calix(4)-arene derivatives (two different cavities) with respect to crown ethers (single holes) are demonstrated. For both groups of macrocyclic ligands, the thermodynamic parameters of complexation are sensitive to the composition of the mixed solvent. As the amount of acetonitrile increases in the MeCN/DMF binary mixture, the stability constant of the complexes increases linearly. All complexes are enthalpy stabilized but entropy destabilized. However, in the case of crown ethers, the thermodynamic parameters of complexation vary monotonically with the solvent composition. A decrease in the concentration

of DMF in the mixture enhances the complex stability due to a more favorable enthalpy and entropy.

In the case of the calix(4)arene derivative, the presence of two cavities leads the ligand to interact through its hydrophilic cavity with the metal cation while interactions with the solvent molecules take place in the hydrophobic cavity. Therefore, the complexation behavior for the systems involving calix(4)arenes is less predictable than those involving crown ethers. For the tetraphenyl ketone derivative and the sodium cation, an increase in the stability of the complex with a decrease of the DMF content in the mixture is also due to a more favorable entropy but the $\Delta_c H°$ values remain constant in all the mixtures.

References and Notes

(1) Popov, A. I.; Lehn, J. M. In *Coordination Chemistry of Macrocyclic Compounds*; Melson, G. A., Ed.; Plenum Press: New York, 1985, Chapter 9.

(2) Izatt, R. M.; Pawlak, K.; Bradshaw, J. S.*Chem. Rev.*, **1991**, *91*, 1721.

(3) Lamb, J. D.; Izatt, R. M.; Christensen, J. J. In *Coordination Chemistry of Macrocyclic Compounds*; Melson, G. A., Ed.; Plenum Press: New York, **1979**, Chapter 3.

(4) Izatt, R. M.; Bradshaw, J. S.; Nielsen, S. A.; Lamb, J. D.; Christensen, J. J. *Chem. Rev.* **1985**, *85*, 271.

(5) Danil de Namor, A. F.; Zapata-Ormachea, M. L.; Hutcherson, R. G. *J. Phys. Chem. B* **1998**, *102*, 7839.

(6) Danil de Namor, A. F.; Zapata-Ormachea, M. L.; Hutcherson, R. G. *J. Phys. Chem. B* **1999**, *103*, 366.

(7) Rounaghi, G.; Eshaghi, Z.; Ghiamati, E. *Talanta* **1997**, *44*, 275.

(8) Rounaghi, G.; Popov, A. I. *Inorg. Chim. Acta* **1986**, *114*, 145.

(9) Lehn, J. M. *Supramolelular Chemistry*; VCH: New York, 1995.

(10) Gutsche, C. D. Calixarenes *In Monographs in Supramolecular Chemistry;* J. F. Stoddart, Ed.; The Royal Society of Chemistry: London 1989.

(11) *Calixarenes: A Versatile Class of Macrocyclic Compounds*; Vicens, J., Böhmer, V., Eds.; Kluwer: Dordrecht **1991**. *Calixarenes 50th Anniversary; Commemorative Volume*; Vicens, J.; Asfari, Z.; Harrowfield, J. M., Eds.; Kluwer: Dordrecht, 1994.

(12) Gutshe, C. D. *Calixarenes Revisited*; Royal Society of Chemistry: Cambridge, 1998.

(13) Danil de Namor, A. F. *Coord. Chem. Rev.* **1999**, 190−192, 283.

(14) Danil de Namor, A. F.; Cleverley, R. M.; Zapata-Ormachea, M. L. *Chem. Rev.* **1998**, *98*, 7, 2495.

(15) Danil de Namor, A. F.; Kowalska, D.; Castellano, E. E.; Piro, O. E.; Sueros Velarde, F. J.; Villanueva Salas, J., submitted for publication, **2001**.

(16) Perrin, D. D.; Armarego, W. L. F. *Purification of Laboratory Chemicals*, 3rd ed.; Pergamon Press: New York, 1988.

(17) Eatough, D. J.; Christensen, J. J.; Izatt, R. M. In *Experiemental in Thermometric, Tritrimetry and Titration Calorimetry*; Brigham Young University Press: Provo, Utah, 1974.

(18) Wilson, E. W.; Smith, D. F. *Anal. Chem.* **1969**, *41*, 13, 1903.

(19) Hill, J. O.; Ojelund, G.; Wadsö, I. *J. Chem. Thermodyn.* **1969**, *1*, 111.

(20) Fernandez-Prini, R. In *Physical Chemistry of Organic Solvent Systems*; Covington, A. K., Dickinson, T., Eds.; Plenum Press: New York, 1973.

(21) *Handbook of Chemistry and Physics*, 62nd ed.; Ed. Weast, R. C.; Astle, M. J. CRC Press, Boca Raton, Fl., 1982.

(22) Cox, B. G.; Hedwig, G. R.; Parker, A. J.; Watts, D. W. *Aust. J. Chem.* 1974, 27, 477.

(23) Danil de Namor, A. F.; Pulcha Salazar, L. E.; Llosa Tanco, M. A.; Kowalska, D.; Villanueva Salas, J.; Schulz, R. A. *J. Chem. Soc., Faraday Trans.* **1998**, *94*, 3111.

(24) Danil de Namor, A. F.; Kowalska, D. *J. Phys. Chem. B*, **1997**, *101*, 1643.

(25) Ghousseini, L. Ph.D. Thesis, University of Surrey, 1985.