Conductance and Thermodynamic Study of the Complexation of Ethyl *p-Tert*-butylcalix[4]arene Tetraacetate with Alkali Metal and Silver Ions in Various Solvents

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Abstract

The first conductometric studies for the complexation reactions between alkali metal and silver cations with ethyl *p-tert*-butylcalix[4]ester in the cone conformation using acetonitrile and methanol in the temperature range 278–303 K are reported. The observed molar conductivities Λ were found to decrease significantly for mole ratios less than unity. The conductivity data were analysed using a computer program based upon 1:1 stoichiometry. Stability constants of the resulting 1:1 complexes were determined, indicating that the sequence of stability in acetonitrile is $Na^+ > Li^+ > K^+$. The ΔH^0 and ΔS^0 values of the calix[4]ester complexation reactions were determined from the temperature dependence of the complexation constants, and their significance are discussed.

Introduction

Many compounds which are able to interact in a supramolecular [1] fashion with a wide variety of guests are known. One group of such compounds that are currently under intensive investigation are the calixarenes and their derivatives. Calixarenes are cyclic oligomers formed by the base-induced condensation of formaldehyde and *p-tert*-butylphenol [2]. Calixarenes can be modified at two main sites: at the "lower" rim and/or at the "upper" rim [3]. The aims of these chemical modifications are: (i) to enhance the selectivity and efficiency of their complexation properties, (ii) to control their conformers, and (iii) to enhance their solubilities [4]. There are many instances of lower rim functionalization to form for example, esters, ketones [5], amides, acids, and calixcrowns [2, 6].

Spectrophotometry [5], potentiometry, NMR spectroscopy, calorimetry and to only a very small extent, conductometry, have been mainly used to determine the stability constants and other thermodynamic parameters of the complexation between calixarene derivatives and cations which have focused mainly on alkali metal ions [6].

There are a number of reports that deal with the complexation of alkali metal ions in pure solvents with *tert*-butylcalix[4]arene-*O*-tetraesters which are in the cone conformation. Some of these reports involve the use of UV spectrometry to determine stability constants in methanol or acetonitrile at 298 K [5]. Other reports describe the use of classical titration calorimetry and/or titration microcalorimetry to determine thermodynamic parameters, ΔG^0, ΔH^0 and ΔS^0 at 298 K [7, 8]. There are however, no reports dealing with the effect on the complexation properties, of using pure solvents at different temperatures, nor with the combined effects of changes in solvent and temperature.

The aim of the present work was to determine the temperature- and solvent composition-dependence of ΔG^0, ΔH^0 and ΔS^0 for the complexation of alkali metal and silver ions with ethyl *p-tert*-butylcalix[4]arene-*O*-tetraester ("calixester") in the cone conformation, in the temperature range 278–303 K using conductometry: (a) in pure methanol, and (b) in acetonitrile. Since conductivity has only been used to determine some of the stability constants of complexation reactions of some calixesters with alkali metal ions [6], another aim of this study was to explore the reliability of the conductivity technique as another method for studying the thermodynamics of the complexation of calixesters with metal cations.

Experimental

Ethyl *p-tert*-butylcalix[4]arenetetraethanoate ("calixester") was prepared according to a published procedure [5]. After fractional distillation, methanol (HPLC grade, Scharlau, assay 99.6%) and acetonitrile (HPLC grade, GCC, assay 99.8%) were used as solvents for the calixester and the salts. Conductivities of methanol and acetonitrile were less than 2.0×10^{-7} S cm^{-1} and 1.0×10^{-7} S cm^{-1} respectively. The following salts were obtained from the suppliers indicated: LiClO$_4$ (Aldrich, 95+%), NaClO$_4$ (Sigma, 99%), KClO$_4$ (Fluka chemica, >99.5%), RbClO$_4$ (Aldrich), CsNO$_3$ (Aldrich, 99%) and AgNO$_3$ (Degussa). All salts were dried under vacuum at 70 °C prior to use. Solutions having metal ion concentrations of approximately

1.0–1.1×10^{-4} M were prepared by dissolving a known mass of each salt in the respective solvent. These solutions were also used as solvents for preparing the calixester solutions with concentrations of approximately 1.49–1.56×10^{-3} M. Conductance measurements were carried out with a microprocessor conductivity meter (WTW/LF 537). A calibrated conductivity cell (WTW/Tetracon 96) having a cell constant of 0.618 Cm^{-1} was used. The cell was calibrated using KCl solutions. The temperature of the reaction was controlled to ± 0.1 °C using a thermostatted circulator water bath (HAAKE D1) equipped with a refrigeration unit. In order to determine the complex formation constants with the different metal ions used, 50 ml of the desired salt solution was placed in a specially-designed water-jacketed cell (150 ml, pyrex^{R}) which was equipped with a magnetic stirrer and was connected to the thermostatted circulator water bath at the required temperature. The conductance of the solutions were measured at each of the thermostatted temperatures. Known amounts (0.50 mL) of the solutions of the calixester were added in a stepwise manner using a calibrated pipette having a precision of ± 0.1 mL. The conductivity of the mixture was then measured after stirring and temperature equilibration. This procedure was repeated in the same manner for each addition. Statistical analyses and curve fitting were conducted using *Sigmaplot v 3.0*.

Mathematical treatment

When a ligand (L) such as the calixester forms a $1:1$ complex with a univalent ion (M$^+$), the equilibrium equation is written as:

$$M^+ + L \rightleftharpoons ML^+, \qquad (1)$$
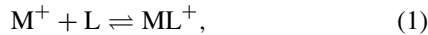
where M$^+$, L and ML$^+$ represent the free solvated cation, the free ligand and the complex, respectively. The thermodynamic equilibrium constant K for the association is given by Equation (2):

$$K = [ML^+]f(ML^+)/[M^+][L]f(M^+)f(L), \qquad (2)$$

where [ML$^+$], [M$^+$], [L] and f represent the equilibrium molar concentrations of the complex, the free cation, the free ligand and the activity coefficients of the species indicated, respectively. Under the dilute conditions used, the activity coefficient of the uncharged ligand, $f(L)$, can be assumed to be unity. The use of the Debye-Hückel limiting law of electrolytes leads to the conclusion that $f(ML^+) \cong f(M^+)$, so that Equation (2) can be rewritten as:

$$K = [ML^+]/[M^+][L]. \qquad (3)$$

If α is taken as the fraction of total metal remaining as free solvated cation the following equations result:

$$[M^+] = \alpha[M^+]_T, \qquad (4)$$

$$[ML^+] = (1 - \alpha)[M^+]_T, \qquad (5)$$

$$[L] = [L]_T - (1 - \alpha)[M^+]_T, \qquad (6)$$

where [M$^+$]$_T$, [L]$_T$, [M$^+$], [ML$^+$] and [L] represent the concentrations (mol dm^{-3}) of total cation, total ligand, free uncomplexed univalent cation, the complex, and the free solvated ligand, respectively. The observed conductivities of the resulting solution can be calculated by using Equation (7):

$$\Lambda = 10^3 \kappa / [M^+]_T, \qquad (7)$$

where κ is the measured conductivity of the formed solution in S.cm^{-1}. Λ can be related to α by Equation (8):

$$\Lambda = \alpha \Lambda_o + (1 - \alpha)\Lambda_c, \qquad (8)$$

where Λ_o and Λ_c represent the molar conductivities of the ligand-free metal salt solution and the complexed cation salt solution, respectively. Substituting Equations (4) and (5) into Equation (3), gives Equation (9):

$$K = (1 - \alpha)/\alpha[L]. \qquad (9)$$

The value of α can be found by rearranging Equation (8) to give Equation (10):

$$\alpha = \Lambda - \Lambda_c / \Lambda_o - \Lambda_c. \qquad (10)$$

Substitution of Equation (10) into Equations (6) and (9) gives (11) and (12) respectively from which Equation (13) can be derived:

$$[L] = [L]_T - \{\Lambda_o - \Lambda / \Lambda_o - \Lambda_c\}[M^+]_T, \qquad (11)$$

$$K = (\Lambda_o - \Lambda)/[L](\Lambda - \Lambda_c), \qquad (12)$$

$$K = a_1 a_2 /(a_2 a_3 [L]_T - a_1 a_3 [M^+]_T), \qquad (13)$$

where

$$a_1 = \Lambda_o - \Lambda, \qquad (14)$$

$$a_2 = \Lambda_o - \Lambda_c, \qquad (15)$$

$$a_3 = \Lambda - \Lambda_c. \qquad (16)$$

Rearranging Equation (13) gives a quadratic equation in Λ as follows:

$$a \overset{2}{\Lambda} + b\Lambda + c = 0, \qquad (17)$$

where

$$a = K[M^+]_T, \qquad (18)$$

$$b = a_2(K[L]_T + 1) - [M^+]_T K(\Lambda_o + \Lambda_c), \qquad (19)$$

$$c = [M^+]_T K \Lambda_o \Lambda_c - a_2(K[L]_T \Lambda_c + \Lambda_o). \quad (20)$$

Solving for Λ from Equation (17) gives

$$\Lambda_{cal} = (-b/2a) + (b^2 - 4ac)^{1/2}/2a. \quad (21)$$

In Equation (17) Λ is treated as a calculated quantity in the simplex program [9], while K and Λ_c are adjustable parameters which are needed to evaluate a, b, and c. The procedure followed in calculating K from the conductivity data involved the use of an approximate value for Λ_c taken from the corresponding Λ vs $[L]_T/[M^+]_T$ graph. This value, together with the known quantities $[L]_T$, $[M^+]_T$, Λ_o and Λ, provides a value for $[L]$ as given by Equation (11). A subroutine is then used to calculate an approximate average value of K for a given series. These approximate values of K and Λ_c are then used as reasonable starting guesses in the program to solve for Λ_{cal} as given by Equation (21). The reported values of K and Λ_c corresponding to the condition, $3(\Lambda - \Lambda_{cal})^2$ is a minimum, at which point the program exits.

Results and discussion

In principle, complex formation can be followed by observing changes in physical properties such as heat of reaction, change in color, conductance or NMR chemical shifts which can occur as a result of the complexation. Such changes provide information about the stoichiometry and stabilities of the complexes. In this study, complex formation and the accompanying changes in thermodynamic parameters were studied by following the change in molar conductance. The molar conductance Λ versus $[L]_T/[M^+]_T$ plots in acetonitrile and methanol at 25 °C are given in Figures 1 and 2 respectively. $[L]_T$ and $[M^+]_T$ are the total concentrations of calixester, and of the metal cations, respectively. Such plots of Λ versus $[L]_T/[M^+]_T$ can provide direct evidence for the stoichiometry of complex formation. As can be seen, in Figure 1, addition of calixester to solutions of Li$^+$, Na$^+$ or K$^+$ but not for M$^+$ = Rb$^+$, Cs$^+$ or Ag$^+$ result in a continuous decrease in the molar conductance of the solution. Also, as shown in Figure 2, addition of the calixester to the solutions of Na$^+$ and Ag$^+$ but not to the solutions of Li$^+$, K$^+$, Rb$^+$ or Cs$^+$ result in a continuous decrease in the molar conductance of the resulting solutions. These decreases in molar conductance indicate a lower mobility of the metal-calixester complexes in acetonitrile and methanol, as compared to the solvated cations alone. The decreases in molar conductance for the metal ions which are shown in each of Figures 1 and 2 start to level off at a mole ratio of 1. The corresponding slopes of these plots for these ions also change at the point where the calixester-to-cation mole ratio is 1, implying the formation of stable 1 : 1 complexes. It is noteworthy that the sharp decreases of the slopes for the Li$^+$ and Na$^+$ metal-calixester systems in acetonitrile indicate that these complexes are more stable compared to the K$^+$-calixester complex. Figures 1 and 2 reveal that despite increases in the calixester concentration,

negligible changes in Λ are found for Rb$^+$, Cs$^+$ and Ag$^+$ in acetonitrile, and for Li$^+$, K$^+$, Rb$^+$ and Cs$^+$ in methanol. As a result therefore, determinations of their complex formation constants with the calixester were not possible. Similar phenomena have also been observed in the conductance behavior of some metal-crown ether complexes [10]. Such behavior can indicate either that these M$^+$-calixester complexes are unstable, or else that the solvated cations and their complexes have equal mobilities. For the solutions in which M$^+$ = Rb$^+$ or Cs$^+$, it can also be argued that either these solvated cations and their complexes in acetonitrile and methanol solutions have equal mobilities or, due to their large sizes, they might be not completely shielded by the ligand. It is more likely that the latter is the case, since it has been shown that it is possible to determine their stability constants in acetonitrile and methanol using UV spectrophotometry [5]. The stability constant however for the Cs$^+$ complexation with the calixester could not be determined calorimetrically since no heat changes could be detected using a microcalorimeter [7].

For M$^+$ = Li$^+$ or K$^+$ in methanol, it is possible that the mobilities of solvated cations are equal to the mobilities of their complexes. In order to account for the negligible change in Λ of Ag$^+$ in acetonitrile, it can be postulated that the Ag$^+$-calixester complex is unstable. It is known that calixesters are poor complexing agents for silver cation in either acetonitrile or benzonitrile [7] due to the fact that Ag$^+$ cations strongly complex with acetonitrile and benzonitrile [11].

The solvent has a significant effect on the stability of the complexes, as the values of log K for the complexes in both acetonitrile and methanol show in Tables 1 and 2. As an example, the conductance behaviour of Na$^+$ changed significantly when the solvent was changed from acetonitrile to methanol. As can be seen in Figure 1, there is a very sharp decrease in the slope of the Na$^+$-calixester curve at the point when $[L]_T/[M^+]_T = 1$, whereas in methanol there is a smooth change in the slope at the same mole ratio. This indicates that the stability of the complex is higher in acetonitrile than in methanol. Indeed, its stability constant, log K, in acetonitrile is equal to 7.40 while in methanol it is 5.48. These differences in log K values reflect the important role played by the solvent in the complexation processes involving macrocyclic ligands. Thus the stability of the complexes increases with decreasing power of the solvents as expressed by the Gutmann number [12]. Danil de Namor *et al.* [8] have argued that acetonitrile enters the hydrophobic cavity of the calixester, which consequently allows the hydrophilic cavity to become better preorganized to interact with cations in acetonitrile than in methanol.

In order to achieve a better understanding of the thermodynamics of calixester-metal complexes, it is useful to consider the enthalpic and entropic contributions to these complexation processes. These thermodynamic parameters were evaluated from the temperature-dependance of the formation constants of the calixester complexes of K$^+$ and Ag$^+$, respectively. Figures 3 and 4 show the effect of temperature on the stability constants, where log K decreases
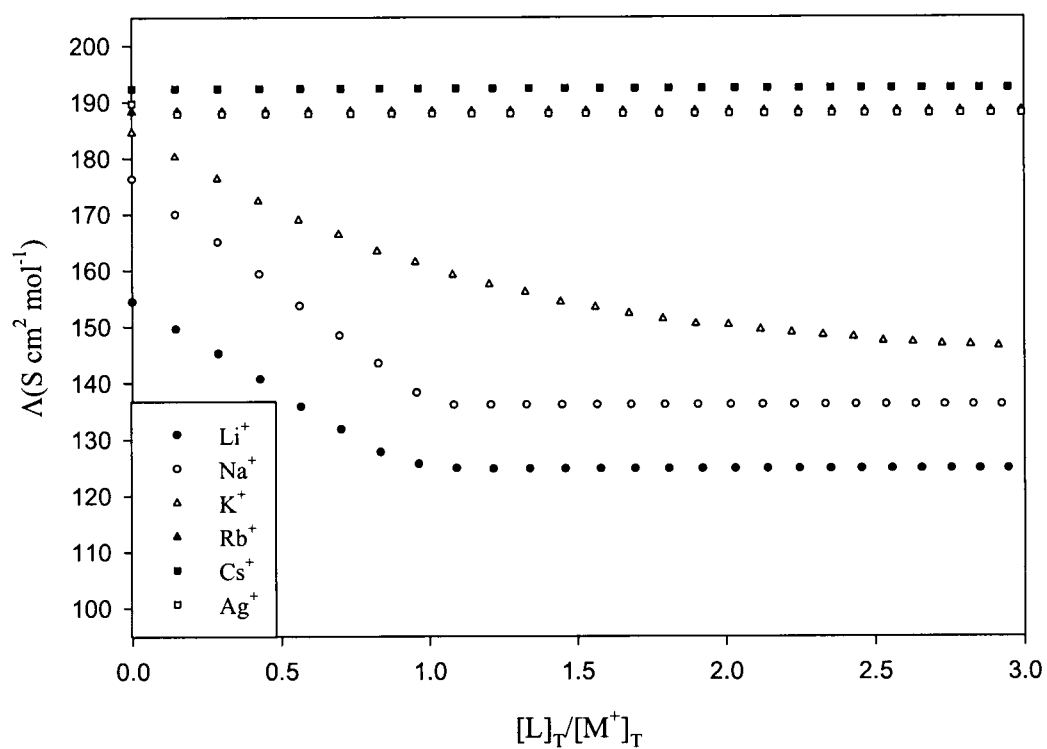
Figure 1. Molar conductance vs. $[L]_T/[M^+]_T$ for different calixester-alkali metal and silver ion complexes in acetonitrile at 25 °C.
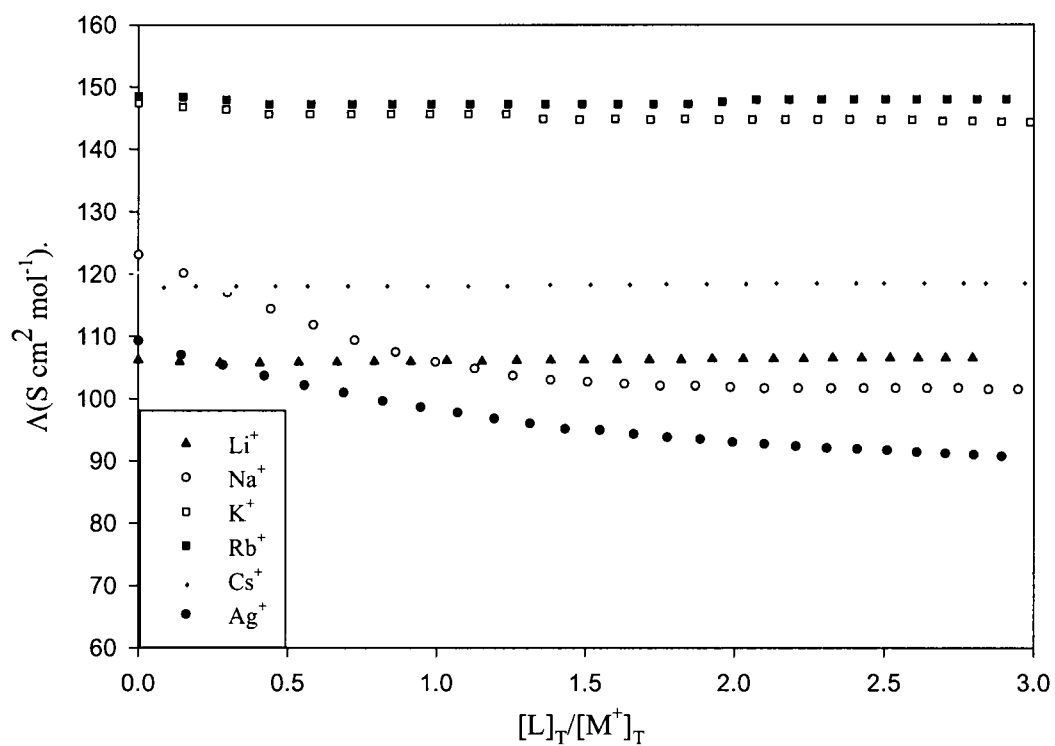

Figure 2. Molar conductance vs. $[L]_T/[M^+]_T$ for different calixester-alkali metal and silver ion complexes in methanol at 25 °C.

Table 1. Thermodynamic parameters for the complexation of alkali-metal and silver cations with ethyl *p-tert*-butylcalix[4]arenetetraethanoate in acetonitrile and methanol at 298 K

cation	log K	ΔG^0, kJ mol^{-1}	ΔH^0, kJ mol^{-1}	ΔS^0, J k^{-1} mol^{-1}
		Acetonitrile		
Li$^+$	6.25 ± 0.49	−33.21 ± 1.6	−45.83 ± 2.3	−42.36 ± 2.1
Na$^+$	7.40 ± 0.47	−41.54 ± 2.1	−56.72 ± 2.8	−50.93 ± 2.6
K$^+$	4.35 ± 0.07	−24.89 ± 1.2	−42.56 ± 2.1	−59.30 ± 2.9
		Methanol		
Na$^+$	5.48 ± 0.22	−31.41 ± 1.5	−33.56 ± 1.7	−7.23 ± 0.36
Ag$^+$	4.25 ± 0.06	−24.51 ± 1.2	−35.66 ± 1.8	−37.40 ± 1.8

Table 2. Published values of thermodynamic parameters for the complexation of alkali-metal cations with ethyl *p-tert*-butylcalix[4]arenetetraethanoate in acetonitrile and methanol at 298 K

cation	log K	ΔG^0, kJ mol^{-1}	ΔH^0, kJ mol^{-1}	ΔS^0, J k^{-1} mol^{-1}
		Acetonitrile		
Li$^+$	6.20 ± 0.05[a]; 6.4[b]	−35.39 ± 0.12[a]; −36.53[c]	−48.78[a]; −48.78[c]	−44.9[a]; −41.1[c]
Na$^+$	7.68 ± 0.08[a]; 5.8[b]	−43.81 ± 0.20[a]; −33.11[c]	−69.20 ± 0.96[a]; −61.55[c]	−85.1[a]; −95.4[c]
K$^+$	4.04 ± 0.03[a]; 4.5[b]	−23.06 ± 0.07[a]; −25.69[c]	−45.75 ± 0.45[a]; −43.85[c]	−76.1[a]; −60.9[c]
Rb$^+$	2.05 ± 0.03[a]; 1.9[b]	−11.70 ± 0.07[a]; −10.85[c]	−23.34 ± 1.36[a]; −18.67[c]	−39.0[a]; −26.2[c]
Cs$^+$	2.8[b]	−15.98[c]	−11.48[c]	+15.1[c]
		Methanol		
Li$^+$	2.6[b]	−14.84[c]	+5.05[c]	+66.7[c]
Na$^+$	5.0[b]	−28.54[c]	−45.60[c]	−57.2[c]
K$^+$	2.4[b]	−13.70[c]	−14.22[c]	−1.7[c]
Rb$^+$	3.1[b]			
Cs$^+$	2.7[b]			

[a]Reference 7.
[b]Reference 5 (standard deviations on the mean: F_{n-1} = 0.2–0.3 log unit).
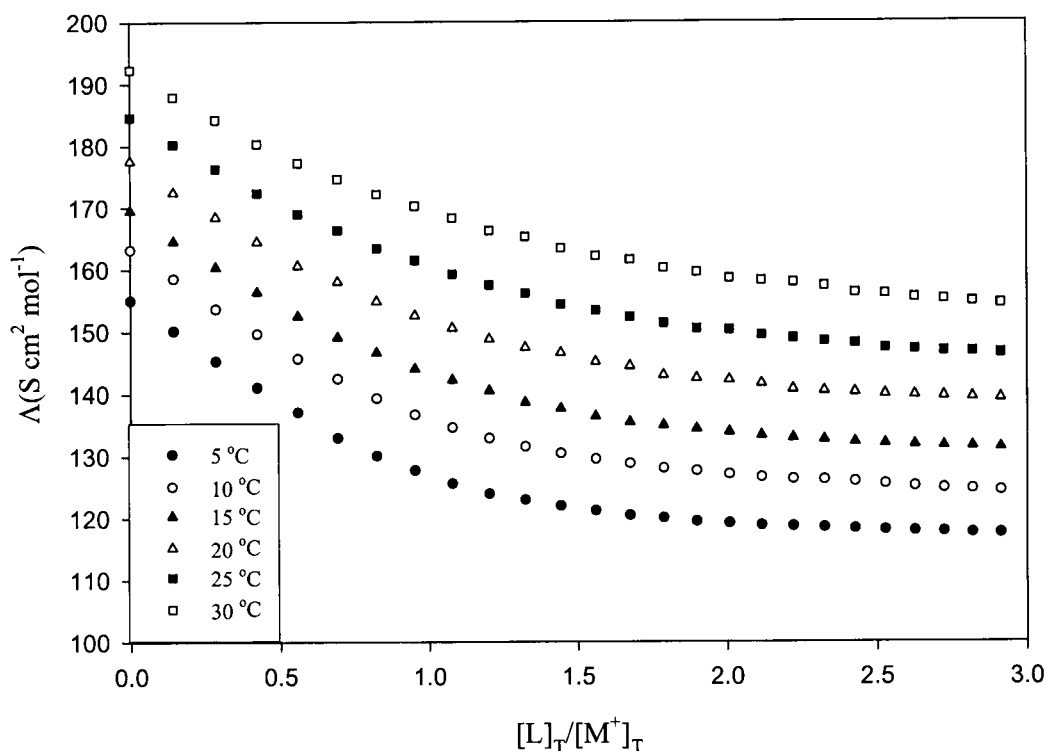[c]Reference 8.

Figure 3. Molar conductance vs. [calixester]/[K$^+$] curves, in acetonitrile at different temperatures.
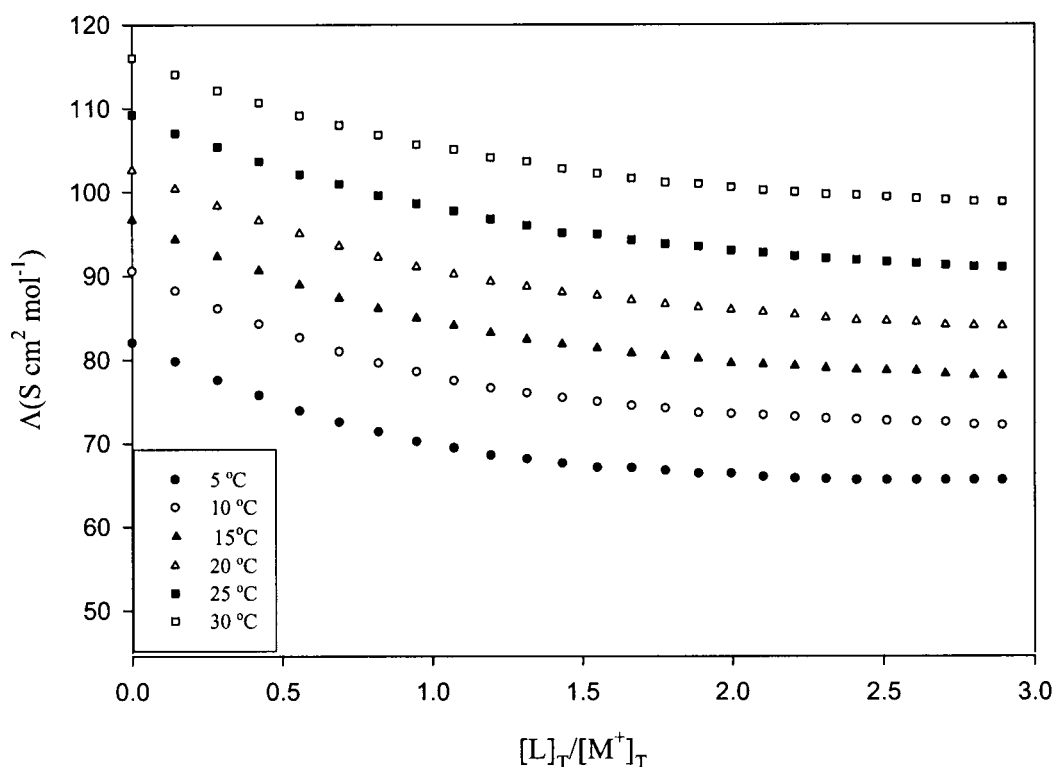
Figure 4. Molar conductance vs. [calixester]/[Ag$^+$] curves, in acetonitrile at different temperatures.

as the temperature increases indicating that the binding of metal cations with the ligand is an exothermic process. Table 1 includes the stability constants (log K), Gibbs free energies (ΔG^0), enthalpies (ΔH^0) and entropies (ΔS^0) for the complexation of alkali metal and silver cations in acetonitrile and methanol at 298 K. Table 2 shows the corresponding values reported by Arnaud-Neu *et al.* [5] and Danil de Namor *et al.* [7, 8] using UV-spectrophotometry and calorimetric methods, respectively. Comparison of the log K (and ΔG^0) and also of the ΔH^0 results obtained in the present study for the 1 : 1 complexes of Li$^+$, Na$^+$ and K$^+$ in acetonitrile with those reported by Danil de Namor *et al.* [7] shows that they are in close agreement with the complex stability sequence being: Na$^+$ > Li$^+$ > K$^+$. Our results however have larger standard deviations. Neither our results nor those of Danil de Namor *et al.* [7] are in agreement with those obtained by Arnaud-Neu *et al.* [5] who reported only log K values, with the complex stability sequence being Li$^+$ > Na$^+$ > K$^+$. We were unable, however, to compare our results on the complexation of calixester with Ag$^+$ in methanol due to the lack of any related published data. It is worth noting that the ΔS^0 values obtained for the Na$^+$-calixester complex in methanol is higher than in acetonitrile, implying that Na$^+$ is more extensively solvated in methanol than in acetonitrile. Consequently, when the cation loses most of its solvation shell upon complexation with the calixester, the entropy change becomes more positive. This result is consistent with that obtained by the calorimetric studies [8].

In conclusion, it can be seen that the conductometric technique can be considered as a simple and reliable method which can complement other methods for the thermody-

namic studies of the complexation reactions of calixarene derivatives with alkali metal cations. A follow-up study using ^{1}H NMR [13] will be reported in due course.

Acknowledgements

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