SUPRAMOLECULAR CHEMISTRY: THERMODYNAMIC ASPECTS

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Resumen

Un sumario de las actividades del Laboratorio de Termoquímica en el campo de la Química Supramolecular con particular referencia a la termodinámica de procesos de formación de complejos con cationes en diferentes medios es presentado.
La relevancia de estos estudios sobre diferentes areas es demostrada con ejemplos representativos.

Palabras clave: Química supramolecular, termodinámica, procesos en solución.

Abstract

A summary of the activities of the Thermochemistry Laboratory in the field of Supramolecular Chemistry with particular reference to the thermodynamics of cation complexation processes in different media is given.
The relevance of these studies in several areas is demonstrated with representative examples.

Key words: supramolecular chemistry, thermodynamics, solution processes.

Introduction

Supramolecular Chemistry as stated by Lehn is *'the chemistry of the intermolecular bond, being concerned with the structures and functions of chemical species of greater complexity than the molecules themselves'* [1, 2]. However the notable ex-

pansion of this area in the last three and half decades has required a continuous revision of the definition of Supramolecular Chemistry as given by Lehn [3]. Through this paper macrocycles are considered as molecules characterised by holes or cavities and these are referred as hosts while guests are ionic and neutral species. The hosts are neutral and therefore these are non-electrolytes in solution.

The roots of Supramolecular Chemistry date back to the end of the 19th century, beginning of the twentieth century

when fundamental concepts such as coordination, recognition and fixation were introduced by Werner [4], Fischer [5] and Ehrlich [6] respectively. The beginning of Supramolecular Chemistry took place with the realisation that alkali-metal cations were selectively complexed by natural [7] and synthetic [8] macrocyclic ligands. It was however the discovery of synthetic macrocycles in the late sixties, early seventies which triggered the extraordinary development of this area of research. Since then Supramolecular Chemistry has developed to an extent that it is considered today as one of the most active areas of research. Thus the pioneering work by Pedersen (crown ethers) [9], Lehn (cryptands) [1] and Cram (spherands) [10] was acknowledged with the Nobel Prize Award in 1987.

Thermodynamic aspects of Supramolecular Chemistry. Objectives

The synthetic developments in the field of Supramolecular Chemistry have been greatly motivated by the interest in finding ligands which are able to interact selectively with one species relative to another in a given medium. Quantitative assessments of selectivity have been made and these are based on the stability constant ratios of their complexes in a given solvent and at a given temperature. However such assessments requires accurate thermodynamic data, the derivation of which is by no means trivial particularly for processes involving ionic species. Thus for a given cation, M^{n+}, and a given ligand, L, to give a 1:1 metal-ion complex, $M^{n+}L$ in a given solvent, s, the complexation process may be represented by eq. 1

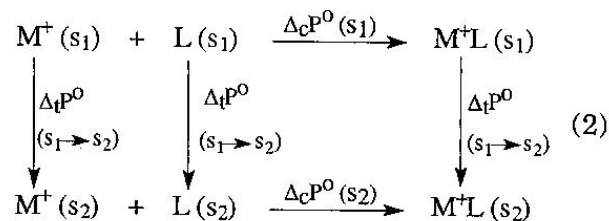$$M^{n+} (s) + L (s) \rightarrow M^{n+}L (s) \qquad (1)$$
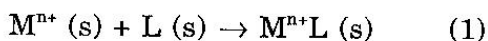
The formulation of eq. 1 [11] not only requires an investigation of the composition of the complex but also requires knowledge regarding whether the free and the complex electrolytes are predominantly in their ionic forms in solution as demanded by eq. 1. Although information regarding the behaviour of the free metal cation salts

in different solvents is available in the literature [12, 13], this has been often ignored. In contrast, knowledge regarding the solution properties of macrocycle based electrolytes in different solvents is very limited [11]. Therefore it can be safely stated that a great deal of thermodynamic data involving ionic species are based on the assumption that if the free metal-ion salt if fully dissociated in a given medium, the same behaviour is expected for the complex metal salt since this is formed by a large cation and as such is unlikely to interact with the anion. Although this may be often the case, the validity of such assumption needs to be experimentally verified in order to fulfil the requirements of eq.1 which explicitly demands that the free and the complex cations are predominantly as ionic species in solution. These have been recently discussed by the author [11].

Distinctive features of our Laboratory relative to others worldwide in this field involve investigations on

i) the medium effect on the complexation process (eq. 1) by the thermodynamic characterisation of the reactants (the metal cations, M^+ and the ligand, L) and the product (metal-ion complex, M^+L) participating in the binding process [14]. This statement is illustrated in the following thermodynamic cycle introduced by us in 1977 [15]. In this case, univalent cations are considered
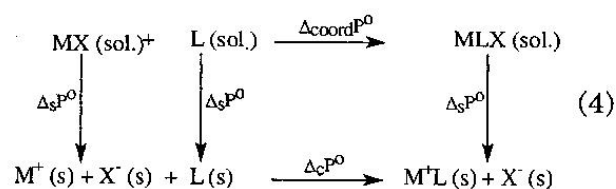
$$
\begin{array}{ccccc}
M^+ (s_1) & + & L (s_1) & \xrightarrow{\Delta_c P^O (s_1)} & M^+L (s_1) \\
\Big\downarrow \Delta_t P^O & & \Big\downarrow \Delta_t P^O & & \Big\downarrow \Delta_t P^O \\
(s_1 \to s_2) & & (s_1 \to s_2) & & (s_1 \to s_2) \\
M^+ (s_2) & + & L (s_2) & \xrightarrow{\Delta_c P^O (s_2)} & M^+L (s_2)
\end{array}
\qquad (2)
$$

In eq. 2; $\Delta_c P^\circ$ ($P^\circ = G^\circ$, H°, S°) are the thermodynamic parameters of complexation in a reference solvent (s_1) and other solvent (s_2) where $\Delta_t P^\circ$ are the corresponding data for the transfer of the free cation, the ligand and the metal-ion complex from s_1 to s_2.

Accordingly, the relationship between complexation and transfer thermodynamics is shown in eq. 3

$$\Delta_c P^\circ (s_1) - \Delta_c P^\circ (s_2) =$$
$$\Delta_t P^\circ (M^+) (s_1 \to s_2) + \Delta_t P^\circ (L) (s_1 \to s_2) -$$
$$\Delta_t P^\circ (M^+L) (s_1 \to s_2) \qquad (3)$$

This equation unambiguously demonstrates that the medium effect on the complexation process is determined by the differences in solvation of the reactants and the product [11].

ii) the coordination process ($\Delta_{coord} P^\circ$) [16,17] referred to the reactants and the product in their pure physical state as shown in eq. 4

MX (sol.)$^+$ L (sol.) $\xrightarrow{\Delta_{coord}P^0}$ MLX (sol.)

$\Delta_s P^0$ $\Delta_s P^0$ $\Delta_s P^0$ (4)

M$^+$ (s) + X$^-$ (s) + L(s) $\xrightarrow{\Delta_c P^0}$ M$^+$L (s) + X$^-$(s)

In eq. 4; $\Delta_s P^\circ$ are the thermodynamics of solution of the free (MX) and the complex metal-ion salt (MLX) and the ligand (L).

Coordination data have proved to be a useful mean to check the accuracy of the thermodynamic values involved in the cycle since for a given system, $\Delta_{coord} P^\circ$ should be the same (within the experimental error) independently of the solvent from which these are derived. In addition, $\Delta_{coord} P^\circ$ is dependent on the counter-ion constituent of the salt and as such it has been possible to evaluate the anion effect on the process shown in eq. 4 [17]. Another important point to stress is that related with the difficulties associated with the determination of the thermodynamics of cation complexation process in low dielectric media. In these media, extensive ion-pair formation takes place and therefore additional processes besides complexation are involved. Therefore the data obtained are not referred to a well defined process. In this situation, combination of coordination and solution data can lead to the calculation of complexation data in low dielectric medium.

iii) the factors which contribute to the selective extraction of metal cations from aqueous solutions to an organic phase in the presence and in the absence of the macrocycle [18, 19].

Some of the areas covered are now briefly described. These include synthetic macrocycles (cryptands, crown ethers, calixarenes and their lower rim derivatives) and natural macrocycles such as cyclodextrines. Representative ligands are shown in Fig. 1. These contributions will be described in chronological order.

1. Thermodynamics of cryptands and metal cations

Cryptands are macrobicyclic ligands characterised by the presence of a three dimensional hydrophilic cavity lined with oxygens and nitrogen donor atoms as shown in Fig. 1. The size of this cavity is determined by the size of the bridges connecting the nitrogen atoms. It is the presence of the intramolecular cavity, the distinctive feature of cryptands relative to crown ethers (holes) and consequently, cation complexes with the former ligand (cryptand) are much more stable than cation coronand complexes [20, 21].

The availability of complexation Gibbs energies [22-25], enthalpies and entropies [26-31] for univalent cations and cryptand 222 (222) in a wide variety of dipolar aprotic solvents including those characterised by their low mutual solubility in water (which can be therefore used for the direct partitioning of electrolytes) led to the calculation of the entropy, $\Delta_1 S^\circ$ for the process described by eq. 5 [32-34]
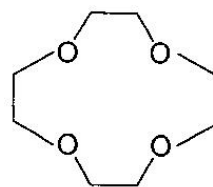
$$M^+ (gas) + 222 (s) \xrightarrow{\Delta_1 S^0} M^+ 222 (s) \quad (5)$$
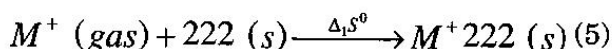
This is essentially the transfer of the cation from the gas phase to the solvated cryptand to form the solvated cryptate, M$^+$222 in the appropriate solvent. The $\Delta_1 S^\circ$ values in the various dipolar aprotic solvents (N,N-dimethylformamide, dimethylsulphoxide, propylene carbonate, acetonitrile, nitromethane) were found to be
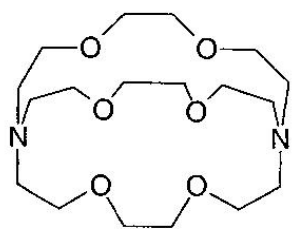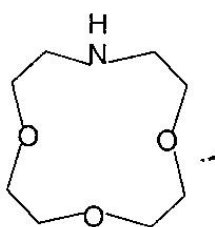
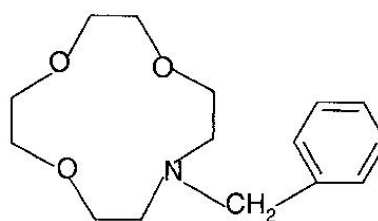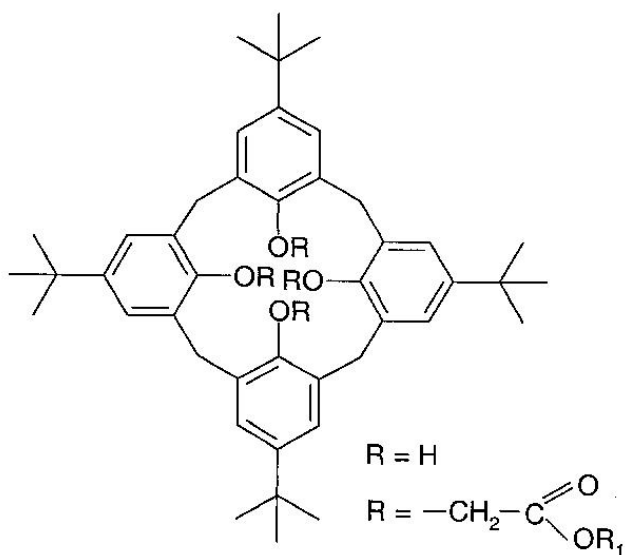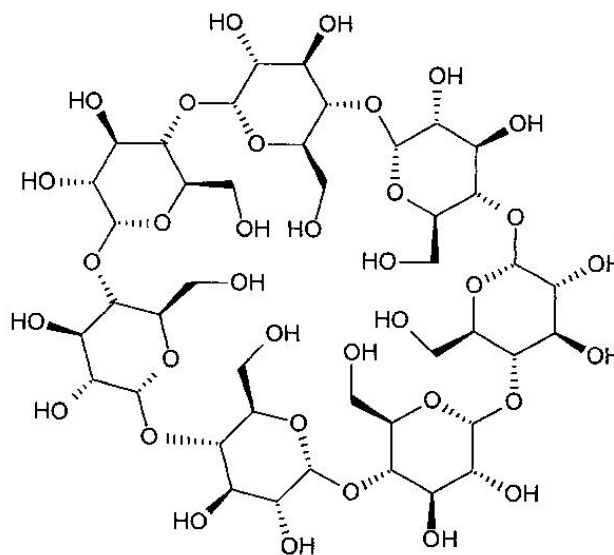Cryptand 222
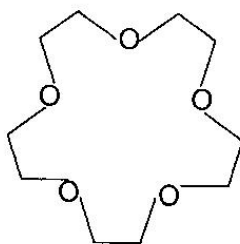
15-crown-5

12-crown-4

1-aza-12-crown-4

1-benzyl-1-aza-12-crown-4

$R = H$

$R = -CH_2-C{\overset{O}{\underset{OR_1}{}}}$

Lower Rim *p-tert*-butylcalix(4)arene
and its derivative

β-Cyclodextrin

Fig. 1. Representative Macrocycles.

constant. Since $\Delta_t S^\circ$ may be equalled to the contribution of both, the solvation entropy of the cation, $\Delta_{solv} S^\circ$ and the entropy of complexation of the same cation with the ligand, $\Delta_c S^\circ$ (eq. 6),

$$\Delta_t S^\circ = \Delta_{solv} S^\circ + \Delta_c S^\circ \qquad (6)$$

It follows that the following correlation holds,

$$\Delta_c S^\circ = \text{constant} - \Delta_{solv} S^\circ \qquad (7)$$

Thus, a linear correlation between the $\Delta_c S^\circ$ values and $\Delta_{solv} S^\circ$ in non-aqueous solvents (dipolar aprotic solvents) was shown (Fig. 2).

The important points derived from this correlation are summarised as follows.

(i) Fig. 2 essentially shows agreement between two alternative extrather-modynamic (the $Ph_4 As Ph_4 B$ convention [35] and the cryptand convention [36]) means of estimating entropies of transfer of ions from water to different solvents. The validity of eq.7 for non-aqueous solvents, mainly dipolar aprotic solvents, allows the calculation of entropies of solvation for alkali metal and silver cations (and therefore $\Delta_t S^\circ$) from entropies of complexation data between these cations and cryptand 222 and vice versa.

(ii) The correlation found was not so expected, since there is information in the literature on the formation of 'exclusive' complexes between some cryptands and some metal cations [37]. It is expected that cation desolvation occurs upon complexation. However there are degrees of desolvation. These findings unambiguously demonstrated that these univalent cations are almost fully desolvated when entering the cavity of the macrocycle and it is on this basis that a linear correlation is found. This was indeed the first quantitative correlation ever shown between entropies of solvation and entropies of complexation in these systems.

Further work carried out with cryptand 222 and bivalent and trivalent cations led us to the conclusion that the effective shielding effect of the ligand for univalent cations in dipolar aprotic media appears to weaken considerably for bivalent (Ba^{2+}) cations and breaks completely for some of the lanthanide cations [38]. As far as bivalent and trivalent cations are concerned further experimental work needs to be carried out in order to confirm these findings.

2. Effect of cyclodextrins on the transfer of drugs (substituted sulfonamides) from water to chloroform [39]

As shown in Fig. 1 cyclodextrins are macrocycles constituted by glucose units and characterised by a hydrophobic cavity which is able to form inclusion complexes with a large variety of guests. Particular emphasis has been placed on the applications of cyclodextrins in the pharmaceuti-
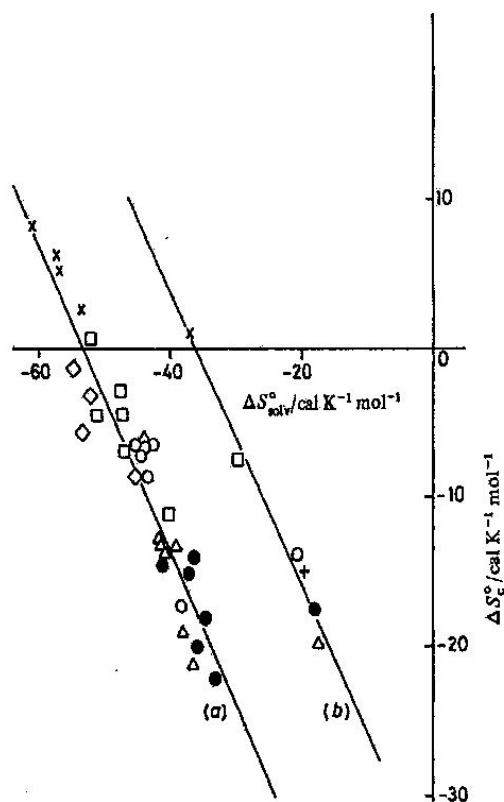

Fig. 2 Linear correlation between entropies of complexation of cryptand 222 with metal ions end entropies of solvation of these ions (a) in non-aqueous solvents and (b) in water. ×, Li+; □, Na+; O, K+; △, Rb+; •, Cs+; □, Ag+; +, Tl+. 1 cal = 4.184 J
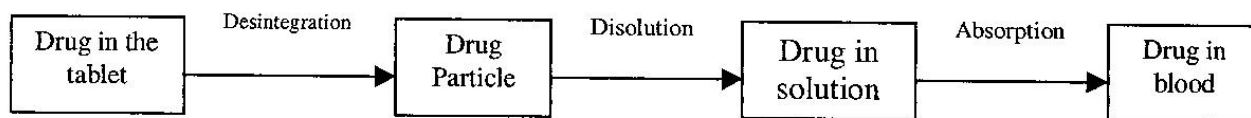
Fig. 3. Drug Bioavailability

cal industry [40]. An interesting aspect is that related to the solubility enhancement observed for slightly soluble drugs by the addition of cyclodextrins (nontoxic macrocycles). If the different steps involved in the bioavailability of a drug are considered (see Fig. 3), an important one is that related to the transfer across the absorption barrier. The effect of cyclodextrins on the transfer of N'-substituted sulfonamides (S) was investigated in the water-chloroform solvent system. It is only on the basis of its low dielectric constant that chloroform has been often used as a representative of the membrane. The fact that (i) no complexation appears to take place between cyclodextrin and these drugs in chloroform and (ii) the solubility of the complex (S-CD) in chloroform was approximately the same as that of the drug in this solvent, it was concluded that the overall transfer process involves two steps

(i) the dissociation of the sulfonamide – cyclodextrin complex in water

$$S\text{-}CD_{(H_2O)} \xrightarrow{K_d} S_{(H_2O)} + CD_{(H_2O)} \qquad (8)$$
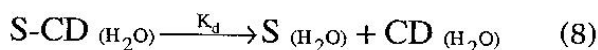
(ii) the partition of the drug from water to chloroform.

$$S_{(H_2O)} \xrightarrow{K_p} S_{(Cl_3CH)} \qquad (9)$$
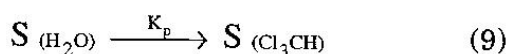
Therefore $K_{ex} = 1/K_s \times K_p$. Since $K_p > K_{ex}$ it was then concluded that from the thermodynamic view point, sulfonamides are more favourably transferred in the absence of β-cyclodextrin than in its presence. Indeed the advantage of increasing the drug solubility and the disadvantage of decreasing the absorption of the drug by the addition of cyclodextrins needs to be carefully considered to achieve the bioavailability of the drug.

Other studies involving cyclodextrins are these concerning the interaction of these ligands with anions constituents of antigenic determinants. This work demonstrated the formation of inclusion or axial type of complexes in water while equatorial or lid type of complexes were suggested in N, N-dimethylformamide [41].

3. Thermodynamic, structural and conductimetric studies of electrolytes relevant in lithium battery technology

Considerable attention has been focused in recent years on the solution chemistry of lithium electrolytes due to their wide range of biological [42] and industrial applications [43, 44]. Major international efforts are competitive in the development of rechargeable lithium batteries particularly those based on the 'rocking chair' concept [45-49] as an environmentally friendly and safer alternative to metallic lithium and the more conventional lead acid and nickel cadmium systems. For either system considered (metallic lithium or the rocking chair system) a major problem in their development is that related to electrolyte solutions [50]. For metallic lithium battery systems, the use of an inert solvent (non-protic) of medium or low permittivity brings about problems related to low solubilities and low conductivities due to the formation of ion pairs (nonconducting species) and in some cases dimers. Rocking chair systems are concentration cells in which lithium is the only ion transported and where metallic lithium is replaced by lithiated carbon or the intercalate $LiTiS_2$. Although similar problems to

those found in metallic lithium battery are found with these systems, an additional advantage of the latter is that more aggressive aprotic solvents can be used. In an attempt to overcome the problems associated with electrolyte solutions, macrocycles were explored. Indeed the large variety of neutral macrocycles produced in the last thirty five years [1, 2], many of them able to interact with alkali-metal cations including lithium has opened the possibility of exploring the use of these new electrolytes in the development of lithium batteries. However the isolation of the metal-ion complex salts requires information regarding the stability of complex formation in the appropriate medium. Therefore we proceeded with the thermodynamic characterisation of these systems. The crown ethers ligands selected for these studies were those shown in Fig. 1 and the electrolytes considered were those containing cations of relatively large size (lithium coronand) and highly polarisable anions such as hexafluoroarsenate, AsF_6^-, tetrafluoroborate, BF_4^- and trifluoromethanesulfonate, $CF_3SO_3^-$. Based on this information which showed that the complexation of lithium and crown ethers in solvents relevant to battery technology was strong enough (stability constant of ~ 10^4) to be isolated, we proceeded with the structural (^{1}H and ^{13}C NMR in CD_3CN) and thermochemical characterisation of these new electrolytes [51, 52].

Table 1 lists representative data for the standard enthalpies of solution of lithium and lithium coronand salts in acetonitrile and in propylene carbonate at 298.15 K. Standard enthalpies of solution result from the contribution of the solvation process (exothermic) and that of the crystal lattice (endothermic). As far as lattice enthalpies for ionic solids containing the same anion are concerned, it is expected that as the size of the cation increases (in going from lithium to lithium coronand) the amount of heat supplied to break up the solid decreases. On this basis and provided that in the solid no specific interactions occurs between the ligand and the anion, the heat associated with the crystal lattice process for the uncomplex lithium salt is expected to be considerably larger (more endothermic) than that of the lithium coronand salt. The data in Table I reflect that while the dissolution of the uncomplexed salt takes place with a release of energy (exothermic process), the reverse is true for the lithium coronand salts. We found these findings striking since they provided a strong indication that the complexed salts are much less solvated that the free lithium salt and therefore the implications of these results on the conductivity of these electrolytes (an important aspect to consider in the development of lithium batteries) could be foreseen. Thus Fig. 4 shows the conductimetric titration

Table I. Standard Enthalpies of Solution of Lithium and Lithium Coronand Salts in Acetonitrile and Propylene Carbonate at 298.15 K

Compound	$-\Delta_s H°$ / kJ mol^{-1} [a]	
	Acetonitrile	Propylene Carbonate
LiCF$_3$SO$_3$	-15.59	-12.50
[Li 15C5] CF$_3$SO$_3$	9.17	13.34
[Li 12C4] CF$_3$SO$_3$	15.69	20.74
[Li 1A12C4] CF$_3$SO$_3$	14.66	18.74
[Li 1BA12C4] CF$_3$SO$_3$	18.96	22.50

Abbreviations; 15C5: 15-crown-5; 12C4: 12-crown-4; 1A12C4: 1-aza-12-crown-4; 1BA12C4: 1-benzyl-1-aza-12-crown-4   [a] From refs. 51 and 52.
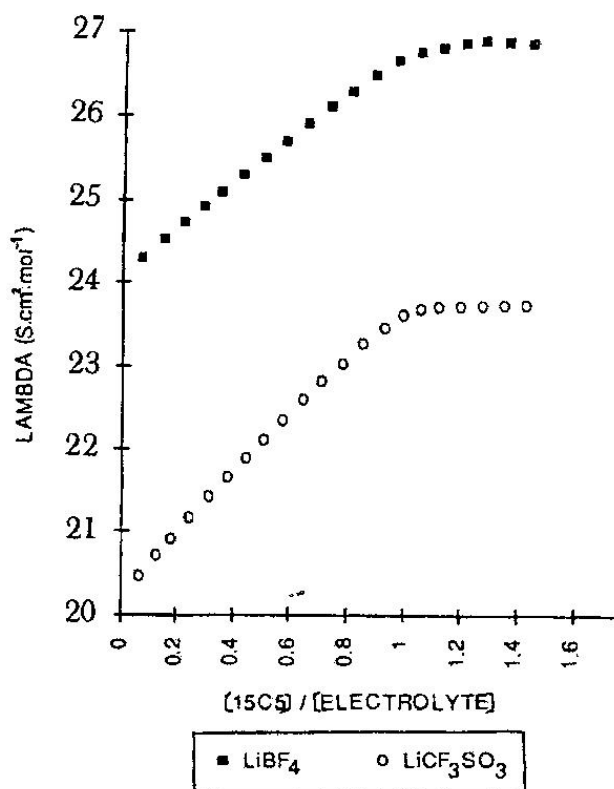
Fig. 4. Conductimetric curves for the titration of lithium (trifluoroborate and trifluoromethane-sulphonate as counter-ions) and 15 crown 5 in propylene carbonate at 298.15 K.

curves for the titration of $LiCF_3SO_3$ and $LiBF_4$ with 15 crown 5 in propylene carbonate. These results unambiguously demonstrate that as the equilibrium position is shifted by the addition of the macrocycle as to favour the formation of the less solvated electrolyte (lithium coronand) a considerable increase in conductance is observed as predicted from thermochemical data. Within this context, it is important to emphasise the relevance of fundamental thermodynamics in the selection of macrocyclic based electrolytes for use in battery technology.

4. Calixarene Chemistry.

Calixarenes (products of the base-induced condensation of *p*-substituted phenols and formaldehyde) and their derivatives have received considerable attention

in recent years to an extent that several books and review articles have been published [53-56]. The current widespread interest in calixarenes is due to several factors. Among these are the remarkably simple way used for the synthesis of the parent compounds and the possibility of obtaining a large variety of derivatives by lower and upper rim functionalisation of the parent compounds. Among these macrocycles, calix(4)arenes are of particular interest. Indeed, these ligands in their 'cone' conformation are characterised by the presence of two cavities: one hydrophobic situated between the benzene rings and able to interact with neutral species (Fig. 5) while the other is hydrophilic and able to interact with ionic species. It is indeed the presence of hydrophilic and hydrophobic regions which makes these macrocycles suitable hosts for neutral and ionic species. The thermodynamics of calixarene chemistry has been reviewed [11] by us in 1998.

Calixarenes have found a wide range of applications. A good account on the industrial uses of these macrocycles prior to 1992 has been given by Perrin and Harris [57].

Our contribution in this area involves

i) The interaction of parent calixarenes with amines [58-60].

ii) Detailed thermodynamic information on calixarene ester derivatives and alkali- metal cations [16-18, 61].

iii) Synthesis, characterisation (^{1}H NMR and X ray diffraction studies) and complexing properties of lower rim calixarene derivatives containing aliphatic and alicyclic amines as pendent arms [62, 63] (Fig. 6), mixed functional groups (Fig. 7) [64-66] and more recently partially functionalised calix(4)arenes (Fig. 8) [67].

iv) Solution thermodynamics of pyridinocalix(4)arenes (Fig. 9) and monovalent cations [68,69] The crystal structure of the monoacetonitrile sodium complex of **9a** is shown in Fig. 10.

v) Complexation thermodynamics of calixarene derivatives and lanthanide cations [70]

Hydrophobic Region
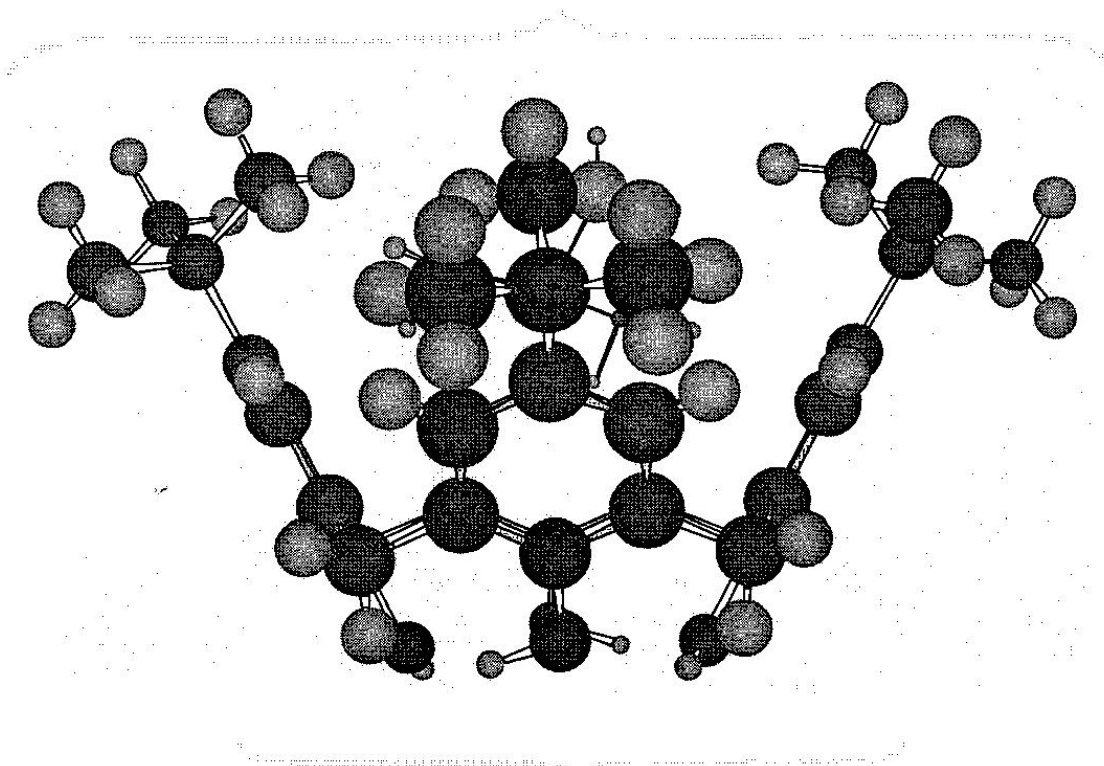
Hydrophilic Region

Fig. 5. Calixarenes.

vi) Quantitative assessments of the individual processes involved in the extraction of alkali-metal and silver cations from aqueous solution to the non-aqueous phase in the presence and absence of calix(4)arene derivatives [18, 19].

As far as the thermodynamics of these systems are concerned, there are misleading statements in the literature either because the medium selected for these measurements was not suitable or unawareness of the limitations of the methodology used in the derivation of thermodynamic data and these have been extensively discussed by our group [11].

Eq. 3 has been extensively used to determine the factors which contribute to the selective behaviour of calixarene deriva-tives for metal cations. It has been demonstrated that in analysing the medium effect on the interaction of these macrocycles with metal cations, the differences in solvation of the reactants and the product must be carefully considered. This is exemplified by inserting the appropriate thermodynamic quantities in eq. 2. In doing so, two solvent systems (MeCN – MeOH and MeCN – DMF) are considered as shown below. For the interpretation of the data, the relationship shown in eq. 3 is considered. This equation implies that the optimum conditions for higher stability in s_2 ($\Delta_c G°$ more negative) relative to s_1 requires a solvent (s_2) which is a better solvator for the metal-ion complex (negative $\Delta_t G°$ from s_1 to s_2) and a poorer solvator for the ligand and the free

6a

6b

6c

6d

6e

6f
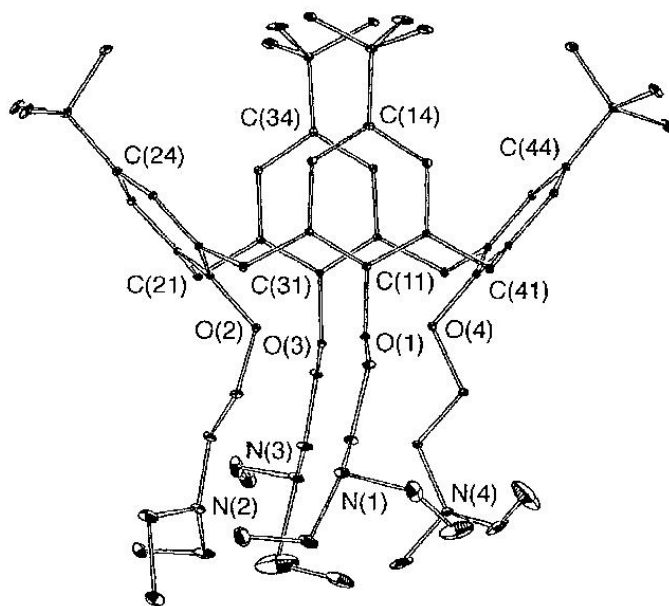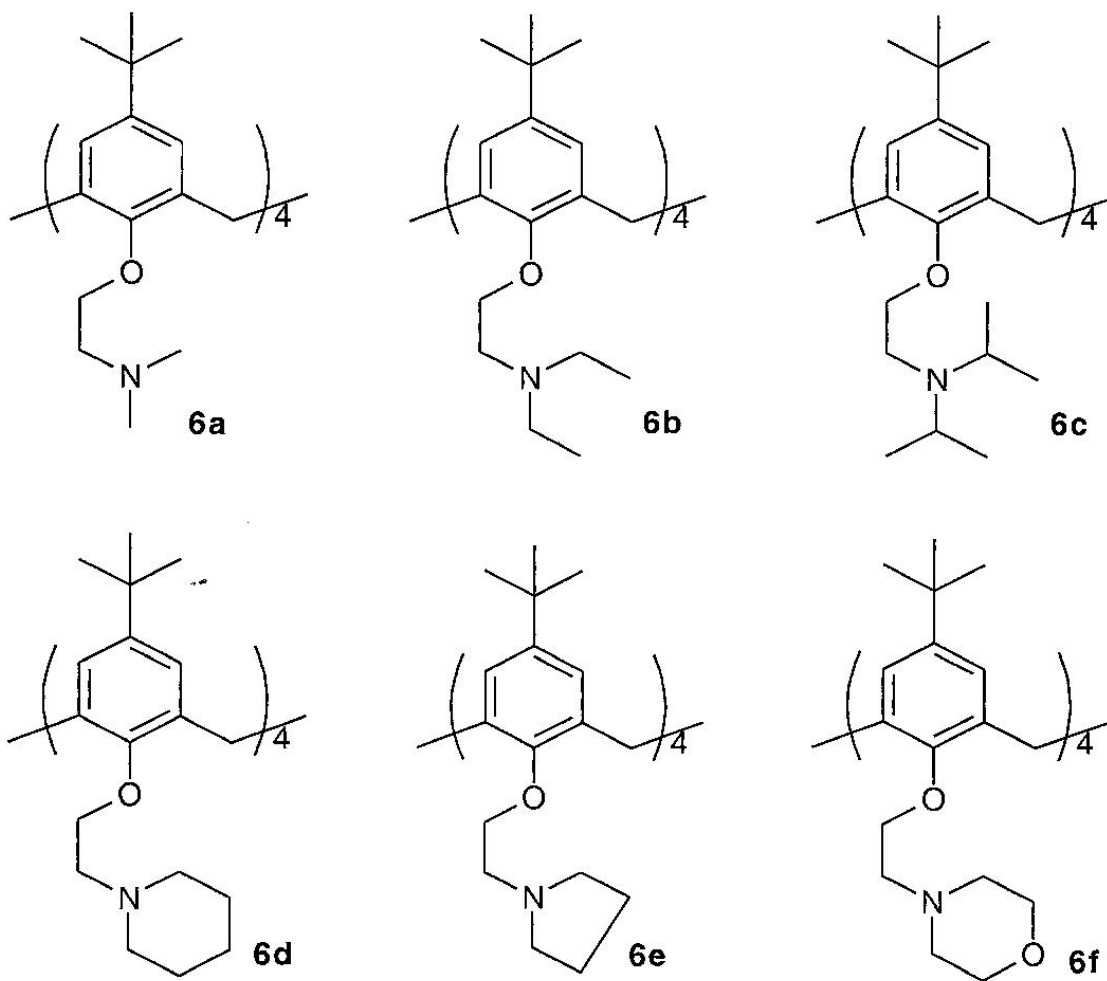
Crystal Structure of **6b** (Ref.63)

Fig. 6. Calix(4)arene Derivatives Containing Aliphatic and Alicyclic Amines as Functionals Groups.

7a   **7b**   **7c**

7d   **7e**   **7f**

7g   **7h**
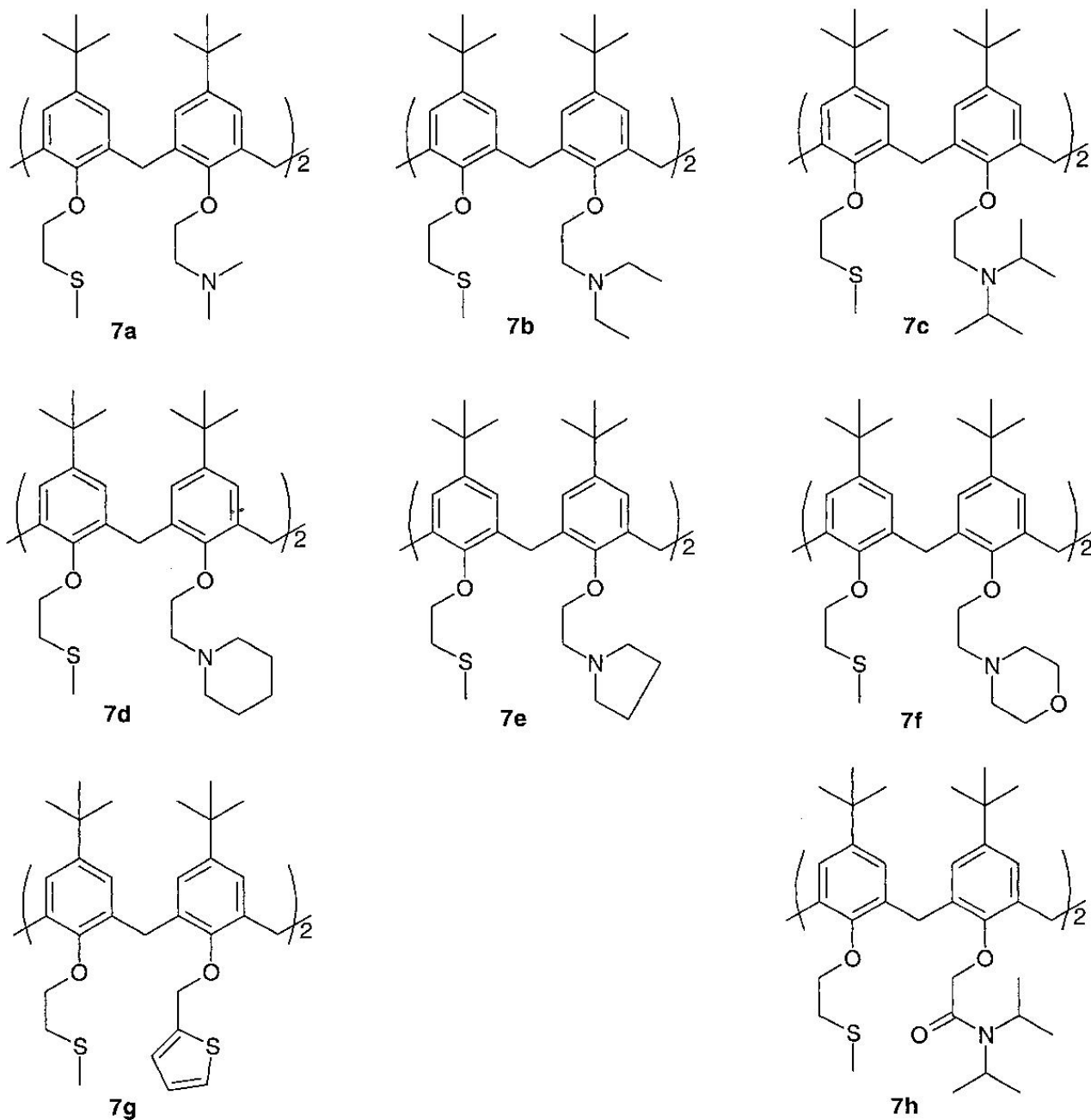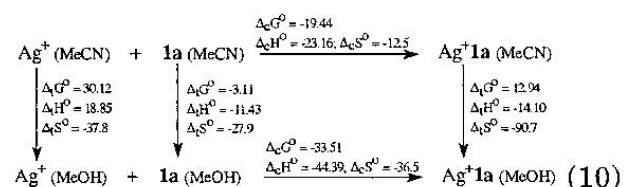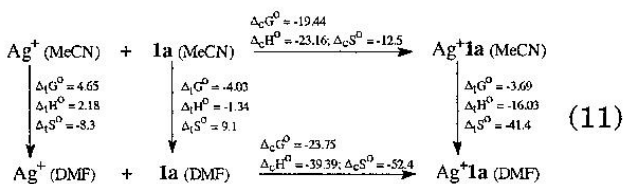
Fig. 7 Calix(4)arene Derivatives with Mixed Functional Groups.

metal cation [positive $\Delta_t G°$ ($s_1 \rightarrow s_2$)] than the reference solvent (s_1)

MeCN - MeOH[a]

$$Ag^+ \text{(MeCN)} + \mathbf{1a} \text{(MeCN)} \xrightarrow[\Delta_c H° = -23.16; \Delta_c S° = -12.5]{\Delta_c G° = -19.44} Ag^+\mathbf{1a} \text{(MeCN)}$$

$\Delta_t G° = 30.12$ $\Delta_t G° = -3.11$ $\Delta_t G° = 12.94$
$\Delta_t H° = 18.85$ $\Delta_t H° = -11.43$ $\Delta_t H° = -14.10$
$\Delta_t S° = -37.8$ $\Delta_t S° = -27.9$ $\Delta_t S° = -90.7$

$$Ag^+ \text{(MeOH)} + \mathbf{1a} \text{(MeOH)} \xrightarrow[\Delta_c H° = -44.39; \Delta_c S° = -36.5]{\Delta_c G° = -33.51} Ag^+\mathbf{1a} \text{(MeOH)} \quad (10)$$

MeCN - DMF[a]

$$Ag^+ \text{(MeCN)} + \mathbf{1a} \text{(MeCN)} \xrightarrow[\Delta_c H° = -23.16; \Delta_c S° = -12.5]{\Delta_c G° = -19.44} Ag^+\mathbf{1a} \text{(MeCN)}$$

$\Delta_t G° = 4.65$ $\Delta_t G° = -4.03$ $\Delta_t G° = -3.69$
$\Delta_t H° = 2.18$ $\Delta_t H° = -1.34$ $\Delta_t H° = -16.03$
$\Delta_t S° = -8.3$ $\Delta_t S° = 9.1$ $\Delta_t S° = -41.4$

$$Ag^+ \text{(DMF)} + \mathbf{1a} \text{(DMF)} \xrightarrow[\Delta_c H° = -39.39; \Delta_c S° = -52.4]{\Delta_c G° = -23.75} Ag^+\mathbf{1a} \text{(DMF)} \quad (11)$$

In eqs 10 and 11, $\Delta G°$ and $\Delta H°$ values are given in kJ mol^{-1} while the units of $\Delta S°$ values are JK^{-1} mol^{-1}
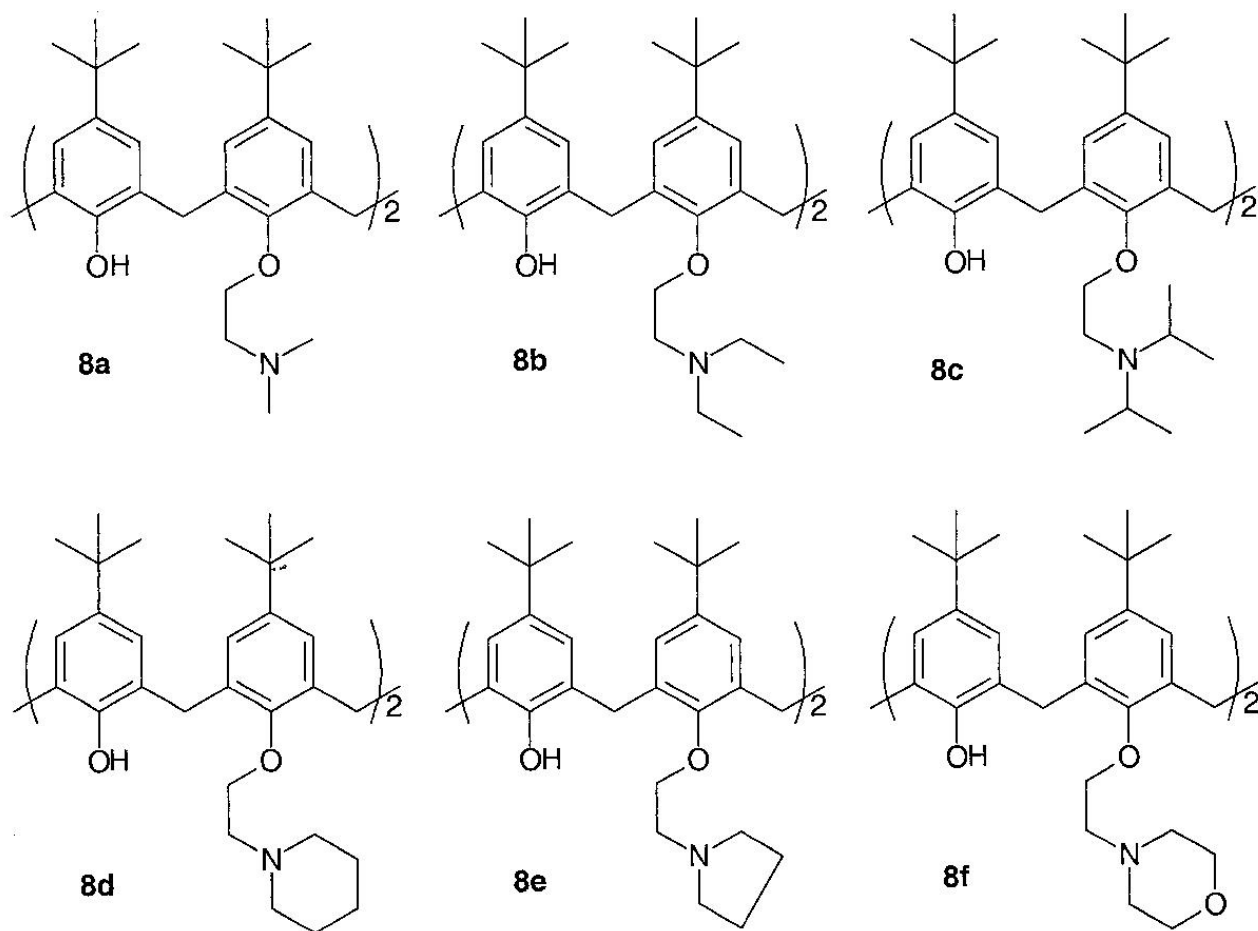
- 79 -

Fig. 8. Partially Functionalised Calix(4)arene Derivatives.
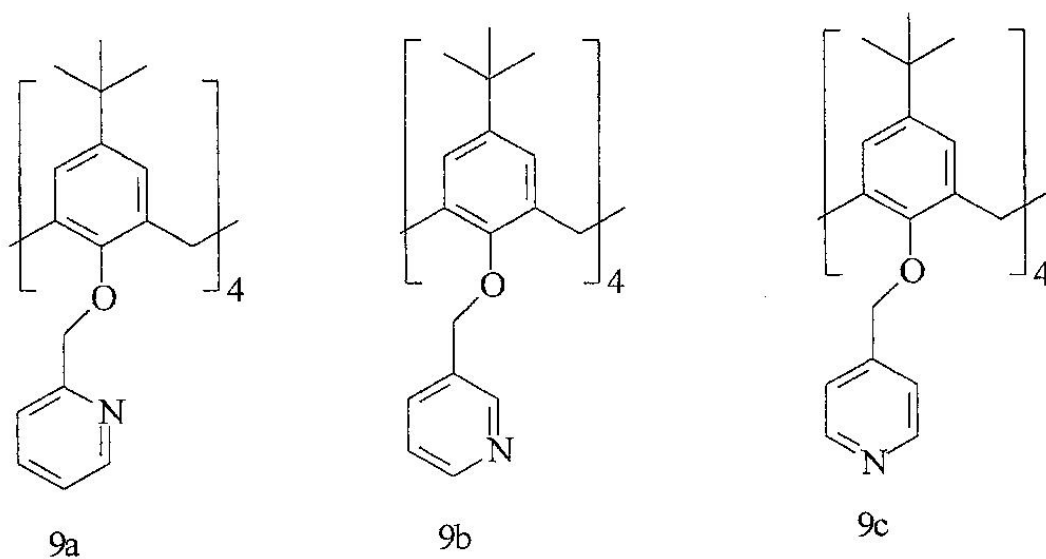

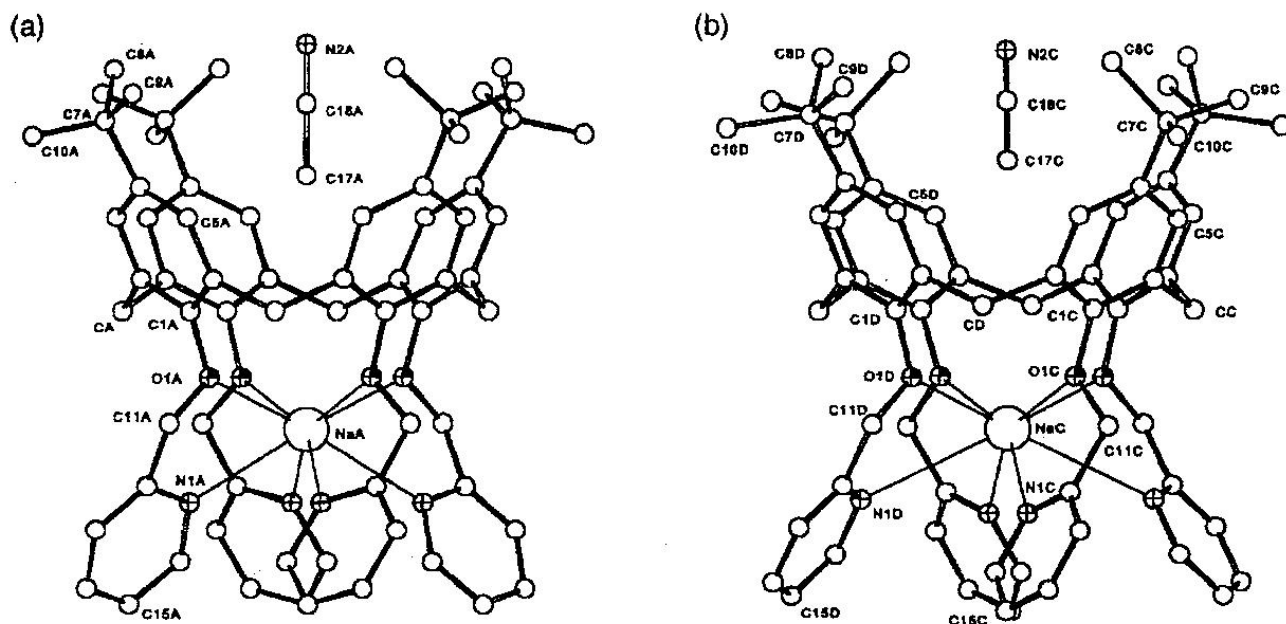Fig. 9. Pyridinocalix(4)arene Derivatives.

(a)

(b)

Fig. 10. (a) Side views of one of two independent sodium and acetonitrile complex of 9a sited on a fourfold crystallographic axis and (b) Side view of the complex sited on a twofold axis [Ref. 68].

As far as the MeCN – MeOH solvent system is concerned, in terms of Gibbs energies, the higher stability in MeOH relative to MeCN is controlled by the lower solvation of the free cation in the former solvent relative to the latter. Taking into account eq. 3 it follows that the higher solvation of the ligand and the lower solvation of the metal-ion complex in MeOH with respect to acetonitrile contribute unfavourably to the stability of the complex in this solvent relative to acetonitrile.

In terms of enthalpies, the transfer value for the metal-ion complex from MeCN to MeOH is similar to that of the ligand. These findings do not necessarily imply that the thermochemical origen of solvent-metal ion complex interactions is necessarily similar to that of the ligand in these media. However, the data unambiguously demonstrates that the higher solvation of the metal-ion complex in MeCN relative to MeOH, which contrasts with that of the ligand (the opposite is observed) is entirely due to the greater loss of entropy of the former relative to the latter in moving from MeCN to methanol.

Quite a different picture emerges when the MeCN-DMF solvent system is considered. Unlike for the system involving MeOH, the $\Delta_t G°$ value for the free cation from MeCN to DMF is relatively small indicating that these solvents do not offer a differentiating medium for this cation. Since the ligand contribution does not favour complexation in DMF, the slightly higher stability of complex formation in this solvent is due to both; the higher solvation of the metal-ion complex in DMF relative to MeCN and the lower solvation of the silver cation in DMF with respect to MeCN.

The above examples demonstrate that the medium alters the stability of the complex in a selective manner which is not only dependent on the solvation of the free cation (as it is often assumed) but it is the result of the changes in solvation that the reactants and the product undergo upon complexation in the various solvents.

It goes without saying that a detailed interpretation of the thermodynamics associated with the complexation processes involving macrocycles and guest species requires very accurate data. A suitable

Table II. Enthalpies of coordination in kJ mol⁻¹ derived from solution and complexation data in different solvents at 298.15 K

Solvent	$\Delta_s H°$ AgClO$_4$	$\Delta_s H°$ **7b**	$\Delta_s H°$ Ag**7b**ClO$_4$	$\Delta_c H°$	$\Delta_{coord} H°$
MeOH	-14.22	12.93	26.10	-44.39	-71.78
EtOH	-19.18	11.68	28.56	-36.34	-72.40
MeCN	-47.50	24.36	25.77	-23.16	-72.07
DMF	-51.19	23.02	3.89	-39.39	-71.45

Data from refs. 63 and 66.

mean of checking the reliability of thermodynamic data is provided by the calculation of the thermodynamics of coordination (see eq. 4) [17, 52] which as stated above, these data should be the same independently of the solvent from which these are derived. Table II shows enthalpies of coordination for the system involving AgClO$_4$ and **7b** derived from four different solvents through eq. 4. Excellent agreement is found between these set of data which clearly reflect the accuracy of the data.

Final remarks

Although the above discussion has been mainly centered on thermodynamic aspects of macrocyclic chemistry, our research covers a wide range of activities, the main goal is the introduction of these ligands into polymeric frameworks for the production of recyclable materials for the removal of ionic and neutral species of environmental concern from natural sources. I should also emphasise that collaborative research with other institutions in the UK and abroad has been very rewarding. Within this context, I gratefully acknowledge the financial support provided by the European Union, the EPSRC (UK), the US Army (USA), Zeneca (UK) and The British Council (UK). This support has allowed the development of extensive links with other institutions in the UK, Europe and Latin America. As far as the latter is concerned, the outcome of these links is well documented [71-74].

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