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### Thermochemical Investigations of Preferential Solvation in Nonelectrolyte Solutions. Estimation of Preferential Solvation from Measured Solute Solubilities in Binary Solvent Mixtures

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# THERMOCHEMICAL INVESTIGATIONS OF PREFERENTIAL SOLVATION IN NONELECTROLYTE SOLUTIONS. ESTIMATION OF PREFERENTIAL SOLVATION FROM MEASURED SOLUTE SOLUBILITIES IN BINARY SOLVENT MIXTURES

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A relatively simple method is developed for studying preferential solvation in nonelectrolyte solutions based on the Nearly Ideal Binary Solvent (NIBS) and Extended NIBS model mathematical representation for how solute chemical potential varies with solvent composition. Expressions are derived for estimating the extent of preferential solvation around a nonelectrolyte solute from measured solubilities in binary solvent mixtures containing nonspecific and specific interactions. The computational method is illustrated with published data for anthracene, pyrene and carbazole dissolved in mixtures containing dibutyl ether plus *n*-octane and isooctane. Numerical values based on the derived expressions show the dissimilar preferential solvation behavior expected in mixtures containing the two noncomplexing polycyclic aromatic hydrocarbon solutes as compared to the complexing carbazole solute, which is believed to form a 1:1 molecular complex with dibutyl ether.

**KEY WORDS:** Nonelectrolytes, chemical potential, molecular complex.

## INTRODUCTION

Preferential solvation and molecular complexation occur in many systems having biological and chemical significance. Calculation of meaningful solvation numbers and association constants is an important first step in understanding molecular interactions and solution nonideality. Interpretation of solution nonideality generally has followed two dissimilar lines: the “physical” approach originated by van Laar<sup>1</sup> and the “chemical” approach proposed by Dolezalek.<sup>2</sup> The physical approach may be described in terms of a random distribution of molecules throughout the entire solution while the chemical approach is characterized in terms of a specific geometric orientation of one molecule with respect to an adjacent molecule. Even in systems known to contain specific interactions, the need to properly account for nonspecific interactions has been recognized. Arnett *et al.*<sup>3</sup> with their “pure base” calorimetric method for determining enthalpies of hydrogen bond formation attempted to separate

specific and nonspecific interactions. The sensitivity of the numerical results to the selection of "model" compound and inert solvent raised important doubts regarding the overall effectiveness of this particular method.<sup>4</sup> Saluja *et al.*<sup>5</sup> used a somewhat similar rationale in their comparison of enthalpies of transfer of alkanes and alkenes from the vapor state to methanol, dimethylformamide, benzene and cyclohexane, with the more exothermic values for the alkanes in methanol and dimethylformamide attributed to dipole-induced dipole interactions between the solvent and the polarizable  $\pi$ -bond. Bertrand<sup>6</sup> demonstrated that neglect of nonspecific interactions in the chloroform + triethylamine system can lead to an appreciable error in the enthalpy of complex formation determined with the Ideal Associated Solution (IAS) model. Fenby *et al.*<sup>7</sup> presented a similar analysis of the diethyl ether + chloroform system. Smith and Hepler<sup>8</sup> extended the IAS model to ternary triethylamine + chloroform + cyclohexane mixtures and introduced two empirical "correctional" terms to account for triethylamine-cyclohexane and chloroform-cyclohexane nonspecific interactions. The correctional terms, which are calculable from measured binary data, have shown only limited success when used to predict excess enthalpies and excess volumes of ternary acetone + cyclohexane + chloroform<sup>9,10</sup> and acetone + bromoform + alkane mixtures.<sup>11,12</sup> Thermodynamic consequences of the IAS model have been presented in three recent monographs,<sup>13-15</sup> several articles<sup>16-23</sup> and a review<sup>24</sup>, along with discussions of applications to select nonelectrolyte systems.

Many of the remaining methods for estimating complex formation constants can be classified as solubility methods. That is, the increase in solubility of a solute at constant fugacity in a complexing-inert solvent mixture, relative to the solubility in pure inert solvent, is generally attributed to the formation of molecular complexes. This primary assumption is common to several thermodynamic methods, such as the partition of solutes between two immiscible liquid phases, the measurement of infinite dilution gas-liquid chromatographic partition coefficients and the increased solubility of solids. The techniques for calculating formation constants are essentially identical for all solubility methods, as are the difficulties in properly assessing what portion of solubility enhancement is due to nonspecific interactions.

Thermodynamic studies of very weak association complexes are particularly difficult as nonspecific interactions can contribute significantly to the overall solution nonideality. Furthermore, it is conceivable that the observed nonideal behavior might simply result from preferential solvation of the solute by one of the solvent components without the actual formation of a distinct, new chemical species. Often there is no definitive experimental evidence to prove the existence of molecular complexes in solution, and in such instances, the measured thermodynamic properties can be described equally well by one of the many associated solution models<sup>13-15,20,21,25-30</sup> or semi-empirical local composition/preferential solvation models<sup>13-15,30-35</sup> derived during the past several years.

For the most part, local composition models have evolved from the so-called hypothetical "two fluid treatment" or from empirical modifications of the Flory-Huggins model, as was the case with the Wilson equation.<sup>36</sup> Recently, Ben-Naim<sup>37</sup> proposed a new definition and method of measuring nonelectrolyte solute preferential solvation in binary solvent mixtures by evaluating the Kirkwood-Buff integrals.<sup>38,39</sup>

The extent of preferential solvation is estimated from the variation of solute chemical potential as a function of binary solvent composition<sup>37</sup>

$$(\partial\mu_S^{\text{liq}}/\partial n_A)_{T,P,n_B,n_S} = (\partial\mu_A^{*\text{liq}}/\partial n_A)_{T,P,n_B,n_S} - RT \bar{V}_A/V_{\text{soln}} \quad (1)$$

$$\lim_{\rho_S \rightarrow 0} (\partial\mu_S^{*\text{liq}}/\partial n_A)_{T,P,n_B,n_S} = RT(\rho_A + \rho_B)^2(G_{BS} - G_{AS})/\eta \quad (2)$$

where  $\mu_S^{\text{liq}}$  is the chemical potential of the solute (component  $S$ );  $X_A$ ,  $\mu_A$  and  $n_A$  refer to mole fraction, number density and number of moles of solvent component  $A$ , respectively; and the difference  $G_{BS} - G_{AS}$  is a simple measure of the difference in the relative affinities of  $B$  towards  $S$  and  $A$  towards  $S$ . Note that  $\eta$  is a calculable quantity through an inversion of the Kirkwood–Buff theory.<sup>40</sup> Since  $\eta$  must always be positive, the sign of  $G_{BS} - G_{AS}$  will be dictated by the sign of the partial derivative. It is only in the very special case when  $G_{BS} = G_{AS}$  that there is no preferential solvation. Hall,<sup>41</sup> Newman<sup>42</sup> and more recently Marcus<sup>43</sup> discussed a similar approach for estimating preferential solvation around ionic solutes.

Application of Eqs (1) and (2) has been limited to a large extent by the unavailability of experimental data determined at constant number of moles of solute. Through thermodynamic modeling though, it is possible to circumvent this problem by mathematically expressing variation of solute chemical potential as a function of composition. The resulting expression can be differentiated with respect to  $n_A$ , holding  $T$ ,  $P$ ,  $n_B$  and  $n_S$  constant. In this paper, I present a method for calculating preferential solvation around a nonelectrolyte solute molecule from measured saturation solubilities in binary solvent mixtures, using the Nearly Ideal Binary Solvent (NIBS)<sup>44–48</sup> and Extended NIBS<sup>25–28</sup> models as two possible mathematical representations of solute chemical potential. The nature of solubility measurements preclude experimental determinations at constant  $n_S$ . Previously published anthracene,<sup>49</sup> pyrene<sup>50</sup> and carbazole<sup>26,28</sup> solubilities in select binary alkane + dibutyl ether solvent mixtures are used to illustrate the method.

## THE NIBS MATHEMATICAL REPRESENTATION OF SOLUTE CHEMICAL POTENTIAL

The specific form of the basic NIBS model which has been most successful in describing the excess chemical potential of solutes is based on a simple mixing model of a multicomponent system:

$$\Delta\bar{G}^{\text{mix}} = RT \sum_{i=1}^N n_i \ln \phi_i + \left( \sum_{i=1}^N X_i \bar{V}_i \right) \left( \sum_{i=1}^N \sum_{j>i}^N \phi_i \phi_j A_{ij} \right) \quad (3)$$

in which  $n_i$  is the number of moles of component  $i$ ,  $\phi_i$  is the volume fraction,  $\bar{V}_i$  is the pure component molar volume and  $A_{ij}$  is a binary interaction parameter that is independent of composition. Application of Eq. (3) to a ternary system containing

solute  $S$  and solvents  $A$  and  $B$  takes the form of

$$\Delta G^{\text{mix}} = RT[n_A \ln \phi_A + n_B \ln \phi_B + n_S \ln \phi_S] + (n_A \bar{V}_A + n_B \bar{V}_B + n_S \bar{V}_S) \times [\phi_A \phi_B A_{AB} + \phi_A \phi_S A_{AS} + \phi_B \phi_S A_{BS}] \quad (4)$$

with the chemical potential of the solute, relative to the pure liquid ( $\mu_S^{\text{liq}}$ ), being

$$\mu_S^{\text{liq}} - \mu_S^{\text{liq}} = RT[\ln \phi_S + 1 - \bar{V}_S/\bar{V}_{\text{soln}}] + \bar{V}_S[\phi_A(1 - \phi_S)A_{AS} + \phi_B(1 - \phi_S)A_{BS} - \phi_A \phi_B A_{AB}] \quad (5)$$

obtained through the appropriate differentiation.

Having a thermodynamic mixing model that accurately describes solute chemical potential at different compositions, the quantity  $(\partial \mu_S^{\text{liq}}/\partial n_A)_{T,P,n_B,n_S}$  can be easily evaluated

$$(\partial \mu_S^{\text{liq}}/\partial n_A)_{T,P,n_B,n_S} = -\frac{RT\bar{V}_A}{V_{\text{soln}}} - \frac{RT\bar{V}_S}{V_{\text{soln}}} \left[ \frac{(\phi_B + \phi_S)\bar{V}_A}{\bar{V}_A} - \frac{\phi_B \bar{V}_A}{\bar{V}_B} - \frac{\phi_S \bar{V}_A}{\bar{V}_S} \right] + (\bar{V}_S \bar{V}_A/V_{\text{soln}})[\phi_A \phi_S A_{AS} + (1 - \phi_S)(\phi_B + \phi_S)A_{AS} - \phi_B A_{BS} - \phi_B(\phi_B + \phi_S)A_{AB} + \phi_A \phi_B A_{AB}] \quad (6)$$

and substituted back into Eqs (1) and (2). These mathematical manipulations yield the following expression

$$(\partial \mu_S^{\text{liq}}/\partial X_A)_{T,P,n_B,n_S} = -\frac{RT\bar{V}_S \bar{V}_A}{(X_B + X_S)\bar{V}_{\text{soln}}} \left[ \frac{(\phi_B + \phi_S)}{\bar{V}_A} - \frac{\phi_B}{\bar{V}_B} - \frac{\phi_S}{\bar{V}_S} \right] + \frac{\bar{V}_S \bar{V}_A}{(X_B + X_S)\bar{V}_{\text{soln}}} [\phi_A \phi_S A_{AS} + (1 - \phi_S)(\phi_B + \phi_S)A_{AS} - \phi_B A_{BS} - \phi_B(\phi_B + \phi_S)A_{AB} + \phi_A \phi_B A_{AB}] \quad (7)$$

which in the limit of  $\rho_S \rightarrow 0$  simplifies to

$$\lim_{\rho_S \rightarrow 0} (\partial \mu_S^{\text{liq}}/\partial X_A)_{T,P,n_B,n_S} = -\frac{RT\bar{V}_S \bar{V}_B}{(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B)^2} \left[ 1 - \frac{\bar{V}_A}{\bar{V}_B} \right] + \frac{\bar{V}_A \bar{V}_B \bar{V}_S}{(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B)^2} [A_{AS} - A_{BS} + (\phi_A^0 - \phi_B^0)A_{AB}] \quad (8)$$

where the superscript (0) denotes the initial binary solvent composition, calculated as if the solute were not present. Examination of Eqs (2) and (8) reveals that the preferential solvation around a solute molecule is governed by the solvent molar volume ratio, the three binary  $A_{ij}$  parameters and the binary solvent composition which appears as a difference in the last term in Eq. (8). Even in the very special case when the solvent components form an ideal solution ( $A_{AB} = 0$ ) and both solute-solvent  $A_{ij}$  interaction parameters are equal, preferential solvation would still result from solvent molecular size disparity.

To date, the basic NIBS model has been extensively tested in regards to its ability to predict iodine, benzil, *p*-benzoquinone, biphenyl, benzoic acid, anthracene, pyrene and thianthrene solubilities in a wide range of binary organic solvent mixtures containing only nonspecific interactions. Past studies<sup>13,49-55</sup> document that predicted values based on Eq. (5) differ from experimentally measured solubilities by less than  $\pm 4\%$  for many of the more than 120 systems studied. Any of this previously published data can be used to illustrate our calculational method for estimating preferential solvation from measured solute solubilities in binary solvent mixtures. After careful considerations, I decided to use anthracene<sup>49</sup> and pyrene<sup>50</sup> solubilities in binary isooctane + dibutyl ether and *n*-octane + dibutyl ether mixtures. Both solutes are fairly insoluble in these mixtures, and the observed maximum mole fraction solubilities of  $X_{\text{Anthra}}^{\text{sat}} = 0.00361$  and  $X_{\text{Pyrene}}^{\text{sat}} = 0.0298$  should be sufficiently small so as to correspond with the limiting  $\rho_S \rightarrow 0$  condition used to derive Eqs (2) and (8). Furthermore, the three solvents are of comparable molecular size, and any preferential solvation must therefore arise from differences in the  $A_{ij}$  interaction parameters.

Solubilities are thermodynamically related to the solute chemical potential and Eq. (5) as follows:

$$RT \left[ \ln(a_S^{\text{solid}}/\phi_S^{\text{sat}}) - (1 - \phi_S^{\text{sat}}) \left( 1 - \frac{\bar{V}_S}{X_A^0 \bar{V}_A + X_B^0 \bar{V}_B} \right) \right] = (1 - \phi_S^{\text{sat}})^2 \bar{V}_S [\phi_A^0 A_{AS} + \phi_B^0 A_{BS} - \phi_A^0 \phi_B^0 A_{AB}] \quad (9)$$

where the activity of the solid solute ( $a_S^{\text{solid}}$ ), relative to the hypothetical subcooled liquid, is calculated from

$$\ln a_S^{\text{solid}} = - \frac{\Delta \bar{H}_S^{\text{fus}}(T_{\text{mp}} - T)}{RTT_{\text{mp}}} + \frac{\Delta \bar{C}_p(T_{\text{mp}} - T)}{RT} - (\Delta \bar{C}_p/R) \ln(T_{\text{mp}}/T) \quad (10)$$

the molar heat of fusion ( $\Delta \bar{H}_S^{\text{fus}}$ ) at the normal melting point temperature ( $T_{\text{mp}}$ ) and the differences between the molar heat capacities of the liquid and solid. If the solute solubility is sufficiently small so that  $(1 - \phi_S^{\text{sat}})^2 \approx 1$ , then it is mathematically possible to eliminate  $a_S^{\text{solid}}$  from Eq. (8) since the expression contains the difference of two solute-solvent interaction parameters, and *a priori* knowledge of  $\Delta \bar{H}_S^{\text{fus}}$  is not required. Enthalpy of fusion data is readily available in the chemical literature for both anthracene and pyrene, and numerical values of  $a_{\text{Anthr}}^{\text{solid}} = 0.01049$ <sup>49</sup> and  $a_{\text{Pyrene}}^{\text{solid}} = 0.1312$ <sup>50</sup> will be used in all calculations.

Inspection of Eq. (9) reveals that for systems obeying this simple mixing model, both  $A_{AS}$  and  $A_{BS}$  binary interaction parameters can be numerically determined from the appropriate binary reduction

$$RT \left[ \ln(a_S^{\text{solid}}/\phi_S^{\text{sat}}) - (1 - \phi_S^{\text{sat}}) \left( 1 - \frac{\bar{V}_S}{\bar{V}_A} \right) \right] = (1 - \phi_S^{\text{sat}})^2 \bar{V}_S A_{AS} \quad (11)$$

$$RT \left[ \ln(a_S^{\text{solid}}/\phi_S^{\text{sat}}) - (1 - \phi_S^{\text{sat}}) \left( 1 - \frac{\bar{V}_S}{\bar{V}_B} \right) \right] = (1 - \phi_S^{\text{sat}})^2 \bar{V}_S A_{BS} \quad (12)$$

using measured solute solubilities in the two pure solvents as input parameters in Eqs (11) and (12). Furthermore, the  $A_{AB}$  parameter can be calculated from liquid-vapor equilibria data, or from published excess Gibbs free energies relative to either Raoult's Law ( $\Delta\bar{G}_{AB}^{\text{ex}}$ ) or the Flory-Huggins definition of solution ideality ( $\Delta\bar{G}_{AB}^{\text{fh}}$ )

$$\Delta\bar{G}_{AB}^{\text{fh}} = \Delta\bar{G}_{AB}^{\text{ex}} + RT[\ln(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B) - X_A^0 \ln \bar{V}_A - X_B^0 \ln \bar{V}_B] \quad (13)$$

$$\Delta\bar{G}_{AB}^{\text{fh}} = (X_A^0 \bar{V}_A + X_B^0 \bar{V}_B) \phi_A^0 \phi_B^0 A_{AB} \quad (14)$$

or can be estimated *via* the Scatchard-Hildebrand solubility parameter theory<sup>56</sup>

$$\Delta\bar{G}_{AB}^{\text{fh}} = (X_A^0 \bar{V}_A + X_B^0 \bar{V}_B) \phi_A^0 \phi_B^0 (\delta_A - \delta_B)^2 + RT[\ln(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B) - X_A^0 \ln \bar{V}_A - X_B^0 \ln \bar{V}_B] \quad (15)$$

where  $\delta_i$  refers to the solubility parameter of component  $i$ . Unfortunately, excess Gibbs free energies of binary isooctane + dibutyl ether and  $n$ -octane + dibutyl ether could not be found in the chemical literature, and thus numerical values of the two  $A_{AB}$  parameters are based on the Scatchard-Hildebrand solubility parameter theory,

**Table 1** Calculated solute-solvent and solvent-solvent binary  $A_{ij}$  interaction parameters.

Component (i)-Component (j)	$A_{ij}$ (cal cm <sup>-3</sup> )
<i>n</i> -Octane-Anthracene	6.921
Isooctane-Anthracene	9.064
Dibutyl Ether-Anthracene	4.272
<i>n</i> -Octane-Pyrene	8.163
Isooctane-Pyrene	10.558
Dibutyl Ether-Pyrene	5.620
<i>n</i> -Octane-Carbazole	15.241
Isooctane-Carbazole	17.031
Dibutyl Ether-Carbazole	12.532 ( $K_{BS}^{\phi} = 25$ ) <sup>a</sup>
Dibutyl Ether-Carbazole	13.200 ( $K_{BS}^{\phi} = 30$ ) <sup>b</sup>
<i>n</i> -Octane-Dibutyl Ether	0.048
Isooctane-Dibutyl Ether	0.810

<sup>a</sup> For *n*-octane cosolvent.

<sup>b</sup> For isooctane cosolvent.

**Table 2** Solute and solvent pure component properties used in calculations.

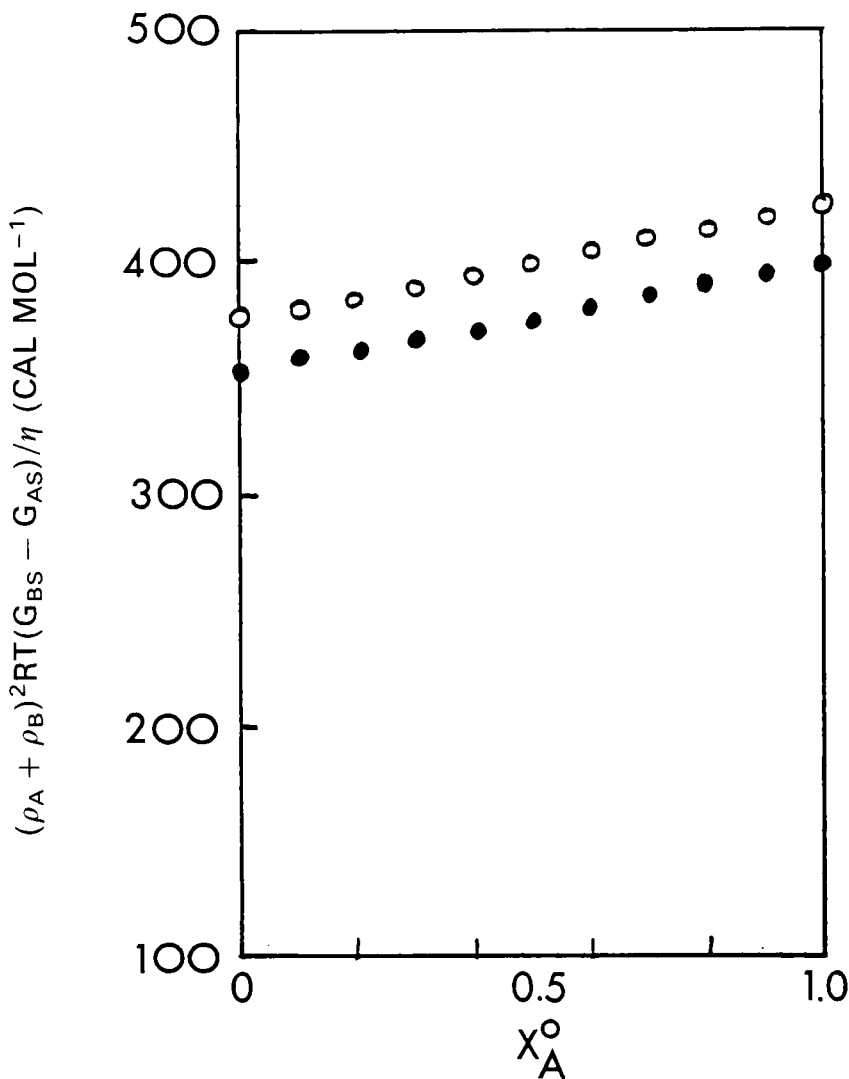
Component (i)	$\bar{V}_i$ (cm <sup>3</sup> mol <sup>-1</sup> )	$\delta_i$ (cal <sup>1/2</sup> cm <sup>-3/2</sup> ) <sup>a</sup>
<i>n</i> -Octane	163.46	7.54
Isooctane	166.09	6.86
Dibutyl Ether	170.41	7.76
Anthracene	150.00	( $a_S^{\text{solid}} = 0.01049$ )
Pyrene	166.50	( $a_S^{\text{solid}} = 0.1312$ )
Carbazole	150.00	( $a_S^{\text{solid}} = 0.009354$ ) <sup>b</sup>

<sup>a</sup> Values taken from literature tabulations.<sup>26,61</sup>

<sup>b</sup> Based on calorimetric data determined by Radomska and Radomski.<sup>62</sup>

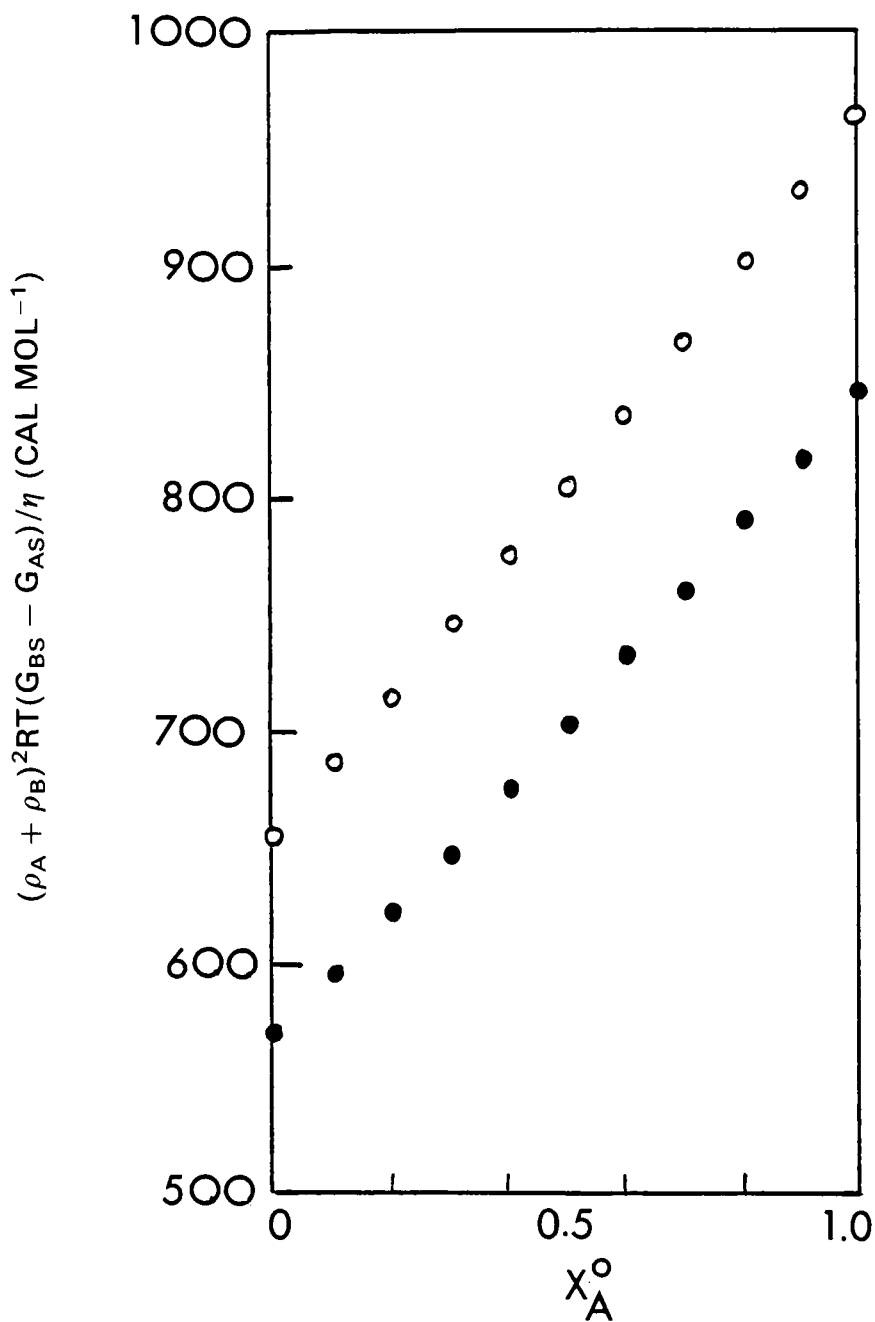
as was also the case in the earlier NIBS solubility comparisons.<sup>49,50</sup> Calculated  $A_{ij}$  parameters and pure solute/solvent properties are listed in Tables 1 and 2, respectively.

Figures 1 and 2 graphically depict calculated values of  $(\rho_A + \rho_B)^2 RT(G_{BS} - G_{AS})/\eta$  versus  $X_A^0$  for both anthracene and pyrene dissolved in binary *n*-octane + dibutyl ether and isooctane + dibutyl ether solvent mixtures. The "extent" of preferential solvation, as measured by the quantity of  $RT(\rho_A + \rho_B)^2 RT(G_{BS} - G_{AS})/\eta$ , shows very little dependence on solvent composition in the case of the two *n*-octane + dibutyl



**Figure 1** Variation of  $(\rho_A + \rho_B)^2 RT(G_{BS} - G_{AS})/\eta$  with solvent composition for anthracene ( $\bullet$ ) and pyrene ( $\circ$ ) dissolved in binary *n*-octane (A) + dibutyl ether (B) mixtures.





**Figure 2** Variation of  $(\rho_A + \rho_B)^2 RT(G_{BS} - G_{AS})/\eta$  with solvent composition for anthracene (●) and pyrene (○) dissolved in binary isooctane (A) + dibutyl ether (B) mixtures.

ether systems. The average molar volume of the solvent is essentially constant (ie.,  $X_A^0 \bar{V}_A + X_B^0 \bar{V}_B = \text{constant}$ ) because the two pure solvents are of comparable molecular size, and the contributions from the  $(\phi_A^0 - \phi_B^0)A_{AB}$  term are not sufficient to cause significant deviations from the horizontal straight line. Calculated values for the two isooctane + dibutyl ether systems show a strong solvent dependence arising from the much larger estimated binary solvent nonideality. Remember the  $A_{AB}$  binary interaction parameter was estimated *via* Eq. (15), and the solubility parameters of the two pure solvents differ by  $0.90 \text{ cal}^{1/2} \text{ cm}^{-3/2}$ .

### THE EXTENDED NIBS MATHEMATICAL REPRESENTATION OF SOLUTE CHEMICAL POTENTIAL

The generalized NIBS mixing model also provides the starting point for the Extended NIBS treatment, but now the four summations in the  $\Delta G^{\text{mix}}$  expression extend over all chemical species believed to be present in the associated solution. Application of Eq. (3) to a quaternary solution ( $A, B, S$  and  $BS$ ) would take the form

$$\begin{aligned}
 B_1 + S_1 &\rightleftharpoons BS & K_{BS}^\phi &= \hat{\phi}_{BS}/(\hat{\phi}_B \hat{\phi}_{S_1}) \\
 \Delta G^{\text{mix}} &= RT[\hat{n}_A \ln \hat{\phi}_A + \hat{n}_{B_1} \ln \hat{\phi}_{B_1} + \hat{n}_{S_1} \ln \hat{\phi}_{S_1} + \hat{n}_{BS} \ln \hat{\phi}_{BS}] \\
 &+ (\hat{n}_A \bar{V}_A + \hat{n}_{B_1} \bar{V}_B + \hat{n}_{S_1} \bar{V}_{S_1} + \hat{n}_{BS} \bar{V}_{BS})[\hat{\phi}_A \hat{\phi}_{B_1} A_{AB_1} + \hat{\phi}_A \hat{\phi}_{S_1} A_{AS_1} \\
 &+ \hat{\phi}_A \hat{\phi}_{BS} A_{ABS} + \hat{\phi}_{B_1} \hat{\phi}_{S_1} A_{B_1 S_1} + \hat{\phi}_{B_1} \hat{\phi}_{BS} A_{B_1 BS} + \hat{\phi}_{S_1} \hat{\phi}_{BS} A_{S_1 BS}] \quad (16)
 \end{aligned}$$

where the carets ( $\hat{\phantom{x}}$ ) denote the "true" compositions in the associated solution. Six  $A_{ij}$  terms are now needed to describe the nonspecific physical interactions. Through mathematical manipulations and clever approximation of the  $A_{B_1 BS}$  and  $A_{S_1 BS}$  parameters, Acree *et al.*<sup>25,27</sup> showed that the solute chemical potential could be expressed in terms of

$$\begin{aligned}
 \mu_S^{\text{liq}} &= \mu_S^{\text{liq}} + RT \left\{ \ln[\phi_A/[1 + \bar{V}_S K_{BS}^\phi \phi_B/(\bar{V}_S + \bar{V}_B)]] + 1 \right. \\
 &- \bar{V}_S \left[ \frac{\phi_S}{\bar{V}_S[1 + \bar{V}_S K_{BS}^\phi \phi_B/(\bar{V}_B + \bar{V}_S)]} + \frac{\phi_A}{\bar{V}_A} + \frac{\phi_B}{\bar{V}_B} \right] \left. \right\} \\
 &+ \bar{V}_S[\phi_A(1 - \phi_S)A_{AS_1} + \phi_B(1 - \phi_S)A_{B_1 S_1} - \phi_A \phi_B A_{AB_1}] \quad (17)
 \end{aligned}$$

of the stoichiometric compositions, volume fraction based  $K_{BS}^\phi$  association constant and only three binary interaction parameters. These latter four quantities are again calculable either from<sup>25,26</sup>

$$\begin{aligned}
 \ln \phi_S^{\text{sat}} &= \phi_A^0 \ln(\phi_S^{\text{sat}})_A + \phi_B^0 \ln(\phi_S^{\text{sat}})_B + \ln[1 + \bar{V}_S K_{BS}^\phi \phi_B^0/(\bar{V}_B + \bar{V}_S)] \\
 &- \phi_B^0 \ln[1 + \bar{V}_S K_{BS}^\phi/(\bar{V}_B + \bar{V}_S)] + \frac{\bar{V}_S \Delta \bar{G}_{AB}^{\text{th}}}{RT(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B)} \quad (18)
 \end{aligned}$$

$$RT \left\{ \ln[a_S^{\text{solid}}/(\phi_S^{\text{sat}})_B] + \ln[1 + \bar{V}_S K_{BS}^\phi/(\bar{V}_B + \bar{V}_S)] - 1 + \frac{\bar{V}_S}{\bar{V}_B} \right\} \approx \bar{V}_S A_{B_1 S_1} \quad (19)$$

the measured solubilities of a sparingly soluble solute in the two pure solvents  $[(\phi_S^{\text{sat}})_A$  and  $(\phi_S^{\text{sat}})_B]$  and binary solvent mixtures, or from experimental  $\Delta\bar{G}_{AB}^{\text{ex}}$  data (see Eqs (13) and (14)). For the noncomplexing solvent,  $A_{AS_1}$  is calculated *via* Eq. (11). The computational method for obtaining the “best”  $K_{BS}^\phi$  value is discussed in detail elsewhere.<sup>26</sup>

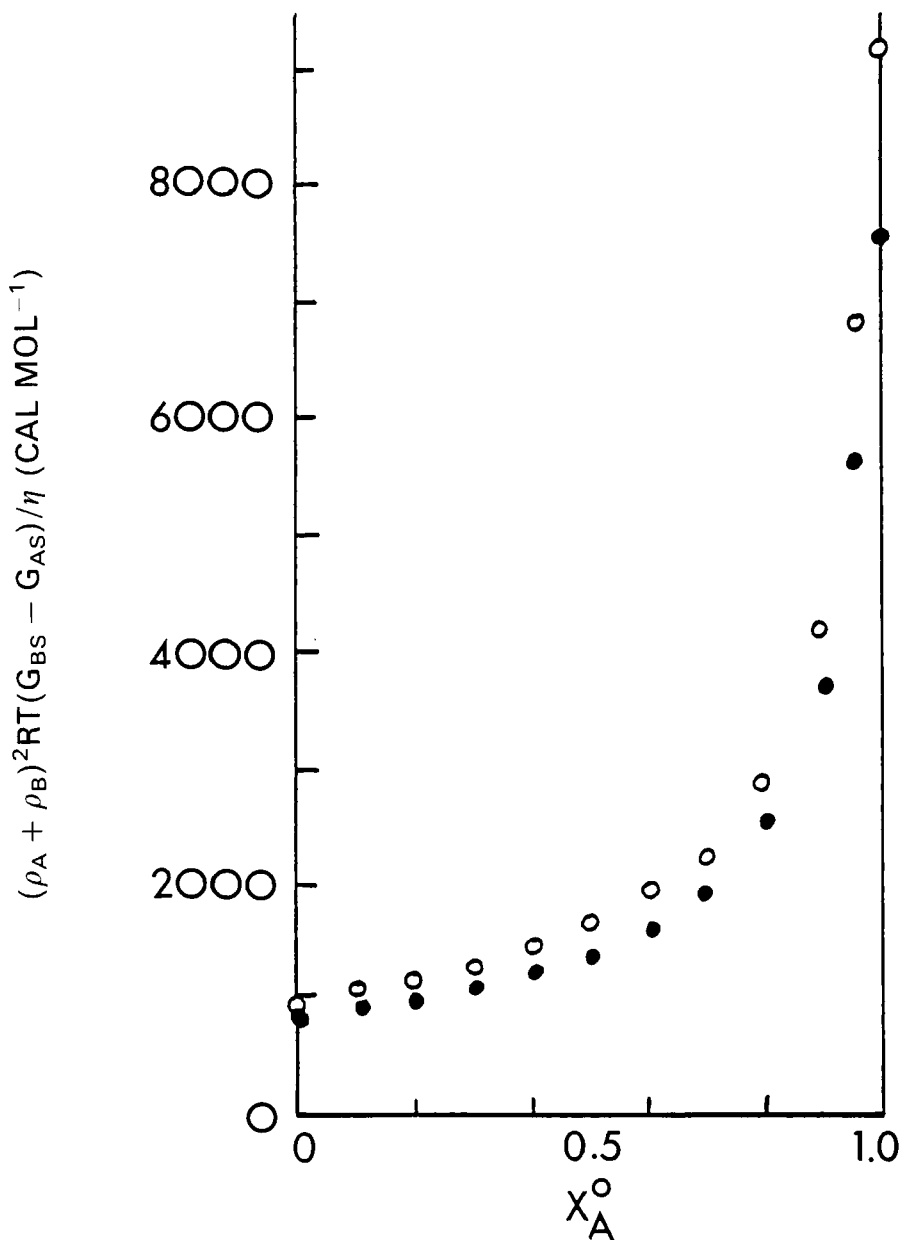
The Extended NIBS model provides a convenient mathematical representation for the compositional dependence of solute chemical potential in systems containing solute-solvent complexation. Differentiation of solute chemical potential with respect to  $n_A$ , followed by suitable algebraic rearrangements, gives the following expression

$$\lim_{\rho_S \rightarrow 0} (\partial \mu_S^{\text{liq}} / \partial X_A)_{T, P, n_B, n_S} = \frac{RT \bar{V}_B \bar{V}_A \bar{V}_S K_{BS}^\phi / (\bar{V}_B + \bar{V}_S)}{(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B)^2 [1 + \bar{V}_S \phi_B^0 K_{BS}^\phi / (\bar{V}_B + \bar{V}_S)]} - \frac{RT \bar{V}_S \bar{V}_B}{(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B)^2} \left[ 1 - \frac{\bar{V}_A}{\bar{V}_B} \right] + \frac{\bar{V}_S \bar{V}_A \bar{V}_B}{(X_A^0 \bar{V}_A + X_B^0 \bar{V}_B)^2} \times [A_{AS_1} - A_{BS_1} + (\phi_A^0 - \phi_B^0) A_{AB_1}] \quad (20)$$

as the solute concentration goes to zero. Equation (20) differs from its NIBS counterpart by only the first term. In the absence of solute-solvent complexation ( $K_{BS}^\phi = 0$ ), Eqs (8) and (20) are identical. Readers should be aware that a number of alternative expressions can be derived by mathematically representing solute chemical potential in terms of different solution models. The solution models used, however, should be based on the Flory–Huggins definition of solution ideality, otherwise the  $-RT \bar{V}_S / V_{\text{soln}}$  term in Eq. (1) will not properly cancel and the resulting  $\lim_{\rho_S \rightarrow 0} (\partial \mu_S^{\text{liq}} / \partial X_A)_{T, P, n_B, n_S}$  derivative will be infinite at  $X_B^0 = 0$ . To use Raoult’s law and mole fraction based association constants, one will need to empirically modify Eq. (1) by replacing  $RT \bar{V}_A / V_{\text{soln}}$  with  $RT / (n_A + n_B + n_S)$ .

Our past solubility studies have focussed primarily on nonelectrolyte solutes dissolved in systems of nonspecific interactions. It was only recently with the development of the Extended NIBS<sup>25–28</sup> and Competitive Associated NIBS<sup>29</sup> models that we began to systematically examine the effect of the inert hydrocarbon cosolvent on solute-solvent association parameters calculated from measured solute solubilities. To date, the only complexing systems studied have been carbazole dissolved in ten alkane + dibutyl ether<sup>26,28,58</sup> and four chloroalkane + dibutyl ether<sup>29,59</sup> mixtures, and pyrene dissolved in six alkane + 1,4-dichlorobutane<sup>60</sup> binary solvent systems. McCargar and Acree<sup>26,28</sup> documented that Eq. (18) described the experimental carbazole solubilities in binary alkane + dibutyl ether mixtures to within an average absolute deviation of 2% using a single carbazole-dibutyl ether association constant, which varied from  $K_{BS}^\phi = 22$  for *n*-heptane to  $K_{BS}^\phi = 30$  for isooctane cosolvent.

Figure 3 graphically depicts variation of  $(\rho_A + \rho_B)^2 RT (G_{BS} - G_{AS}) / \eta$  with  $X_A^0$  for carbazole dissolved in mixtures containing dibutyl ether plus *n*-octane and isooctane. Careful examination reveals that differences in the calculated values are more pronounced at compositions near the pure saturated hydrocarbon cosolvent. Only a small quantity of dibutyl ether is needed to completely surround the nonelectrolyte solute molecule dissolved at infinite dilution. The “extent” of preferential solvation



**Figure 3** Variation of  $(\rho_A + \rho_B)^2 RT(G_{BS} - G_{AS})/\eta$  with solvent composition for carbazole dissolved in binary solvent mixtures containing dibutyl ether plus *n*-octane (●) and isooctane (○). Equilibrium constants used in the calculations were  $K_{BS}^0 = 25$  for *n*-octane and  $K_{BS}^0 = 30$  for isooctane cosolvent.

sharply drops off as  $X_B^0$  increases and then levels out to a relatively constant value, which would be expected when the complexing solvent composition begins to exceed that of the inert hydrocarbon cosolvent. The dramatic dissimilarities noted in Figures 1–3 suggest that this type of treatment can provide valuable quantitative information regarding solute-solvent interactions in solution. Simulated calculations using different  $K_{BS}^\phi$  values indicate that it will be fairly easy to identify systems containing strong solute-solvent interactions as significant nonlinear behavior will be noted in the corresponding  $(\rho_A + \rho_B)^2 RT(G_{BS} - G_{AS})/\eta$  versus  $X_A^0$  plots. It will be extremely difficult, however, to identify very weak molecular complexes since the plots become more linear when  $K_{BS}^\phi$  approaches zero. In closing, readers are reminded that through thermodynamic modeling it is possible to evaluate the “extent” of preferential solvation from measured solute solubilities in binary solvent mixtures. The methodology developed as part of the present study can be applied to a large number of nonelectrolyte solutes, experimental partition coefficient and chromatographic retention data, and different mathematical representations of solute chemical potential.

#### References

1. J. J. van Laar, *Z. Physik. Chem.*, **72**, 723 (1910).
2. F. Dolezalek, *Z. Physik. Chem.*, **64**, 727 (1908).
3. E. M. Arnett, L. Joris, E. Mitchell, T. S. S. R. Murty, T. M. Gorrie and P. v. R. Schleyer, *J. Am. Chem. Soc.*, **92**, 2365 (1970).
4. W. C. Duer and G. L. Bertrand, *J. Am. Chem. Soc.*, **92**, 2587 (1970).
5. P. P. S. Saluja, T. M. Young, R. F. Rodewald, F. H. Fuchs, D. Kohli and R. Fuchs, *J. Am. Chem. Soc.*, **99**, 2949 (1977).
6. G. L. Bertrand, *J. Phys. Chem.*, **79**, 48 (1975).
7. D. V. Fenby, A. Chand, A. Inglese, J.-P. E. Grolier and H. V. Kehiaian, *Austr. J. Chem.*, **30**, 1401 (1977).
8. J. F. Smith and L. G. Hepler, *Can. J. Chem.*, **67**, 1153 (1989).
9. W. E. Acree, Jr., A. I. Zvaigzne and S. A. Tucker, *Phys. Chem. Liq.*, **21**, 169 (1990).
10. W. E. Acree, Jr., A. I. Zvaigzne and S. A. Tucker, *Thermochim. Acta*, in press.
11. P. P. Singh, R. Malik, S. Maken, W. E. Acree, Jr. and S. A. Tucker, *Thermochim. Acta*, in press.
12. P. P. Singh, R. Malik, S. Maken and W. E. Acree, Jr., *J. Chem. Soc., Faraday Trans.*, in press.
13. W. E. Acree, Jr., *Thermodynamic Properties of Nonelectrolyte Solutions*, Academic Press, Inc., Orlando, FL (1984).
14. M. L. McGlashan, *Chemical Thermodynamics*, Academic Press, Inc., New York, NY (1979).
15. J. M. Prausnitz, R. N. Lichtenthaler and E. Gomes de Azevedo, *Molecular Thermodynamics of Fluid-Phase Equilibria*, 2nd edition, Prentice-Hall, Englewood Cliffs, NJ (1986).
16. L. Barta, Z. S. Kooner, L. G. Hepler, G. Roux-Desgranges and J.-P. E. Grolier, *J. Solution Chem.*, **18**, 663 (1989).
17. J.-P. E. Grolier, G. Roux-Desgranges, Z. S. Kooner, J. F. Smith and L. G. Hepler, *J. Solution Chem.*, **16**, 745 (1987).
18. L. G. Hepler, Z. S. Kooner, G. Roux-Desgranges and J.-P. E. Grolier, *J. Solution Chem.*, **14**, 579 (1985).
19. L. Barta, Z. S. Kooner, L. G. Hepler, G. Roux-Desgranges and J.-P. E. Grolier, *Can. J. Chem.*, **67**, 1225 (1989).
20. W. E. Acree, Jr. and S. A. Tucker, *Phys. Chem. Liq.*, **19**, 23 (1989).
21. W. E. Acree, Jr., S. A. Tucker and L. E. Cretella, *Thermochim. Acta*, **145**, 245 (1989).
22. H. P. Singh, P. P. Singh and S. Dagar, *Can. J. Chem.*, **66**, 290 (1988).
23. A. Apelblat, A. Tamir and M. Wagner, *Fluid Phase Equilib.*, **4**, 229 (1980).
24. D. V. Fenby and L. G. Hepler, *Chem. Soc. Reviews*, **3**, 193 (1974).
25. W. E. Acree, Jr., D. R. McHan and J. H. Rytting, *J. Pharm. Sci.*, **72**, 929 (1983).
26. J. W. McCargar and W. E. Acree, Jr., *Phys. Chem. Liq.*, **17**, 123 (1987).
27. W. E. Acree, Jr. and J. W. McCargar, *J. Mol. Liq.*, **37**, 251 (1988).
28. J. W. McCargar and W. E. Acree, Jr., *J. Solution Chem.*, **17**, 1081 (1988).
29. W. E. Acree, Jr. and J. W. McCargar, *J. Pharm. Sci.*, **76**, 575 (1987).

30. C. Panayiotou and J. H. Vera, *Fluid Phase Equilib.*, **5**, 55 (1980).
31. L. L. Lee and K. E. Starling, *Fluid Phase Equilib.*, **21**, 77 (1985).
32. O. B. Nagy, M. waMuanda and J. B. Nagy, *J. Chem. Soc., Faraday Trans. 1*, **74**, 2210 (1978).
33. M. Szpakowska and O. B. Nagy, *J. Phys. Chem.*, **93**, 3851 (1989).
34. M. Szpakowska and O. B. Nagy, *J. Chem. Soc., Faraday Trans. 1*, **85**, 2891 (1989).
35. W. E. Acree, Jr., A. I. Zvaigzne and S. A. Tucker, *J. Chem. Soc., Faraday Trans.*, **86**, 307 (1990).
36. G. M. Wilson, *J. Am. Chem. Soc.*, **86**, 127 (1964).
37. A. Ben-Naim, *Cell Biophys.*, **12**, 255 (1988).
38. J. G. Kirkwood and F. P. Buff, *J. Chem. Phys.*, **19**, 774 (1951).
39. A. Ben-Naim, *Water and Aqueous Solutions*, Plenum, New York, NY (1974).
40. A. Ben-Naim, *J. Chem. Phys.*, **67**, 4884 (1977).
41. D. G. Hall, *J. Chem. Soc., Faraday Trans. 1*, **68**, 25 (1972).
42. K. E. Newman, *J. Chem. Soc., Faraday Trans. 1*, **84**, 1387 (1988).
43. Y. Marcus, *J. Chem. Soc., Faraday Trans. 1*, **85**, 3019 (1989).
44. T. E. Burchfield and G. L. Bertrand, *J. Solution Chem.*, **4**, 205 (1975).
45. W. E. Acree, Jr. and G. L. Bertrand, *J. Phys. Chem.*, **81**, 1170 (1977).
46. W. E. Acree, Jr. and G. L. Bertrand, *J. Solution Chem.*, **12**, 101 (1983).
47. W. E. Acree, Jr. and G. L. Bertrand, *J. Pharm. Sci.*, **70**, 1033 (1981).
48. W. E. Acree, Jr. and G. L. Bertrand, *J. Phys. Chem.*, **83**, 2355 (1979).
49. M. V. Marthandan and W. E. Acree, Jr., *J. Chem. Eng. Data*, **32**, 301 (1987).
50. J. R. Wallach, S. A. Tucker, B. M. Oswald, D. J. Murrall and W. E. Acree, Jr., *J. Chem. Eng. Data*, **34**, 70 (1989).
51. C. L. Judy, N. M. Pontikos and W. E. Acree, Jr., *Phys. Chem. Liq.*, **16**, 179 (1987).
52. S. A. Tucker and W. E. Acree, Jr., *Phys. Chem. Liq.*, **19**, 73 (1989).
53. W. E. Acree, Jr., *Int. J. Pharm.*, **18**, 47 (1984).
54. W. E. Acree, Jr. and S. A. Tucker, *Phys. Chem. Liq.*, **20**, 31 (1989).
55. W. E. Acree, Jr., S. A. Tucker and A. I. Zvaigzne, *Phys. Chem. Liq.*, **21**, 45 (1990).
56. J. H. Hildebrand, J. M. Prausnitz and R. L. Scott, *Regular and Related Solutions*, Van Nostrand Reinhold, New York, NY (1970).
57. W. E. Acree, Jr., *J. Pharm. Sci.*, **76**, 580 (1987).
58. J. W. McCargar and W. E. Acree, Jr., *J. Pharm. Sci.*, **76**, 572 (1987).
59. J. W. McCargar and W. E. Acree, Jr., *J. Solution Chem.*, **18**, 151 (1989).
60. W. E. Acree, Jr., S. A. Tucker and A. I. Zvaigzne, *J. Chem. Soc., Faraday Trans.*, in press.
61. K. L. Hoy, *J. Paint Technol.*, **42**, 76 (1970).
62. M. Radomska and R. Radomski, *Thermochim. Acta*, **40**, 405 (1980).