Conclusions

The simple ion–water potential functions reported here when used in conjunction with the TIP4P potential for water were demonstrated to provide reasonable descriptions of dilute aqueous solutions of Li$^+$, Na$^+$, F$^-$, and Cl$^-$. The computed heats of solution showed improvement over previous work, and the structural results are in excellent accord with available diffraction data. The average coordination numbers for Li$^+$ and Na$^+$ are 5 and 6, while F$^-$ and Cl$^-$ participate in an average of 6 and 7 linear hydrogen bonds, respectively. Furthermore, the weaker ion–solvent interactions for Cl$^-$ lead to a more flexible first shell exhibiting the broadest range of coordination numbers (6–10).

A scheme for decomposing the computed energetics of the solvation process was provided and led to several observations. (1) The orienting influence of the ions is clearly apparent in the net repulsive solvent–solvent interactions that are found between water molecules in the first shell. (2) The total ion–solvent attraction is about evenly divided between the first shell and the remainder of the solvent. (3) The solvent disruption is localized in the first shell and in the interface to the first shell. This point was supported by hydrogen bonding analyses which revealed that water molecules beyond only 4–5 Å from the ions participate in normal numbers of hydrogen bonds.

The definition of the energy components also permitted comparisons of results for the solutions and gas-phase ion–water clusters. It was found that the energetics associated only with the first solvent shell ($E_{SX1} + E_{11}$) are enough to determine the trends in total heats of solution for the ions. Moreover, the ions plus their first solvation shells are structurally and energetically similar to ion–molecule clusters in the gas phase, if the coordination numbers match the cluster size. These observations provide important links between gas-phase and solution chemistry.

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Thermodynamics of Solution of Naphthalene in Various Water–Ethanol Mixtures[‡]

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Abstract: The temperature dependences of the solubilities of naphthalene in water and in various water–ethanol mixtures up to 0.07 mol fraction of ethanol have been determined. In the concentration range studied, a plot of the free energy of solution of naphthalene vs. mole fraction of ethanol present in the solvent is linear. The heat of solution of naphthalene in water is positive and becomes more positive as the ethanol concentration is increased. This heat term is more than overcome by a concomitant large increase in the entropy of solution, the net effect being to solubilize the hydrocarbon by the addition of ethanol. The addition of ethanol to the aqueous phase is much more effective in solubilizing naphthalene at high rather than at low temperatures. The results may be interpreted in terms of the McMillan–Mayer second virial coefficient. It is suggested that ethanol may exert its influence upon hydrocarbon solubility by loss of water structure as hydrophobic interactions take place between the hydrocarbon and alcohol. Dispersal of an iceberg or clathrate structure probably makes a significant contribution.

Introduction

The solubility of naphthalene in mixed organic solvents has been of interest to the present authors because, by making some reasonable simplifying assumptions, it is possible to relate the variation of solubility of a substrate such as naphthalene to the ability of the enzyme(s) cytochrome P-450 to bind it.[1] Since the solubility of naphthalene was found to vary in an interesting manner with variations in composition of water–ethanol solvent mixtures at a single temperature and since there appears to be very little data available on the effects of addition of organic solvents upon the aqueous solubility of hydrocarbons, it was deemed advisable to study the temperature dependence of the solution process in order to ascertain the effects of added organic solvent (ethanol) upon the thermodynamics of solution for naphthalene. Such a study might shed some light upon the process of solubilization of a simple but fairly bulky aromatic hydrophobic compound by the addition of an organic solvent to the aqueous medium. In addition, the use of a crystalline solid tends to minimize the effects of extraction of ethanol from the aqueous medium by a liquid hydrocarbon were such a compound chosen for the study.

The purpose of the present work is to determine the dependence of the thermodynamics of solution of naphthalene in solvent mixtures ranging from pure water to a water–ethanol mixture containing 0.07 mol fraction of ethanol. In this range, it has been found that the free energy of solution varies in a linear manner with the mol fraction of ethanol present in the ethanol–water solvent. At higher ethanol concentrations than this, a different law applies to the solubilities and will be considered at a later time.

Experimental Section

Materials and Methods. The naphthalene was obtained from the Fisher Scientific Co. and was zone-refined. In addition, it was recrystallized from 95% ethanol before use (mp 80–81 °C). The water–ethanol solutions were prepared from double-glass-distilled water and absolute ethanol (USP) obtained from U.S. Industrial Chemicals Co., New York, NY, and used without further treatment.

Naphthalene (0.2 g) was weighed out and rapidly added to 100 mL of water or water–ethanol mixture in a Virtis homogenizer flask which had been sealed and allowed to equilibrate at the working temperature for 30 min. The Teflon cover was attached and adjusted for best seal. The solution was stirred at approximately 25 000 rpm for a total of 2 min with a Virtis Model 23 homogenizer. Actually, no material increase in solubility could be demonstrated by increasing the stirring time beyond 1 min. The other details of the experimental techniques, equipment, temperature control, and measurement have been previously described.[2,3] The sintered glass tube used for sampling (previously equilibrated at working temperature; see ref 3) was rinsed with the naphthalene solution

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[‡] Supported by the Department of Energy/Morgantown Energy Research Center, Contract DY-77-C-21-8087, and the West Virginia Medical Corporation.

(1) W. L. Backes and W. J. Canady, *Pharmacol. Ther.* **12**, 133 (1981).

(2) R. J. Larese and W. J. Canady, *J. Phys. Chem.*, **65**, 1240 (1961).

(3) R. J. Larese, D. A. Robinson, W. F. Brassine, and W. J. Canady, *J. Phys. Chem.*, **66**, 897 (1962).

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Table I. Constants for Equation 1 in Text

mol fraction of ethanol	$-A$	$-B$	$-C$
0	289.086 46	1.044 94	6.920 38
0.016	291.056 33	1.055 13	6.977 76
0.033	296.054 15	1.080 36	7.131 80
0.051	287.212 94	1.065 73	7.012 70
0.070	275.188 22	1.045 93	6.853 85

Table II. Constants for Equation 2 in Text[a]

mol fraction of ethanol	$-\Delta H^\circ_0$	$\Delta C_p/R$	$-I$
0	1767.4601	17.952 09	120.551 67
0.016	3657.4153	25.703 52	170.772 05
0.033	2705.7566	23.962 85	157.378 62
0.051	2057.1651	23.301 987	151.155 87
0.070	2317.2164	25.773 79	165.772 37

[a] Solubility is expressed in mole fraction. Note: the above values are either second derivatives ($\Delta C_p/R$) or are obtained by a very long extrapolation of the experimental data to a temperature of absolute zero. Therefore, small errors in the experimental range can lead to large deviations. In calculating the thermodynamic parameters in the experimental range, these deviations tend to cancel. Therefore, it is not reasonable, for example, to expect ΔH°_0 to necessarily vary in a similar manner with ethanol concentration as $\Delta H^\circ_{298.2}$ does.

Table III. Experimental Solubilities of Naphthalene in Water and in Various Ethanol–Water Solutions[a]

mole fraction of ethanol	T, °K	solubility × 10^6
0	275.05	1.696
	283.85	2.416
	288.55	3.041
	294.85	3.755
	298.35	4.318
	303.85	5.634
	308.25	6.509
	312.45	7.702
	318.05	9.684
0.0159	275.25	2.066
	283.85	3.112
	289.25	3.735
	294.85	4.961
	298.25	5.864
	303.75	7.643
	308.15	9.143
	312.95	11.436
	318.05	13.881
0.0329	275.00	2.339
	283.85	3.850
	288.74	4.715
	294.85	6.533
	298.55	8.270
	303.65	10.169
	313.15	12.213
	317.55	20.347
0.0508	274.45	2.611
	283.95	4.730
	288.75	6.193
	294.85	8.444
	298.05	10.503
	303.85	13.655
	307.75	16.867
	313.05	23.006
	317.55	29.186
0.0508	273.75	3.221
	283.85	5.950
	288.85	7.904
	294.85	11.853
	298.35	14.689
	303.65	19.824
	308.35	26.737
	313.25	35.009
	317.25	44.635

[a] Solubilities are expressed in mol fraction.

Table IV. Thermodynamic Parameters for the Solution of Naphthalene in Water and in Various Ethanol–Water Solutions at 298.2 K (25 °C)

mol fraction of ethanol	0	0.0159	0.0329	0.0508	0.07
ΔG°, kcal/mol	7.31	7.14	6.97	6.80	6.58
ΔH°, kcal/mol	7.13	7.96	8.82	9.72	10.67
ΔS°, eu	−0.6	2.8	6.2	9.8	13.7
ΔC_p cal/mol deg	36	51	48	46	51

three times to saturate the glass surfaces with hydrocarbon. Finally a 0.5-mL sample was removed from the tube by means of a calibrated Gilson pipet (±0.2%) and placed in a 1-cm cuvette equipped with a glass stopper containing 2.5 mL of absolute ethanol. The pipet was rinsed several times with the aforementioned ethanol–water mixture to effect quantitative transfer. The cuvette was stoppered and inverted a number of times to accomplish complete mixing. The concentration of naphthalene was then determined spectrophotometrically using a Cary 17D spectrophotometer. The reference cuvette contained solvent of the same composition as that of the sample. Both cuvettes were maintained at 25 ± 0.2 °C. Readings were made at 275 nm; in agreement with the findings of Bohon and Claussen,[4] the absorbance of naphthalene was found to follow Beer's law and not to vary with the relatively small changes in ethanol–water ratio in the final mixture. Indeed those authors found this to be true over a much wider range of such solvent variation than used here. Each experimental point is based on at least four measurements. The measurements were found to be repeatable within a mean deviation of approximately 1.8%. Since the samples were obtained at a particular temperature under consideration, and then the aliquot transferred to a cuvette containing ethanol at 25 °C, there would be a systematic error introduced unless the change in density of the solution with temperature were taken into account. In order to avoid this difficulty, the densities of the water and ethanol–water mixtures were fit to an equation of the form

$$\ln \text{density} = (A/T) + B \ln T + C \tag{1}$$

where T is, of course, the absolute temperature.

The data were obtained from the "International Critical Tables".[5] It was found that over the experimental range investigated here, the densities calculated by means of the above equation agreed within better than 3 parts per 10 000 with the measured values. The densities at the particular experimental temperatures used in this work were then calculated and the appropriate volume corrections made for 25 °C. The not unreasonable assumption was made that the density of a given water or ethanol–water solution would not be materially different from the same solution containing something of the order of 10^{-4} M naphthalene. The constants for eq 1 are found in Table I.

Results and Discussion

The corrected solubility values (converted to mole fraction) were then fitted to an equation of the same form as eq 1:

$$\ln \text{sol} = \frac{\Delta H^\circ_0}{RT} + \frac{\Delta C_p}{R} \ln T + I \tag{2}$$

as suggested by Everett and Wynne-Jones[6] to describe the temperature dependence of the ionization of various weak acids and bases. The term ΔH°_0 represents the hypothetical standard partial molar change in heat content at absolute zero, while ΔC_p is, by definition, $\partial \Delta H^\circ/\partial T$; I is an integration constant. The constants for eq 2 are found in Table II and the experimental data from which they were derived are seen in Table III. The thermodynamic parameters at 25 °C are tabulated in Table IV. The standard partial molar free energy change ΔG° is written as

$$\Delta G^\circ = -RT \ln \text{sol} \tag{3}$$

$$\Delta H^\circ = \Delta H^\circ_0 + \Delta C_p T \tag{4}$$

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[4] R. L. Bohon and W. F. Claussen, *J. Am. Chem. Soc.*, **73**, 1571 (1951).

[5] "International Critical Tables", Vol. III, E. W. Washborn Ed., McGraw-Hill, New York, 1928, pp 115–119.

[6] D. H. Everett and W. F. K. Wynne-Jones, *Trans. Faraday Soc.*, **35**, 1380 (1939).

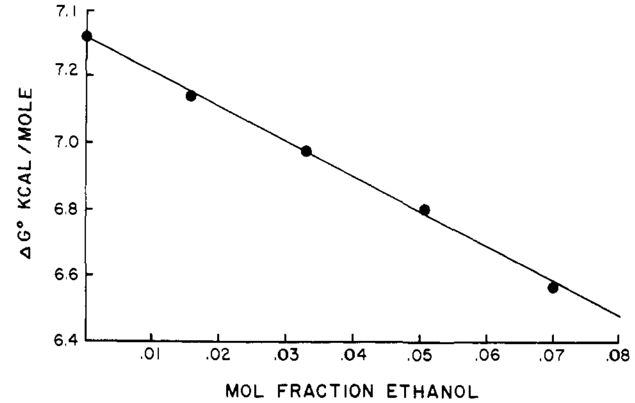

Figure 1. Plot of the free energy of solution of naphthalene vs. mol fraction of ethanol present at 25 °C.

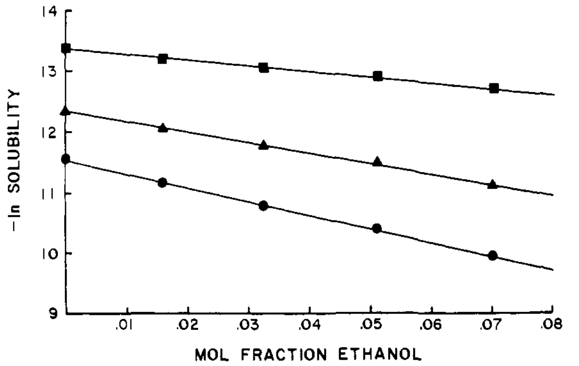

Figure 2. Plot of the logarithm of the solubility of naphthalene vs. mol fraction of ethanol present at 0, 25, and 45 °C.

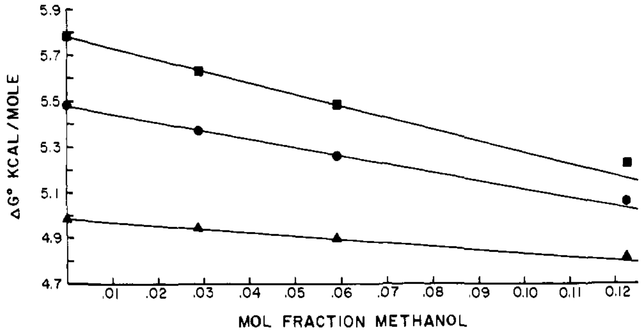

Figure 3. Plot of the free energy of solution of toluene vs. mol fraction of methanol present at, from top to bottom, 0, 25, and 45 °C (data of Ueda et al.[8]).

where ΔH° is the standard partial molar change in heat content at a specific temperature under consideration. The corresponding entropy change ΔS° was calculated in the usual way. It is clear from Table III that solubilization takes place by means of an entropic effect, the heat of solution actually becoming more positive as the ethanol concentration is increased.

A comparison of our data with previous work, insofar as that is possible, is in order. The solubility of naphthalene in water at 25 °C as determined from our measurements is 2.43×10^{-4} M. The value obtained by Bohon and Claussen[4] was 2.68×10^{-4} M. While our value is somewhat lower than that obtained by those authors, it is in very good agreement with the work of Andrews and Keefer[7] who obtained 2.45×10^{-4} M. On the other hand, the corresponding value for the heat of solution at 25 °C of naphthalene in water is, according to our results 7.13 kcal/mol as compared with 7.15 kcal/mole from the data of Bohon and Claussen.[4] Therefore, there is very good agreement between the two studies insofar as the temperature dependence of the solubility in water is concerned although our absolute values tend to be somewhat lower than theirs.

The differences in ΔC_p, the change in heat capacity obtained by us as compared with that of Bohon and Claussen,[4] is easily understandable since this term is a second differential and very small errors lead to rather large differences in the calculated value for ΔC_p. Bohon and Claussen quote a value of 77 or 146 cal/mol·°C in water, whereas our value is 36 cal/mol·°C. We have obtained no evidence to indicate a break in the value of ΔC_p at approximately 22 °C as they report. To give the reader an idea of the sensitivity of ΔC_p to experimental error, an error of ±1% in the one solubility determination at 0 °C leads to a change of about 15% in the calculated value for ΔC_p. Ueda et al.[8] report a ΔC_p of 44 cal/mol·°C for benzene in water as compared with 108 cal/mol·deg as reported by Bohon and Claussen. It is apparent that very large differences may be obtained for ΔC_p while other thermodynamic parameters for a given experiment might agree quite well. It is difficult to discern any definite trend of ΔC_p with the limited variations in solvent composition; one difficulty is due to the fact that, contrary to the cases of benzene and toluene, in dealing with naphthalene, we are working at some distance from T_0, that temperature at which $\Delta H^\circ = 0$. Put another way, we are trying to distinguish between a series of lines that in the experimental range deviate relatively little from linearity. The rate of variation of slope with temperature is therefore particularly difficult to determine accurately over a somewhat limited temperature range. We believe that, overall, our results are in satisfactory agreement with Bohon and Claussen, especially when one takes into account the technical difficulties involved in measuring the water solubilities of hydrocarbons.

The results of Ueda et al.[8] reinforce our confidence in the validity of our own measurements for naphthalene, since plotting their data for the solution of benzene and toluene in water and methanol–water mixtures produces results that are very similar to our observations of naphthalene in ethanol–water mixtures. It should be emphasized that Ueda et al.[8] did not analyze their data in the same way that we have done and did not point out the linear relationships inherent, but apparently unobserved, in their work which we discuss here. They did, however, point out a similar entropic effect.

Figure 1 shows a plot of ΔG° of solution of naphthalene vs. mole fraction of ethanol present in the ethanol–water solutions at 298.2 K; good linearity is observed. Figure 2 shows the natural logarithm of the solubility plotted against mole fraction of ethanol present at three temperatures: 273.2, 298.2, and 318.2 K. The same type of relationship holds for other temperatures as well. The slope of the line increases as the temperature increases, indicating that solubilization of napthalene by ethanol is much more effective at higher temperatures while relatively little solubilization takes place at low temperatures. The solubilization of benzene and toluene by methanol was studied by Ueda et al.[8] and, although they did not note a similar linear relationship to that shown here, a plot of their values of free energy of solution for toluene vs. mole fraction of methanol added (see Figure 3) does indeed show a similar result for the smaller liquid hydrocarbons and methanol–water solutions although some deviation is observed at the very high mole fraction of methanol. Their data for benzene follows the same type of law (not shown).

The results obtained by us shown in Figure 2 and by Ueda et al.[8] may be interpreted qualitatively in terms of the McMillan–Mayer second virial coefficient[9] which is a simple measure of the interaction of the hydrocarbon molecules investigated here with ethanol molecules in water.

The slope of Figure 2 may be written

$$\text{slope} = -(\partial \ln X_n / \partial X_e)u_n,\ T,\ P \qquad (5)$$

where n = naphthalene, e = ethanol, X = mole fraction, and u = chemical potential.

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[7] L. J. Andrews and R. M. Keefer, *J. Am. Chem. Soc.*, **71**, 3645 (1949).

[8] M. Ueda, A. Katayama, T. Urahata and N. Kuroki, *Kagaku To Kogyo (Osaka)*, **54**, 252 (1980).

[9] W. G. McMillan and J. E. Mayer, *J. Chem. Phys.*, **13**, 276 (1945).

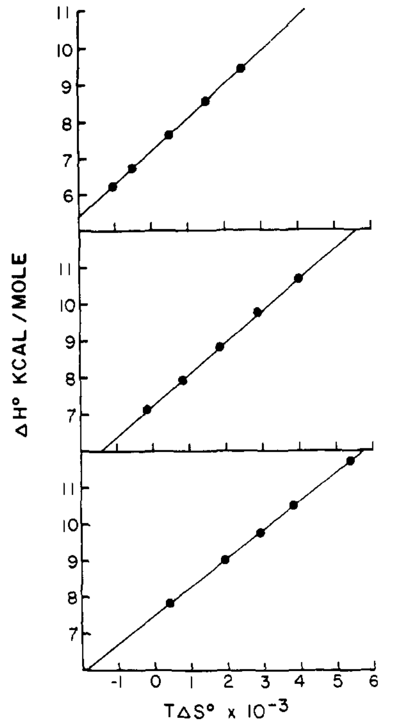

Figure 4. $\Delta H°$ vs. $T\Delta S°$ for the solution of naphthalene in water and various mol fractions of ethanol at, from top to bottom, 0, 25, and 45 °C.

It can be shown[10,11] that if the solutions are dilute and if one neglects some small corrections, the above slope is approximately proportional to

$$\text{slope} = -(\partial \ln C_n/\partial C_e)u_n,\ T,\ u_w \qquad (6)$$

where C is the concentration in molecules per unit volume and w is the solvent. At very low concentrations of ethanol, it is equal to the McMillan–Mayer second virial coefficient, $\bar{P}_{ne}$. It has been shown by Rossky and Friedman[12] for the interactions in water of like hydrophobic molecules (benzene–benzene) that this coefficient is strongly negative, confirming a hydrophobic interaction. Its dependence on temperature is negative; that is to say, the interaction is entropically stabilized. Since naphthalene and similar hydrocarbons are hydrophobic substances, and ethanol and methanol contain hydrophobic moities, one would expect to obtain results similar to those reported here with naphthalene and ethanol as well as those obtained by Ueda et al.[8] with benzene and toluene with methanol.

The McMillian–Mayer theory[9] is one which is useful in applying statistical mechanics to calculate the properties of a model for the solution process. However, in this system a central role is played by an excess Helmholz free-energy term. On the other hand, in dealing with experiments such as those reported here, a corresponding Gibbs free-energy term is involved. Interconversion of these two systems is of great importance; the subject has been considered in detail by Friedman in the above-mentioned references and references contained therein.

Since it is generally observed that linear free-energy relationships involve a compensation between the heat and entropy terms, Figure 4 shows a plot of $\Delta H°$ vs. $T\Delta S°$ for naphthalene at three temperatures, the good linearity confirming this fact. The experimental data of Ueda et al.[8] follow the same kind of law quite nicely for methanol–water solutions of benzene and toluene, (not shown) although, again, they did not point out this fact. They were content to list in tabular form the thermodynamic parameters as a function of per cent methanol of the solvent mixture. Indeed, for naphthalene, $\Delta H°$ and $\Delta S°$ are themselves close to linear in mole fraction of organic solvent added (see Figures 5a and 5b). The question naturally arises as to whether this is a general rule for the solution of aromatic hydrocarbons in mixed aqueous–alcoholic solvents. Figure 6 shows a plot of $\Delta H°$ of solution of toluene vs. mole fraction of methanol calculated from the work of Ueda et al.[8] It is easily seen that a good linear relationship is observed, but at the higher concentration of methanol shown, curvature is to be noted. Therefore, one may conclude that a plot of the heat (or entropy) of solution of an aromtic hydrocarbon vs. mole fraction of alcohol added may, in general, be linear with mole fraction of alcohol added, provided the alcohol concentration

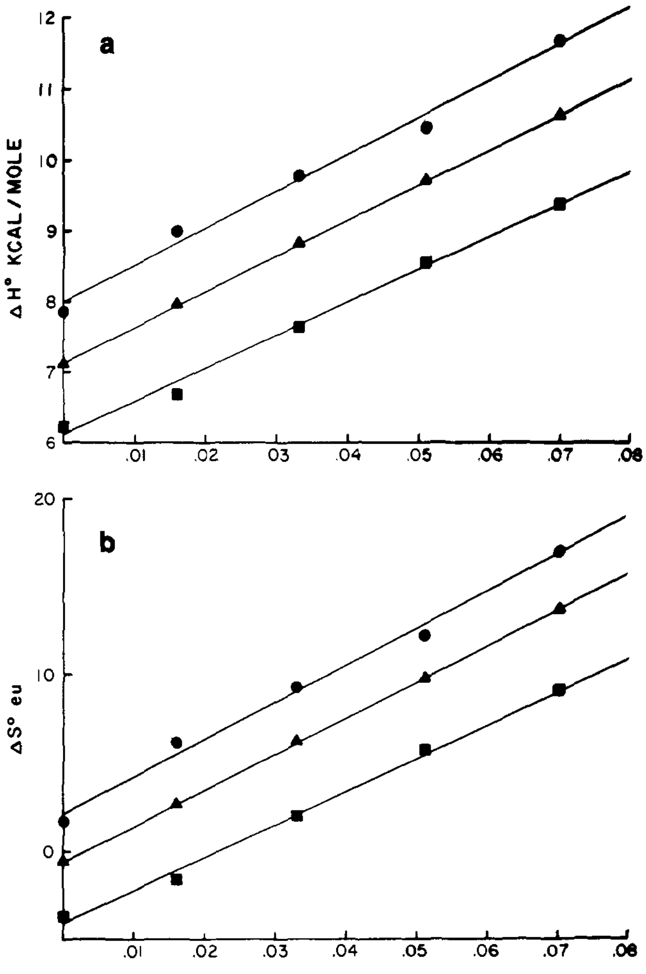

Figure 5. $\Delta H°$ vs. mol fraction of ethanol present for naphthalene. (b) $\Delta S°$ vs. mole fraction of ethanol present for naphthalene. The temperature are, from top to bottom, 0, 25, and 45 °C.

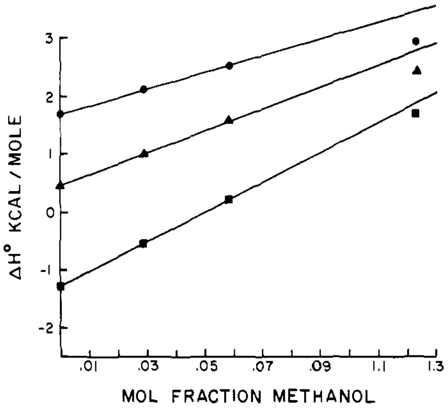

Figure 6. Plot of $\Delta H°$ of solution vs. mol fraction of methanol present for toluene (data of Ueda et al.[8]). The temperatures are, from top to bottom, 0, 25, and 45 °C.

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(10) H. L. Friedman, *J. Solution Chem.*, **1**, 387 (1972).
(11) H. L. Friedman, *J. Phys. Chem.*, **59**, 161 (1955).
(12) P. J. Rossky and H. L. Friedman, *J. Phys. Chem.*, **84**, 587 (1980).

is not too high. Heat–entropy compensation does take place in all cases studied thus far. Such heat–entropy compensations observed in various processes have been discussed in detail by Lumry and Rajender.[13]

Wada and Umeda[14] have observed that the addition of ethanol increases the temperature of maximum density which would infer an increase in water structure. Bosio et al.[15] indicate that variations in density decrease. Franks and Reid[16] have made the suggestion, based on a number of considerations, that ethanol molecules fill voids in the water structure. Overall, this would imply an increased structuring around the ethanol molecules; with fewer voids available, more work must be done to form them in order to dissolve naphthalene molecules. This would explain an increase in $\Delta H°$ for the solution process with the addition of the alcohol.

The great over-balancing entropic effect may be due to the loss of water structure from the environs of both naphthalene and ethanol molecules when hydrophobic interactions take place, the destruction of "iceberg" structure[17] possibly making a significant contribution.

When more data have been acquired with a larger number of hydrocarbons and other compounds of various structural characteristics, it may be possible to deduce empirical laws whereby the effect of the addition of a given organic solvent upon the solubility of a given substance might be predicted with reasonable accuracy.

Ueda et al.[18] have suggested the following for consideration as a convenient way of thinking about the solution process for hydrocarbons. Their solubility data for benzene and toluene were fit to an equation of the following form which is purely empirical:

$$\log N = A(1/T)^2 + B(1/T) + C \qquad (7)$$

where

$$\Delta H° = -R2[2A(1/T) + B] \qquad (8)$$

They further suggest that

$$\Delta H°_{\text{obsd}} = \Delta H°_{\text{iceberg}} + \Delta H_{\text{cavity}} + \Delta \text{H}_{\text{interactn}} \qquad (9)$$

The heat of iceberg formation is negative because of a more ordered structure around the hydrocarbon molecules in solution; the heat of cavity formation is positive and is that contribution required to make a hole for the dissolved hydrocarbon. Structural change in the solvent is not included in this term. The heat of interaction is suggested by them to be the negative heat due to the interactions of hydrocarbon molecules with solvent molecules, including all solvent–solute interactions other than $\Delta H_{\text{iceberg}}$ and ΔH_{cavity}.

They then ascribe the term $-2A(1/T)R$ primarily to disruption and formation of iceberg structure, and the temperature-independent term $-BR$ to cavity formation and to $\Delta H_{\text{interactn}}$ which they suggest might not be too dependent on temperature. The temperature-independent constant C is considered to be an imaginary solubility at an infinite temperature.

We are suggesting a somewhat different approach which has a better foundation in thermodynamics than that of Ueda et al.[18] To make this clear, it might be advantageous to review the rationale of the equation that we have used to express the temperature dependence of the solubility.

The change in heat capacity, ΔC_p is defined according to Kirchoff as

$$(\partial \Delta H°/\partial T)_p = \Delta C_p \qquad (10)$$

A general integration of the above yields

$$\Delta H° = \Delta H°_0 + \int \Delta C_p \, dT \qquad (11)$$

where $\Delta H°_0$ may be regarded as the hypothetical difference in the heat contents of the two states at absolute zero, and $\Delta H°$ is the standard change in heat content at a given finite temperature. Making the not unreasonable assumption that the heat capacities in T, T^2, etc., are not too different in the two states, upon integrating the vant' Hoff equation one obtains eq 2.

Everett and Wynne-Jones[6], by making use of the Nernst heat theorem, evaluated I in eq 2 in terms of $\Delta S°_0$, the hypothetical change in entropy at absolute zero, ΔC_p, and R, obtaining

$$\ln \text{sol} = -\frac{\Delta H°_0}{RT} + \frac{\Delta C_p}{R} \ln T + \frac{(\Delta S°_0 - \Delta C_p)}{R} \qquad (12)$$

and

$$\Delta H° = \Delta H°_0 + \Delta C_p T$$

An equation of the form of eq 2 and 12 was used by Pitzer[19] to describe the temperature dependence of a number of ionization constants for weak acids and bases, and its use was later refined by Everett and Wynne-Jones[6] for a similar purpose with outstanding success.

Therefore, if one postulates, as have Ueda et al.,[18] that there is an essentially temperature-independent heat term due to cavity formation and interaction of solute and solvent, that term is $\Delta H°_0$, the hypothetical change in heat content at absolute zero. The temperature-dependent contributions (presumed to be primarily iceberg) to $\Delta H°$ are determined by ΔC_p. This is true whether one assumes ΔC_p to be independent of temperature (eq 12) or not. The temperature-independent term in the equation used by us for the temperature dependence of the solubility is not the solubility at some infinitely high temperature, but is $(\Delta S°_0 - \Delta C_p)/R$, an entropy containing term.

It should be emphasized again that such terms as $\Delta H°_0$, $\Delta S°_0$, etc., are hypothetical heat and entropy terms at absolute zero that would be real provided that the same laws obtained in the experimental temperature range could be legitmately extrapolated to absolute zero. However, they are not legitimate extrapolations to absolute zero because of various factors including phase changes, etc. Indeed, for example, there are compelling reasons to expect that ΔC_p would tend to approach zero as 0 °K a is approached in a real experiment at such low temperatures.[20]

Since Walde[21] has shown that an equation of the form of eq 7 used by Ueda et al.[8,18] infers that ΔC_p is highly temperature dependent, and since the experimental data do not indicate a large temperature dependence for ΔC_p, one might pose the question as to why that purely empirical equation (eq 7) fits their data for the solution of benzene and toluene so well. According to eq 12 derived from thermodynamic considerations, an equation in T and $\ln T$ should be used rather than in T and T^2. The observation that there is an inversion point in the temperature dependence of the solubilities of both benzene and toluene within the experimental range (roughly 300 K) provides the key to the realization that, to a good approximation, they have actually employed a special form of the equation used by us.[6] At the inversion point, the slope of a ln sol. vs. $1/T$ plot is equal to zero; hence $\Delta H° = 0$. Since the experiments were done by Ueda et al.[8,18] over a relatively limited temperature range not too far removed from T_0 (that temperature at which $\Delta H° = 0$), the relationship between

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[15] L. Bosio, J. Teixeira, and H. E. Stanley, *Phys. Rev. Lett.*, **46**, 597 (1981).

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the two equations becomes apparent.

We express the log of the solubility as a function of temperature by means of eq 12. At T_0 where $\Delta H° = 0$, the natural log of the solubility is

$$\ln \mathrm{sol}_0 = \frac{\Delta H°_0}{RT_0} + \frac{\Delta C_p}{R} \ln T_0 + \frac{(\Delta S°_0 - \Delta C_p)}{R} \qquad (13)$$

Subtracting eq 13 from eq 12, expanding the right-hand term in a power series, and neglecting the terms above those in T^2, one obtains

$$\ln \mathrm{sol} - \ln \mathrm{sol}_0 = -\frac{\Delta H°_0}{R}\left(\frac{1}{T} - \frac{1}{T_0}\right) - \frac{\Delta C_p}{R}\left[\left(\frac{T_0}{T} - 1\right) - \frac{1}{2}\left(\frac{T_0}{T} - 1\right)^2\right] \qquad (14)$$

Keeping in mind that $-\Delta H°_0/R = \Delta C_p T_0/R$, eq 14 can be rewritten as

$$\ln \mathrm{sol} = \frac{\Delta C_p T_0^2}{2R}\left(\frac{1}{T^2}\right) - \frac{\Delta C_p T_0}{R}\left(\frac{1}{T}\right) + \left(\frac{\Delta C_p}{2R} + \ln \mathrm{sol}_0\right) \qquad (15)$$

Since ΔC_p, T_0, R, and ln sol_0 are all temperature independent or assumed to be so. Equation 15 is equivalent to eq 7 as used by Ueda et al.[8,18] where their constants A, B, and C are represented in the following way:

$$A = \Delta C_p T_0^2/2R$$

$$B = -\Delta C_p T_0/R$$

$$C = \Delta C_p/2R + \ln \mathrm{sol}_0$$

Thus it can be seen, for example, that the temperature-independent term C, which they suggested as representing the hypothetical solubility at some infinitely high temperature, actually contains ΔC_p, R, and the logarithm of the solubility at T_0 which falls in the experimental range.

Although the detailed theoretical treatment of aqueous solutions of hydrocarbons, particularly mixed aqueous solutions, promises to be a difficult one, the general thermodynamic laws governing the solubilization process will be useful. For the present we feel that methanol and ethanol exert their influence upon hydrocarbon solubility by bringing about a decrease in water structure and release of somewhat immobilized water molecules as the hydrophobic interaction between hydrocarbon and alcohol takes place in the aqueous medium.

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Registry No. Naphthalene, 91-20-3; ethanol, 64-17-5; water, 7732-18-5.

Behavior of Hexane Dissolved in Dimyristoylphosphatidylcholine Bilayers: An NMR and Calorimetric Study

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Abstract: Deuterium and phosphorous NMR spectroscopy and differential scanning calorimetry (DSC) have been used to examine the behavior of dimyristoylphosphatidylcholine (DMPC) bilayers to which hexane has been added. This work represents a characterization of a two-component bilayer system in which the minor component is a simple example drawn from the large class of hydrophobic molecules that are located in the biological membrane or can be incorporated into it. DSC thermograms indicate that the phase behavior of this mixture is similar to bilayers composed of a mixture of dilauryolphosphatidylcholine and distearoylphosphatidylcholine. Isothermal melting is observed at about 0 °C. Deuterium NMR spectra of perdeuterated hexane dissolved in the bilayer show three overlapping powder patterns indicating that on average both ends of the molecule are experiencing the same environment. Deuterium NMR spectra of macroscopically oriented bilayers indicate that the direction of motional averaging is the same for both the lipids and hexane, i.e., normal to the bilayer plane. The temperature dependence of the hexane–deuterium quadrupole splittings exhibits a maximum at the lipid–bilayer phase-transition temperature. At about the isothermal melting temperature the powder patterns coalesce into a single line, indicating the onset of isotropic motion. Deuterium NMR spectra of acyl-chain perdeuterated DMPC and phosphorus NMR spectra of the DMPC head group change little upon the addition of hexane. Taken together these results indicate that at hexane to lipid ratios of less than about 0.5 the bilayer is intact and the order of the DMPC molecules is little affected by the presence of the alkane. The hexane motion giving rise to three powder patterns is, no doubt, quite complex, but we envision it as encompassing a shuffling between the two monolayers in a direction normal to the bilayer surface as well as rotation about this normal axis and gauche–trans isomerization.

Introduction

The primary structural element of the biological membrane is the lipid bilayer. The interior of the bilayer is a mixture of the alkyl chains of the amphiphilic molecules of which the bilayer is composed. The partition of molecules between the bilayer and the surrounding aqueous phase plays a major role in determining the structure and function of the biological membrane. The most widely accepted architectural model for the membrane is the fluid mosaic model in which integral protein "solute molecules" are dispersed (or aggregated) within the two-dimensional bilayer "matrix" or "solvent".[1] Thus, the interior of the membrane is often taken to be an alkyl solvent in which other molecules are

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