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Solvation thermodynamics of xenon in *n*-alkanes, *n*-alcohols and water

Giuseppe Graziano*

Dipartimento di Scienze Biologiche ed Ambientali Facoltà di Scienze, Università del Sannio Via Port'Arsa,
11 – 82100 Benevento, Italy

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Abstract

Experimental thermodynamic data for the solvation of xenon in *n*-alkanes, water and *n*-alcohols, from different sources, are analysed by means of a general theory of solvation. The standard solvation Gibbs energy change is given by the sum of the work spent to create a cavity suitable to host xenon and the energy gained to turn on xenon-solvent attractive interactions. The latter contribution is larger in magnitude than the former in both *n*-alkanes and *n*-alcohols; the reverse holds in water. This finding is due to the fact that liquid water is characterized by the largest work of cavity creation, because of the smallness of its molecules. Since the two contributions to the solvation Gibbs energy change are both larger in magnitude in *n*-alcohols with respect to *n*-alkanes of the same number of heavy atoms, the xenon solubility is not so different between the two classes of solvents. The H-bonds play an indirect role, being important to determine the density of the liquid. It is also shown that there is no clear correlation between the cohesive energy density of the liquid and the solubility of xenon in the liquid itself. The present study confirms that the effective molecular size of solvent molecules is the principal factor in controlling the solvation Gibbs energy changes.

Keywords: *n*-alkanes and *n*-alcohols; Effective molecular diameter; Solute–solvent and solvent–solvent interactions; H-bonds

1. Introduction

In a recent article Eads [1], in order to support the claim that the presence of H-bonds has a strong influence on solvation thermodynamics, called attention on the $\Delta G^{\bullet}$ (the superscript filled circle indicates the Ben-Naim standard [2,3] that is strongly recommended to analyse solubility data) values of xenon at 20 °C in *n*-alkanes, *n*-alcohols and water, exploiting the experimental data of Pollack and co-workers [4,5]. The experimental values of $\Delta H^{\bullet}$, $\Delta S^{\bullet}$ and $\Delta G^{\bullet}$ for the solvation of xenon in *n*-alkanes [4], water [6], and *n*-alcohols [5], at 20 °C, are collected in Table 1. It is evident that $\Delta G^{\bullet}$ is large positive in water, but negative in both *n*-alkanes and *n*-alcohols. Under Ben-Naim standard conditions, the solvation of xenon is a spontaneous process in both *n*-alkanes and *n*-alcohols, but a non-spontaneous one in water. In addition, the transfer process is dominated by

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*Tel.: +39-0824-305101; fax: +39-0824-23013.
E-mail address: graziano@unisannio.it (G. Graziano).

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Table 1
Ben-Naim standard enthalpy, entropy and Gibbs energy changes associated with the solvation at 20 °C of xenon in *n*-alkanes, water and *n*-alcohols. The values for *n*-alkanes and *n*-alcohols are from [4] and [5], those for water are from [6]

	$\Delta H^{\bullet}$ kJ mol^{-1}	$\Delta S^{\bullet}$ J K^{-1}mol^{-1}	$\Delta G^{\bullet}$ kJ mol^{-1}
n-C_6H_{14}	−9.23	−18.01	−3.95
n-C_7H_{16}	−9.52	−19.65	−3.76
n-C_8H_{18}	−8.86	−17.98	−3.59
n-C_9H_{20}	−8.61	−17.57	−3.46
n-$C_{10}H_{22}$	−8.24	−16.75	−3.33
n-$C_{11}H_{24}$	−7.98	−16.31	−3.20
n-$C_{12}H_{26}$	−8.02	−16.75	−3.11
n-$C_{13}H_{28}$	−8.16	−17.57	−3.01
n-$C_{14}H_{30}$	−7.85	−16.72	−2.95
H_2O	−18.07	−79.14	5.13
CH_3OH	−7.81	−20.05	−1.93
C_2H_5OH	−8.12	−20.15	−2.21
n-C_3H_7OH	−8.51	−20.90	−2.38
n-C_4H_9OH	−8.43	−20.55	−2.41
n-$C_5H_{11}OH$	−8.29	−20.25	−2.35
n-$C_6H_{13}OH$	−8.42	−20.75	−2.34
n-$C_7H_{15}OH$	−8.32	−20.55	−2.30
n-$C_8H_{17}OH$	−8.45	−21.15	−2.25
n-$C_9H_{19}OH$	−7.90	−19.45	−2.22
n-$C_{10}H_{21}OH$	−7.84	−19.35	−2.17
n-$C_{11}H_{23}OH$	−7.77	−19.40	−2.08

enthalpy in both *n*-alkanes and *n*-alcohols, but it is dominated by entropy in water.

Actually the $\Delta G^{\bullet}$ values in *n*-alcohols are slightly smaller in magnitude than those in *n*-alkanes of the same number of heavy (non-hydrogen) atoms: (a) $\Delta G^{\bullet} = -2.35$ kJ mol^{-1} in n-$C_5H_{11}OH$ vs. -3.95 kJ mol^{-1} in n-C_6H_{14}; (b) $\Delta G^{\bullet} = -2.08$ kJ mol^{-1} in n-$C_{11}H_{23}OH$ vs. -3.11 kJ mol^{-1} in n-$C_{12}H_{26}$. In addition, there is a remarkably strong difference between the $\Delta G^{\bullet}$ value of xenon in water, 5.13 kJ mol^{-1}, and those determined in methanol and ethanol, −1.93 and −2.21 kJ mol^{-1}, respectively. Eads claimed that the H-bonding network of liquid water and *n*-alcohols is the main cause of such findings [1].

Moreover, Lazaridis [7], by studying the solvation of methane in suitably constructed water models by means of computer simulations, concluded that 'the essential condition for solvophobicity is that solvent–solvent interactions be much stronger than solute–solvent interactions. A small solvent size does play a role in enhancing the magnitude of the effect, but is not a *sine qua non* condition. Experimental data seem to support this idea. For example, the fact that xenon is more soluble in alkanes than in alcohols of the same size is consistent with the primacy of the cohesive forces.'

Therefore, the rationalization of the solvation thermodynamics of xenon in *n*-alkanes, *n*-alcohols and water seems to be a stringent test for any theoretical approach claiming to describe solvation phenomena. This is exactly the aim of the present work. We would like to show that the solvation thermodynamics of xenon in both *n*-alkanes, *n*-alcohols and water can be rationalized by the suitable application of a well-founded theoretical approach [8–15], that does not explicitly take into account the presence of the H-bonds. Finally, it should not be forgotten that a better understanding of xenon solubility in different solvents might be useful to improve the knowledge of the mechanisms of anaesthesia [16].

2. Theory of solvation

Ben-Naim [17], using statistical mechanics, showed that solvation has to be referred to the transfer of a solute molecule from a fixed point in the ideal gas phase to a fixed point in a liquid phase. Even though this may appear a theoretical definition difficult to use in practice, actually the corresponding standard Gibbs energy change, $\Delta G^{\bullet}$ is readily determined from experimental solubility measurements. In fact, the following relation holds [17]:

$$\Delta G^{\bullet} = RT \times \ln(C_g/C_l) \tag{1}$$

where C_g and C_l are the molar solubility of the solute in the gas phase and in the liquid phase, respectively, in equilibrium conditions.

In order to reach a molecular level understanding of solvation thermodynamics, the transfer process has to be divided in a series of steps. Such steps, however, should have physical reality otherwise the procedure has no value. Insertion of a solute molecule at a fixed point in any liquid can be dissected in two consecutive steps [8–15]: creation of a suitable cavity to host the solute molecule, and turning on the solute–solvent attractive interactions. The first step is an unavoidable consequence of the fact that each molecule occupies a finite volume impenetrable to the others, and that liquids are a condensed state of the matter. The second step accounts for the existence of at least dispersive interactions between the solute molecule and the surrounding solvent molecules. Accordingly, the standard solvation Gibbs energy change is given by [8]:

$$\Delta G^{\bullet} = \Delta G_c + E_a \tag{2}$$

where ΔG_c, the work of cavity creation, and E_a, the average solute–solvent interaction energy, represent the direct perturbation of the solvent [18]. In particular, ΔG_c is the entropic part of the direct perturbation, measuring the probability that a certain region of space is devoid of solvent molecules [19,20]:

$$\Delta G_c = -T \times \Delta S_x \tag{3}$$

where ΔS_x is the excluded volume entropy contribution due to cavity creation in the solvent. As a response to the direct perturbation, solvent molecules undergo a structural reorganization affording further enthalpy and entropy contributions [18]. Thus, the standard solvation enthalpy and entropy changes are given by:

$$\Delta H^{\bullet} = E_a + \Delta H_r \tag{4}$$

$$\Delta S^{\bullet} = \Delta S_x + \Delta S_{nx} \tag{5}$$

where the notation ΔS_{nx} is used to emphasize that it is the non-excluded volume entropy contribution associated with solvation. Clearly, (Eqs. (2)–(5)) mean that the solvent response is a compensating process that does not affect the total Gibbs energy change:

$$T \times \Delta S_{nx} = \Delta H_r \tag{6}$$

The validity of the compensation relation, Eq. (6), was demonstrated first by Ben-Naim [21], who pointed out that $\Delta G^{\bullet}$ depends on the solvent structure, but is not influenced by the solvent structural changes induced by solute insertion. Since then several authors [22–27], using different theoretical routes, have derived relationships corresponding to Eq. (6). A detailed and complete exposition of such theory, along the lines devised by Lee, has been presented in reference [28].

3. Calculation procedure

The work of cavity creation in the different solvents is calculated by means of the formula provided by scaled particle theory [29,30], SPT, assuming spherical both the solute and solvent molecules, on the basis of the equivalent hard sphere concept [31]. To perform calculations the experimental values of the liquid densities at 20 °C are used for all solvents [32], and the pressure is always fixed at 1 atm [30]. In addition, we adopted: (a) for *n*-alkanes, the effective hard sphere diameters determined by Wilhelm and Battino [33], from the analysis of gas solubility data; (b) for *n*-alcohols, the effective hard sphere diameters obtained using the group contributions determined by Ben-Amotz and Willis [34], from the analysis of high pressure equation of state data; (c) for water, $\sigma = 2.80$ Å, estimated from the

Table 2
Molar volume at 20 °C [32], hard sphere diameter [33–36], volume packing density, Lennard-Jones potential parameter, and critical temperature [32] for the solvents used in this work. For Xe, $\sigma = 3.96$ Å and $\varepsilon/k = 230$ K [28]

	υ cm^3 mol^{-1}	σ Å	ξ	ε/k K	T_c K
n-C_6H_{14}	130.52	5.92	0.501	575	507.6
n-C_7H_{16}	146.56	6.25	0.525	630	540.2
n-C_8H_{18}	162.61	6.54	0.542	670	568.7
n-C_9H_{20}	178.74	6.83	0.562	730	594.6
in-$C_{10}H_{22}$	194.92	7.08	0.574	765	617.7
n-$C_{11}H_{24}$	211.20	7.33	0.588	810	639.0
n-$C_{12}H_{26}$	227.51	7.58	0.604	875	658.0
n-$C_{13}H_{28}$	243.75	7.79	0.612	900	675.0
n-$C_{14}H_{30}$	260.09	8.00	0.621	940	693.0
H_2O	18.05	2.80	0.383	120	647.1
CH_3OH	40.49	3.83	0.438	330	512.5
C_2H_5OH	58.37	4.44	0.473	415	514.0
n-C_3H_7OH	74.81	4.93	0.505	500	536.8
n-C_4H_9OH	91.53	5.38	0.537	600	563.0
n-$C_5H_{11}OH$	108.24	5.77	0.560	675	588.1
n-$C_6H_{13}OH$	125.59	6.10	0.570	710	610.3
n-$C_7H_{15}OH$	141.39	6.41	0.587	775	632.6
n-$C_8H_{17}OH$	157.47	6.69	0.600	830	652.5
n-$C_9H_{19}OH$	174.37	6.96	0.610	875	670.7
n-$C_{10}H_{21}OH$	190.78	7.20	0.617	895	687.3
n-$C_{11}H_{23}OH$	207.65	7.42	0.620	895	703.6

location of the first peak in the oxygen-oxygen pair correlation function of water [35,36]. In this respect it is worth noting that the effective hard sphere diameters calculated for n-alkanes using the group contributions of Ben-Amotz and Willis closely corresponds to those determined by Wilhelm and Battino (i.e. 5.92 vs. 5.96 Å for n-C_6H_{14}, 8.00 vs. 7.93 Å for n-$C_{14}H_{30}$, and identical values for n-nonane and n-decane). This indicates the consistency of the two different approaches to obtain effective hard sphere diameters for liquids and the reliability of the values used for n-alcohols. All the σ values are listed in Table 2, together with the volume packing density of each solvent, $\xi = \pi \times \sigma^3 \times N_{\mathrm{Av}}/6 \times \upsilon$, where υ is the molar volume of the solvent itself.

The E_a term measures the van der Waals interaction energy of xenon in the different solvents and is calculated with the formula devised by Pierotti [30], to account for dispersive interactions. The van der Waals interactions of xenon with n-alkanes consist purely of dispersive contributions, whereas those with water and n-alcohols consist of both dispersive and dipole-induced dipole contributions. The latter can be absorbed into the parameterization of dispersive contributions, because both terms depend on the inverse sixth power of distance [37]. The ε/k values used are listed in the fifth column of Table 2. They have been fixed in order to obtain a close correspondence between the experimental $\Delta G^{\bullet}$ data and the $\Delta G_c + E_a$ values, in the assumption that the ΔG_c estimates provided by SPT are numerically exact. Such ε/k numbers are similar to those obtained by Ben-Amotz and colleagues [34,38], using a completely different procedure (i.e. the agreement is really good for n-alcohols for instance, $\varepsilon/k =$ 710 vs. 685 K for n-$C_6H_{13}OH$, and 895 vs. 920 K for n-$C_{11}H_{23}OH$). The strength of attractive interactions among molecules in a liquid is related to the critical temperature, T_c, of the liquid itself [37]: the higher T_c, the stronger the attractive interactions. According to the experimental T_c values listed in the last column of Table 2 [32], the trend of our ε/k estimates along the n-alkane and n-alcohol series is physically reliable. For water we obtained $\varepsilon/k = 120$ K, a number larger than the usual one, approximately 80 K [30], but it accounts for both dispersive and dipole-induced dipole interactions that liquid water can do with the solute. Finally, for xenon we used $\sigma = 3.96$ Å and $\varepsilon/k =$ 230 K in all solvents [28].

4. Results

The ΔG_c values to create a cavity of 3.96 Å diameter at 20 °C, calculated by means of SPT, are listed in the second column of Table 3. Even though they are large positive for all solvents, two clear differences emerge between n-alkanes and n-alcohols. First, ΔG_c is always larger in n-alcohols with respect to n-alkanes. For instance, ΔG_c(in kJ mol^{-1}) = 21.14 in n-$C_5H_{11}OH$ vs. 14.84 in n-C_6H_{14}, and 20.44 in n-$C_{11}H_{23}OH$ vs. 18.09 in n-$C_{12}H_{26}$. Second, ΔG_c increases continuously in the n-alkane series, whereas it is practically constant in the n-alcohol series, with an average value of 20.67 kJ mol^{-1}. Finally, water is characterized by the largest ΔG_c value, 25.13 kJ mol^{-1}.

According to the SPT formula, at a given

Table 3
Estimates of ΔG_c and E_a for the solvation at 20 °C of xenon in n-alkanes, water, and n-alcohols. The values of $\Delta G_c + E_a$ have to be contrasted with the experimental $\Delta G^{\bullet}$ ones. The values of surface tension γ at 25 °C, and of c.e.d. at 20 °C for each liquid are also reported [32]. See text for further details

	ΔG_c kJ mol^{-1}	γ mN m^{-1}	E_a kJ mol^{-1}	$\Delta G_c + E_a$ kJ mol^{-1}	$\Delta G^{\bullet}$ kJ mol^{-1}	c.e.d. J cm^{-3}
n-C_6H_{14}	14.84	17.89	−18.78	−3.94	−3.95	220
n-C_7H_{16}	15.56	19.65	−19.32	−3.76	−3.76	230
n-C_8H_{18}	15.95	21.14	−19.52	−3.57	−3.59	240
n-C_9H_{20}	16.68	22.38	−20.13	−3.45	−3.46	245
n-$C_{10}H_{22}$	16.90	23.37	−20.24	−3.34	−3.33	250
n-$C_{11}H_{24}$	17.37	24.21	−20.56	−3.19	−3.20	255
n-$C_{12}H_{26}$	18.09	–	−21.20	−3.11	−3.11	260
n-$C_{13}H_{28}$	18.20	25.55	−21.18	−2.98	−3.01	260
n-$C_{14}H_{30}$	18.45	26.13	−21.39	−2.94	−2.95	265
H_2O	25.13	71.99	−19.85	5.28	5.13	2300
CH_3OH	20.51	22.07	−22.51	−2.00	−1.93	875
C_2H_5OH	19.67	21.97	−21.94	−2.27	−2.21	685
n-C_3H_7OH	19.89	23.32	−22.26	−2.37	−2.38	600
n-C_4H_9OH	20.71	24.93	−23.14	−2.43	−2.41	545
n-$C_5H_{11}OH$	21.14	25.36	−23.46	−2.32	−2.35	505
n-$C_6H_{13}OH$	20.56	25.81	−22.91	−2.35	−2.34	470
n-$C_7H_{15}OH$	21.00	–	−23.26	−2.26	−2.30	455
n-$C_8H_{17}OH$	21.22	27.10	−23.45	−2.23	−2.25	435
n-$C_9H_{19}OH$	21.19	27.89	−23.43	−2.24	−2.22	420
n-$C_{10}H_{21}OH$	20.99	28.51	−23.11	−2.12	−2.17	410
n-$C_{11}H_{23}OH$	20.44	–	−22.50	−2.06	−2.08	400

temperature and for a given cavity diameter, ΔG_c depends on the size of solvent molecules, σ, and on the volume packing density of the solvent, ξ [8,9,28,39]. The ΔG_c value increases (a) on decreasing σ for a given ξ; (b) on increasing ξ for a given σ. Such dependence can be readily rationalized by means of geometric arguments [8].

In the n-alkane series the ξ effect is dominant leading to an increase of ΔG_c passing from n-C_6H_{14} ($\xi = 0.501$) to n-$C_{14}H_{30}$ ($\xi = 0.621$). In the n-alcohol series there is a nearly perfect balance between the σ effect and the ξ effect, leading to an almost constancy of ΔG_c along the series. The comparison between n-alkanes and n-alcohols of the same number of heavy atoms points out that the n-alcohol is characterized by: (a) a smaller value of σ (i.e. 5.77 Å for n-pentanol vs. 5.92 Å for n-hexane, and 7.42 Å for n-undecanol vs. 7.53 Å for n-dodecane); (b) a significantly smaller value of v (i.e. 108.24 cm^3 mol^{-1} for n-pentanol vs. 130.52 cm^3 mol^{-1} for n-hexane, and 207.65 cm^3 mol^{-1} for n-undecanol vs. 227.51 cm^3 mol^{-1} for n-dodecane). This reflects in a larger value of ξ (i.e. 0.560 for n-pentanol vs. 0.501 for n-hexane, 0.620 for n-undecanol vs. 0.604 for n-dodecane). Therefore, between n-alkanes and n-alcohols, both the σ effect and the ξ effect play a role in rendering ΔG_c larger in the latter with respect to the former liquids. All this reasoning is grounded on the reliability of the σ values selected for the two series of compounds, whose molecules are non-spherical. In this respect, it is worth noting that the group contributions, used by us to estimate the effective hard sphere diameters of n-alcohols, were determined by Ben-Amotz and Willis by means of a robust procedure [34], without any attempt to fit solvation thermodynamic data.

Water is characterized by the largest ΔG_c value because σ is very small. The size of water molecules is the smallest one among all common solvents (i.e. it practically corresponds to the diameter of an oxygen atom), and this circumstance exaggerates the excluded volume effect associated with cavity creation [8,9,28,39–43]. The comparison between n-alkanes, n-alcohols and water clarifies unequivocally that the effective size of solvent molecules is the main determinant of the ΔG_c magnitude (note that water is characterized by the smallest value of ξ, reminiscent of the open structure of ice).

Usually the surface tension of a liquid, γ, is considered a measure of the work necessary to create a cavity in it [44–46], even though such relationship for molecular-sized cavities is not theoretically founded [47,48]. The experimental values of γ at 25 °C for the different liquids considered in this work are listed in the third column of Table 3 [32]. A comparison between the ΔG_c estimates from SPT and the γ values confirms the existence of a rough relationship, but there are inconsistencies (i.e. along the *n*-alcohol series ΔG_c is practically constant, whereas γ slightly increases).

The E_a values, calculated with the Pierotti's formula, are listed in the fourth column of Table 3. They prove to be large negative in all solvents, but some points emerge. First, in *n*-alkanes and *n*-alcohols, E_a is larger in magnitude than ΔG_c so that $\Delta G^{\bullet}$ is negative (i.e. the xenon-solvent van der Waals interactions overwhelm the excluded volume effect due to cavity creation). Second, E_a in water has a magnitude similar to that in *n*-alkanes, but coupled with the largest value of ΔG_c renders $\Delta G^{\bullet}$ large positive for the solvation of xenon in liquid water. Third, E_a is always larger in magnitude in *n*-alcohols with respect to *n*-alkanes, even though it is practically constant along the *n*-alcohol series, while it is slightly increasing along the *n*-alkane series. This is because both ξ and ε/k are larger for *n*-alcohols than for *n*-alkanes of the same number of heavy atoms. Remember that E_a, according to the Pierotti's formula [30], is proportional to both ξ and ε/k.

As already stated in the Calculation procedure section, the ε/k parameters used for *n*-alkanes and *n*-alcohols do not correspond to literature values, but have been determined in order to have close agreement between $\Delta G_c(\mathrm{SPT})+E_a(\mathrm{Pierotti})$ and experimental $\Delta G^{\bullet}$ data. Actually, our ε/k parameters have to be considered effective energy parameters, that as expected, increase passing from *n*-alkanes to *n*-alcohols, and along each series with the chain length. In fact, such ε/k parameters should be a measure of both the polarizability of the molecule, which in turn is proportional to the size and the number of electrons of the molecule, and of its dipole moment [37]. Moreover, it is worth noting that the ε/k values calculated from gas solubility data, liquid-state properties using a cell theory, gas-phase second virial coefficients and viscosities, show often an erratic scatter that blurs expected trends. For a given species the disagreement between different ε/k estimates may be as large as 20% [30,33,37,38].

The values of the cohesive energy density, the ratio of the internal energy change upon vaporization to the molar volume of the liquid [49], c.e.d. $=(\Delta H_{\mathrm{vap}}-\mathrm{RT})/v$, calculated at 20 °C for each liquid [32], are listed in the last column of Table 3. By comparing such numbers with the SPT estimates of ΔG_c, it is evident that there is no correlation. This is expected by considering that the c.e.d. of the liquid does not enter, directly or indirectly, into the SPT formula to calculate the work of cavity creation [39]. Moreover, the numerical values of ξ and those of the c.e.d. do not suggest the existence of a relationship between such two quantities.

A comparison between the experimental $\Delta G^{\bullet}$ data for xenon and the c.e.d. values are performed in Fig. 1 for *n*-alkanes and Fig. 2 for *n*-alcohols. Fig. 1 indicates that there is correlation between $\Delta G^{\bullet}$ and c.e.d. in the case of *n*-alkanes since both quantities increase with the number of carbon atoms. Fig. 2 indicates that there is no correlation in the case of *n*-alcohols: $\Delta G^{\bullet}$ decreases from methanol up to *n*-butanol and then increases, whereas c.e.d. decreases along the whole *n*-alcohol series. The latter trend may be an indication that *n*-alcohols become more alkane-like on increasing the chain length, whereas the trend of $\Delta G^{\bullet}$ could reflect the tendency of xenon atoms to localize in aliphatic-rich regions avoiding the OH groups. This reliable interpretation does not change the matter; rather it confirms that $\Delta G^{\bullet}$ and c.e.d. are not related in the case of *n*-alcohols (for which, being H-bonded liquids, the c.e.d. should play a significant role). In fact, it is worth noting that, passing from CH_3OH to *n*-$C_{11}H_{23}OH$, the c.e.d. decreases markedly from 875 to 400 J cm^{-3}, but the $\Delta G^{\bullet}$ values change very little (see Table 3).

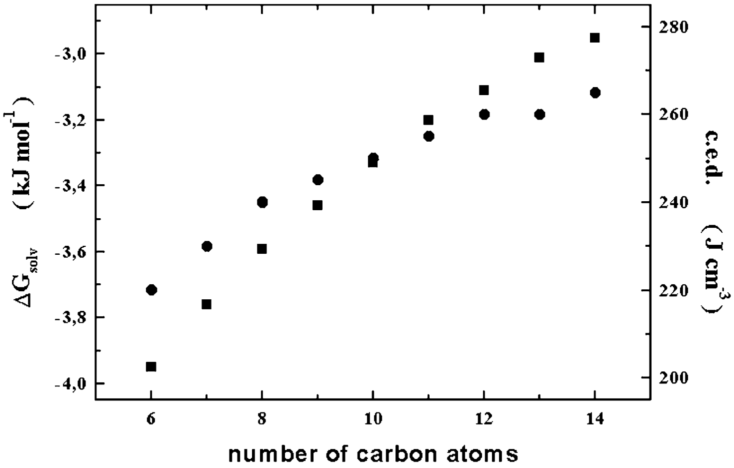

Fig. 1. Plot of the solvation Gibbs energy change for xenon in *n*-alkanes at 20 °C (■), and plot of the cohesive energy density of the same alkanes at 20 °C (●). The data are from [4] and [32], respectively.

This finding is in clash with the claim by Lazaridis [7] that cohesive forces among solvent molecules should be the main determinant for the solubility of non-polar species. Real experimental data do not support the explanation originated by a biased analysis of computer simulations.

On the basis of Eq. (4) and Eq. (5), we determined estimates for $\Delta H_r = \Delta H^{\bullet} - E_a$ and $\Delta S_{nx} = \Delta S^{\bullet} - \Delta S_x$. Such estimates have solely a qualitative significance and are listed in Table 4. It is evident that: (a) ΔH_r is large positive in both *n*-alkanes and *n*-alcohols, whereas it is very small in water; the same statement is true for ΔS_{nx} since Eq. (6) holds; (b) ΔS_x is negative in all solvents, but water is characterized by the largest ΔS_x in absolute value. These findings are absolutely not new. It has been already pointed out that [8,9,28]: (a) the small size of water molecules renders very

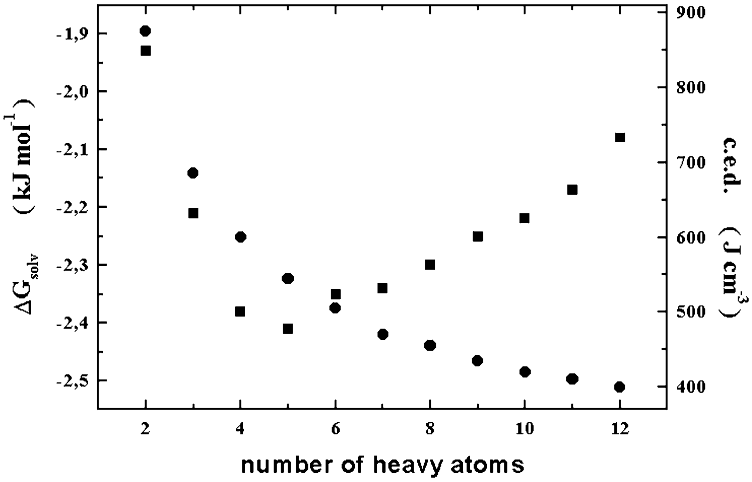

Fig. 2. Plot of the solvation Gibbs energy change for xenon in *n*-alcohols at 20 °C (■), and plot of the cohesive energy density of the same alcohols at 20 °C (●). The data are from [5] and [32], respectively.

Table 4
Estimates of ΔH_r, ΔS_x and ΔS_{nx} for the solvation at 20 °C of xenon in *n*-alkanes, water, and *n*-alcohols. See text for further details

	$\Delta H^{\bullet}$ kJ mol^{-1}	E_a kJ mol^{-1}	ΔH_r kJ mol^{-1}	$\Delta S^{\bullet}$J K^{-1} mol^{-1}	ΔS_x J K^{-1} mol^{-1}	ΔS_{nx} J K^{-1} mol^{-1}
n-C_6H_{14}	−9.23	−18.78	9.55	−18.01	−50.62	32.61
n-C_7H_{16}	−9.52	−19.32	9.80	−19.65	−53.08	33.43
n-C_8H_{18}	−8.86	−19.52	10.66	−17.98	−54.41	36.43
n-C_9H_{20}	−8.61	−20.13	11.52	−17.57	−56.90	39.33
n-$C_{10}H_{22}$	−8.24	−20.24	12.00	−16.75	−57.65	40.90
n-$C_{11}H_{24}$	−7.98	−20.56	12.58	−16.31	−59.25	42.94
n-$C_{12}H_{26}$	−8.02	−21.20	13.18	−16.75	−61.71	44.96
n-$C_{13}H_{28}$	−8.16	−21.18	13.02	−17.57	−62.08	44.51
n-$C_{14}H_{30}$	−7.85	−21.39	13.54	−16.72	−62.94	46.22
H_2O	−18.07	−19.85	1.78	−79.14	−85.72	6.58
CH_3OH	−7.81	−22.51	14.70	−20.05	−69.96	49.91
C_2H_5OH	−8.12	−21.94	13.82	−20.15	−67.10	46.95
n-C_3H_7OH	−8.51	−22.26	13.75	−20.90	−67.85	46.95
n-C_4H_9OH	−8.43	−23.14	14.71	−20.55	−70.65	50.10
n-$C_5H_{11}OH$	−8.29	−23.46	15.17	−20.25	−72.11	51.86
n-$C_6H_{13}OH$	−8.42	−22.91	14.49	−20.75	−70.13	49.38
n-$C_7H_{15}OH$	−8.32	−23.26	14.94	−20.55	−71.64	51.09
n-$C_8H_{17}OH$	−8.45	−23.45	15.00	−21.15	−72.39	51.24
n-$C_9H_{19}OH$	−7.90	−23.43	15.53	−19.45	−72.28	52.83
n-$C_{10}H_{21}OH$	−7.84	−23.11	15.27	−19.35	−71.60	52.25
n-$C_{11}H_{23}OH$	−7.77	−22.50	14.73	−19.40	−69.73	50.33

large ΔG_c and so the magnitude of ΔS_x; (b) water molecules are able to avoid the loss of H-bonds adopting specific orientations around the solute molecule, in view of the tetrahedral H-bonding functionality, and this renders ΔH_r small; (c) both *n*-alkane and *n*-alcohol molecules were not able to restore all the solvent–solvent interactions around a solute molecule and so ΔH_r is large positive.

5. Discussion

We considered solvents whose molecules are highly non-spherical, such as long-chain *n*-alkanes and *n*-alcohols, for which the application of SPT is questionable. The correctness of SPT results is entirely grounded on the reliability of the σ values selected. We are confident that the effective hard sphere diameters determined by Wilhelm and Battino [33], and by Ben-Amotz and Willis [34] are reliable, because: (a) the two procedures adopted are robust; (b) there is good agreement for the liquids considered by both research groups, i.e. *n*-alkanes.

Solvent molecular size is a very important point also to clarify the role played by H-bonds for the solubility of xenon in water and *n*-alcohols. Madan and Lee [50] pointed out that H-bonds bunch-up water molecules well beyond their van der Waals diameter. In fact, $\sigma = 2.8$ Å for H-bonded water molecules, whereas the van der Waals diameter of an oxygen atom is 3.2 Å [50]. The bunching-up effect is operative also in liquid alcohols since the effective diameter of methanol and ethanol is 3.83 and 4.44 Å [34], respectively, whereas the van der Waals diameter is 4.10 and 4.66 Å [51,52], respectively. This holds also for the other *n*-alcohols, even though the bunching-up effect diminishes on increasing the chain length.

In order to single out the role played by H-bonds in determining the magnitude of the work of cavity creation, one should investigate what happens upon turning off H-bonds. This thought transformation has to be done without changing the molar volume of the liquid. The latter is a fundamental requirement, because, to compare ΔG_c in a liquid with H-bonds turned on and in the same liquid with H-bonds turned off, the characteristic number density of the liquid itself has to be preserved [39,50,53]. Indeed, according to a general theorem of statistical mechanics [54],

ΔG_c can be directly calculated by determining the probability that the desired region of space is devoid of solvent molecules from molecular-scale density fluctuations [19,55]. If the molar volume were not considered constant, the number density of water or alcohol molecules in a given box would decrease markedly upon turning off H-bonds, rendering the comparison meaningless.

In the assumption that the molar volume does not change, it is readily verified that ΔG_c becomes larger by considering the van der Waals diameter values as the size of water, methanol and ethanol in SPT calculations, because ξ increases markedly (i.e. ξ passes from 0.383 to 0.572 for water, from 0.438 to 0.537 for methanol, and from 0.473 to 0.547 for ethanol). Specifically, to create, at 20 °C, a cavity suitable to host a xenon atom in water $\Delta G_c = 59.13$ kJ mol^{-1} for $\xi = 0.572$ vs. 25.13 kJ mol^{-1} for $\xi = 0.383$; in methanol, $\Delta G_c = 31.74$ kJ mol^{-1} for $\xi = 0.537$ vs. 20.51 kJ mol^{-1} for $\xi = 0.438$; and in ethanol $\Delta G_c = 27.36$ kJ mol^{-1} for $\xi = 0.547$ vs. 19.67 kJ mol^{-1} for $\xi = 0.473$. Clearly, there is also an increase in the magnitude of the E_a term, but it is markedly smaller than that of ΔG_c. In particular, E_a for xenon passes from -19.85 to -23.60 kJ mol^{-1} in water, from -22.51 to -24.92 kJ mol^{-1} in methanol, and from -21.94 to -23.71 kJ mol^{-1} in ethanol. The above numbers suggest that the H-bonds, by reducing the size of solvent molecules, would render ξ smaller which reflects in smaller ΔG_c estimates [39]. In other words, the existence of H-bonding interactions among solvent molecules could aid the solvation of non-polar solutes [8,39,50]. It is worth mentioning that a similar conclusion was reached independently by Shinoda [56], and Hvidt [57], on the basis of a completely different thermodynamic analysis.

Clearly H-bonds play an important role in determining the density of water and alcohols, together with the size and shape of molecules. To pass from *n*-hexane to *n*-hexanol, an oxygen atom has to be formally added to the *n*-hexane molecule. As a consequence one should expect an increase in molar volume. In contrast the molar volume of *n*-hexanol is less than that of *n*-hexane, 125.59 vs. 130.52 cm^3 mol^{-1} at 20 °C. The numbers listed in the second column of Table 2 show that such a decrease is a general feature. There is no doubt that the formation of H-bonds in *n*-alcohols is the cause of the decrease in molar volume with respect to *n*-alkanes. A similar comparison cannot be done in the case of water. However, since volume expands upon ice formation, one can conclude that H-bond formation increases the molar volume in water. This is a peculiar property of water, due almost certainly to the formation of a three-dimensional network of H-bonds, a feature that is absent in *n*-alcohols for the presence of the alkyl chain [58].

However, even though water and methanol would never be liquid at room temperature without H-bonds, the presence of the latter is not the fundamental factor in controlling the solubility of nonpolar compounds. The qualitative difference between the $\Delta G^\bullet$ values of xenon in water and *n*-alcohols (positive in the former, but negative in the latter) points out that the uniqueness of water for the solubility of non-polar species is not due to H-bonds, but to the smallness of its molecules. In addition, computer simulations have shown that, in both a hard sphere liquid and a Lennard-Jones liquid constituted by spherical particles of the same size of water molecules and characterized by the same number density of liquid water, the values of the work of cavity creation are close to those determined in detailed models of liquid water, taking into full account the presence of H-bonds [50,59,60]. In this respect the results by Gomez and colleagues are illuminating [60].

Ben-Naim and Marcus [61], in analysing the solvation thermodynamics of xenon in *n*-alkanes, suggested an explanation that cannot be considered entirely satisfactory because it does not take into account the excluded volume entropy contribution due to cavity creation. In addition, Ben-Naim [62] wrote the following for the solvation thermodynamics of xenon in *n*-alcohols: 'all we can conclude from the present experimental data is that the average binding energy of xenon to its surroundings is reduced on replacing an environment of *n*-alkane by *n*-alkanol. This reduction is probably due to the decrease in the overall coordination number of the various groups around the xenon atom'. Such statements are not reliable nor correct, as well demonstrated by the results of our analysis.

In this respect it is worth re-stating that: (a) the molecule of *n*-alcohol does not have the same size of the *n*-alkane molecule having the same number of heavy atoms; (b) the *n*-alcohol has a packing density greater than that of the corresponding *n*-alkane. These are fundamental issues, carefully considered by us, but not very appreciated by other authors.

Even Pollack and co-workers [63,64] tried to correlate their solubility data to the sum of the work of cavity creation and the xenon-solvent interaction energy, but their attempt reached limited success. In all probability the main cause was the procedure adopted to obtain effective hard sphere diameters for the various solvents. Pollack and co-workers determined σ by fitting experimental heat of vaporization data with respect to a formula provided by SPT. In this manner they obtained too large σ values for *n*-alcohols (for instance: for CH_3OH, 4.08 vs. 3.83 Å; for n-$C_6H_{13}OH$, 6.27 vs. 6.10 Å; for n-$C_{11}H_{23}OH$, 7.59 vs. 7.42 Å). Actually, Pierotti [30] pointed out that such procedure is not applicable to H-bonded liquids.

Eads [1], on the basis of lattice models for both *n*-alcohols and *n*-alkanes, developed under several simplifying assumptions, suggested that in the first class of solvents the insertion of a nonpolar solute affects both the H-bond entropy and the chain conformational entropy, whereas solely the latter mechanism is operative in the second class of solvents. However, except for a mathematical fitting of the $\Delta G^{\bullet}$ values for xenon in *n*-alcohols and *n*-alkanes, his analysis did not provide any significant results.

In conclusion, the ability of a theoretical approach to rationalize a set of thermodynamic solvation data cannot in itself prove the correctness of the model. However, it has to be noted that the very same theoretical approach has been able to satisfactorily explain most of the data and features constituting the hydrophobicity puzzle [8–15,65–70]. This strengthens the reliability of the present analysis, indicating that the effective size of solvent molecules is the principal factor in determining the xenon solubility.

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