Enthalpies of Interaction of Alkanes and Alkenes with Polar and Nonpolar Solvents

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Abstract: Enthalpies of solution (ΔH_s) of a variety of straight-chain, branched, and cyclic alkanes and alkenes have been measured in methanol, *N,N*-dimethylformamide (DMF), benzene, and cyclohexane. ΔH_s values and enthalpies of evaporation (ΔH_v) have been combined to give enthalpies of transfer from the vapor to each solvent [$\Delta H(\mathrm{v} \rightarrow \mathrm{S})$]. These solvation enthalpies are more exothermic for alkenes than for the corresponding alkanes in transfers to methanol, DMF, and benzene, by about 0.3, 0.6, and 0.3 kcal/mol, respectively, but are similar in transfers to cyclohexane. This suggests that solute-solvent interactions in cyclohexane are mainly dispersion forces. The more exothermic interaction of DMF with alkenes relative to alkanes results mainly from dipole-induced dipole interactions of the solvent dipoles with the highly polarizable π bonds. Only very large steric factors inhibit the approach of solvent dipoles to the π bond, as in *trans*-1,2-di-*tert*-butylethylene, where the alkene-alkane difference in $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ is diminished. $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ of alkanes and alkenes are reduced by nearly 1 kcal/mol for each quaternary carbon atom present, possibly arising from steric hindrance to solvation by dispersion forces.

The thermodynamics of solvation—the interaction of solute molecules with bulk solvents—is important in chemistry for several reasons.[1] In the past decade extensive investigations have been carried out in an attempt to understand the molecular interactions of nonpolar and polar solutes in aqueous systems, which may determine the conformations and physicochemical properties of proteins and model compounds of interest to biological, biomedical, and pharmaceutical sciences.[1-4] However, there are surprisingly few experimental data describing the effects of systematic change in the structure of nonpolar solutes such as aliphatic hydrocarbons in nonaqueous solvents.[5,6] In such systems a large number of complications and unresolved questions regarding a realistic model for the structure of liquid water are avoided. Measurements of these interactions in appropriate systems are needed to provide the basis for both empirical and theoretical generalizations, and also to contribute to an understanding of the intermolecular forces which become operative as gaseous nonpolar molecules interact with solvents.

Important insights into solution reactivity and equilibria have recently resulted from studies of gaseous state behavior in combination with measurements of the interactions of he gaseous species with solvents.[6-8] The effect of solvation on equilibrium constant and rates of reaction was well exemplified in an earlier paper by Arnett[10] on medium effects in the hydrolysis of *tert*-butyl chloride in water-alcohol mixtures. The absence of data on the solvation of alkenes has been a barrier to the correlation of consistent force field calculations for these compounds with heats of hydrogenation in solvents.[9] Recently the necessity of understanding solvation of nonpolar hydrocarbon molecules in nonaqueous solvents has been emphasized in an attempt to explain the phenomenon of "reversed" or "inverted" micelles[5] in the solvents benzene and cyclohexane.

The first step in the investigation of solvation of organic molecules must be to form empirical rules from an extensive set of data. This will permit the estimation of thermodynamic properties of molecules for which the direct experimental determination is not possible, such as unstable intermediates and transition states in various reactions, and compounds which are too insoluble for direct measurement of the heats of solution. Empirical generalizations based on an extensive set of data may also permit the identification of faulty experimental results. Previous investigations of physicochemical properties of tetraalkylammonium ions,[11] aliphatic hydrocarbons,[12] and alcohols[12,13] in polar solvents, and of aromatic compounds in polar and nonpolar solvents[14-16] have indicated a linear dependence of properties on chain lengths, i.e., an approximately constant increment in properties for each addition of a methylene group. Group additivity has also been demonstrated for several functional groups in aliphatic alcohols and aromatic systems. Thus, it is convenient to work with the derived group contribution data, rather than the total molecule, while searching for trends which may suggest dominating underlying molecular interactions.[16]

In this investigation we attempt to study the solvation of an extensive series of alkanes and alkenes, including straight-chain, branched, and cyclic compounds, by measurements of enthalpies of solution (ΔH_s) in a spectrum of four solvents, hydrogen-bonding methanol (MeOH), polar *N,N*-dimethylformamide (DMF), polarizable benzene (Bz), and nonpolar cyclohexane (c-C_6).

We have examined enthalpies of transfer of solutes from the vapor to each solvent to establish trends which provide information on the extent to which polar and dispersion interactions of alkanes and alkenes with each solvent can be affected by steric factors.

Experimental Section

Reagents. The methanol, DMF, and benzene were reagent grade materials stated to contain 0.02-0.05% water, and were obtained from MCB and Fisher. Cyclohexane of 99+% purity was obtained from Aldrich. Methanol was dried with 3A molecular sieve, and the other solvents, with 4A molecular sieve. Solvent samples (70 mL) were transferred by syringe. Most of the hydrocarbons were obtained from Chemical Samples Co., and a few others from Aldrich. All are stated to be 99-99.9% pure, and are liquids at room temperature.

Calorimetry. The principal features of the calorimetric apparatus and procedure have been described previously.[17] The vacuum jacketed calorimeter was immersed in a water bath maintained at 25.00 °C with regulation to ±0.001 °C. Liquid samples were introduced into the calorimeter using 20 or 50 μL syringes with the needle tip immersed in solvent. Sample concentrations were $1\text{-}10 \times 10^{-3}$ M; ΔH_s values were concentration independent within this range and are thus essentially infinite dilution values. The reported values of ΔH_s are averages of 3-6 determinations, and have an estimated accuracy of ±0.05 kcal/mol. The measured values of ΔH_s have been corrected (lowered) by 0.03 kcal/mol to allow for the difference in temperature of the sample in the syringe (22-23 °C) and that of the solvent in the calorimeter. This correction is the average of ΔH_s values measured by introducing each liquid solvent from the syringe into the calorimeter containing the same solvent. After introduction of any liquid sample the new baseline was normally attained within 1-1.5 min. The amplified thermistor signal was recorded at an overall sensitivity of 200 μV (0.02 °C) full scale. A typical heat of solution involved a change of 0.1-0.4 kcal, and a temperature change of 0.003 to 0.01 °C.

Table I. Enthalpies of Solution[a] of Alkanes and Alkenes at 25 °C

Solute	ΔH_s (MeOH)	ΔH_s (DMF)	ΔH_s(Bz)	ΔH_s (c-C_6)
Alkanes				
n-C_5	1.01	1.75	1.07	0.26
n-C_6	1.20	2.03	1.17	0.27
n-C_7	1.39	2.22	1.33	0.39
n-C_8	1.58	2.52	1.44	0.44
n-C_9	1.82	2.83[b]	1.62	0.47[b]
n-C_{10}	2.05	3.23[b]	1.80	0.56[b]
n-C_{12}	2.44	3.79	2.15	0.69
n-C_{14}	2.78	4.41	2.44	0.75
n-C_{16}	3.20	5.05	2.70	0.97
2,2,-Me_2C_4	1.06	2.02	1.15	0.14
2,2,-Me_2C_5	1.22	2.24	1.31	0.23
2,2,4-Me_3C_5	1.36	2.38	1.43	0.20
2-MeC_8	1.66	3.01	1.58	0.35
Et_4C	1.46	2.88	1.55	0.19
2,2,4,4-Me_4C_5	1.39	2.64[b]	1.46	0.07[b]
2,2,5,5,-Me_4C_6	1.61	2.92[b]	1.58	0.19[b]
c-C_5	0.94	1.48	0.69	0.03
c-C_6	1.18	1.84	0.87	0.02
c-C_8	1.54	2.22	0.96	−0.03
Me-c-C_6	1.21	1.98	0.97	−0.01
1.1-Me_2-c-C_6	1.28	2.19	1.09	0.05
cis-1,2-Me_2-c-C_6	1.31	2.15	1.09	0.05
Alkenes				
1-C_5	0.50	1.00	0.62	0.28
1-C_6	0.78	1.33	0.76	0.31
1-C_7	1.01	1.58	0.86	0.35
1-C_8	1.18	1.91	0.97	0.39
1-C_9	1.41	2.20	1.11	0.43
1-C_{10}	1.58	2.45[b]	1.27	0.49[b]
1-C_{12}	2.03	3.09	1.57	0.61
1-C_{13}	2.30	3.39	1.70	0.67
1-C_{14}	2.47	3.61	1.86	0.75
1-C_{15}	2.64	3.93	2.01	0.78
cis-2-C_8	1.19	1.89	0.95	0.43
trans-2-C_8	1.18	1.90	0.97	0.39
cis-4-C_8	1.12	1.89	0.94	0.33
trans-4-C_8	1.18	1.94	0.96	0.39
3-Et-2-Me-2-pentene	1.11	1.64	0.95	0.32
trans-*t*-BuCH=CH-*t*-Bu	1.26	2.86[b]	1.40	0.18[b]
Cyclopentene	0.59	0.80	0.39	0.21
Cyclohexene	0.89	1.23	0.42	0.10
1-Me-cyclohexene	0.93	1.36	0.67	−0.01
Benzene	0.33	0.01	0.00	0.91
1,5-Hexadiene	0.45	0.74	0.43	0.39

[a] kcal/mol ± 0.05. [b] Reference 6.

Results

Enthalpies of solution (ΔH_s) of a variety of alkanes and alkenes have been measured in the solvent methanol, *N,N*-dimethylformamide, benzene, and cyclohexane. The results are presented in Table I. ΔH_s values and published[18,19] values of enthalpies of evaporation at 25 °C have been combined to give enthalpies of transfer for the compounds from the vapor state to each solvent [$\Delta H(v \rightarrow S)$]. These solvation enthalpies are presented in Table II. Values of the familiar quantity the enthalpy of transfer from one solvent to another solvent, $\Delta H_{tr}(S_1 \rightarrow S_2)$, can be derived from either Table I or Table II: $\Delta H_{tr}(S_1 \rightarrow S_2) = \Delta H_s(S_2) - \Delta H_s(S_1) = \Delta H(v \rightarrow S_2) - \Delta H(v \rightarrow S_1)$.

Discussion

We shall discuss phenomenologies obtained from the present study of enthalpies of solution of 22 alkanes and 21 alkenes, with particular emphasis on diagnostic factors such as methylene group additivity, double bond solvation, and steric factors arising from chain branching. A complete discussion of solvent-solute interactions would require that enthalpy data be supplemented by entropy (or free energy) data. It is likely, however, that the resulting conclusions would be similar, because a comparison[16] of transfer enthalpies from MeOH to DMF [$\Delta H_{tr}(M \rightarrow D)$] for aromatic and aliphatic compounds with available ΔG_{tr} values[20] indicates that the latter are largely determined by the enthalpy contribution. Thus the transfer enthalpies are meaningful indications of the change in solvation of a molecule upon transfer from one organic solvent to another. The similarity of enthalpy and free energy terms does not hold in transfers involving water as a solvent, where nonpolar solutes may promote solvent structure, and entropy effects may completely overshadow the enthalpy contribution to the free energy value.[21] Enthalpies of transfer of a given solute from the vapor to various solvents are related to the corresponding free energies in the same way as the enthalpies and free energies of transfer from one solvent to another, and are thus valid indications of relative solvation. Similarly, $\Delta H(v \rightarrow S)$ values for a series of hydrocarbon solutes transferred to a given solvent show a linear enthalpy-entropy relationship.[22] $\Delta H(v \rightarrow S)$ values for the various solutes must again be a valid measure of relative solvation, because these values are linearly related to values of $\Delta G(v \rightarrow S)$.

Enthalpies of Solvation. Heats of solution (ΔH_s) are the resultant of at least three enthalpy terms, the heat of evaporation (ΔH_v) of the solute (or lattice energy, if a solid), the energy of cavity formation (the enthalpy required to separate solvent molecules within the bulk solvent so that the solute molecules can be accommodated), and the enthalpy of interaction of solvent with solute. The first two of these terms are endothermic, and the third is exothermic. ΔH_v can be independently determined as can ΔH_S, but this permits the evaluation only of the sum of the cavity and interaction energies. The sum, often referred to as the solvation enthalpy, is equivalent to the enthalpy of transfer from vapor to solvent [$\Delta H(v \rightarrow S)$]. However, some estimate of the relative hole energies are available from the experimental values of $\Delta H(v \rightarrow S)$, and the observation (discussed below) that nonpolar alkanes, which interact with polar and nonpolar solvents almost entirely by dispersion forces, interact to a rather similar extent with each of the solvents. Differences in $\Delta H(v \rightarrow S)$ for the four solvents studied (with a given alkane solute) reflect principally differences in the solvent cavity energies. Values of $\Delta H(v \rightarrow S)$ become more exothermic in the order DMF < MeOH < Bz < c-C_6, and this probably represents the order of decreasing energy of cavity formation within the solvent structures. The same solvent order is observed for the methylene group increment, $\Delta H(v \rightarrow S)C_{n+1} - \Delta H(v \rightarrow S)C_n$. In practically all cases the heats of solution are endothermic, which indicates that solvation enthalpies [$\Delta H(v \rightarrow S)$] are smaller than enthalpies of evaporation.

The alkanes represent a relatively uncomplicated series of homologs. A number of quantities increase linearly with the number of carbon atoms present, including the heat of evaporation at 25 °C (ΔH_v, 12 kcal/mol/CH_2), the molar volume (16 mL/mol/CH_2), heats of solution in each of the four solvents, and $\Delta H(v \rightarrow S)$. It is likely that the solvent cavity energy and the solute-solvent interaction energy also vary linearly with increasing solute chain length. Since enthalpies of transfer between solvent [$\Delta H_{tr}(S_1 \rightarrow S_2)$] represent differences between ΔH_s or $\Delta H(v \rightarrow S)$ values, these, too, vary linearly with chain length.

Alkenes. The solvation enthalpies of 1-alkenes are more exothermic than for the corresponding alkanes by about 0.3 (MeOH), 0.6 (DMF), and 0.3 kcal/mol (Bz), but are similar in transfers from vapor to cyclohexane (0–0.1 kcal/mol more exothermic for alkanes). This suggests that solute-solvent

Table II. Enthalpies of Transfer[a] of Alkanes and Alkenes from Vapor to Solvents at 25 °C

Solute	$-\Delta H(\mathrm{v} \rightarrow \mathrm{MeOH})$	$-\Delta H(\mathrm{v} \rightarrow \mathrm{DMF})$	$-\Delta H(\mathrm{v} \rightarrow \mathrm{Bz})$	$-\Delta H(\mathrm{v} \rightarrow \mathrm{c\text{-}C_6})$[b]	ΔH_v^{25}[c]
Alkanes					
n-C_5	5.31	4.57	5.25	6.06	6.32
n-C_6	6.34	5.51	6.37	7.27	7.54
n-C_7	7.35	6.52	7.41	8.35	8.74
n-C_8	8.34	7.40	8.48	9.48	9.92
n-C_9	9.28	8.27[d]	9.48	10.63[d]	11.10
n-C_{10}	10.23	9.05[d]	10.48	11.72[d]	12.28
n-C_{12}	12.21	10.86	12.50	13.96	14.65
n-C_{14}	14.22	12.59	14.73	16.25	17.00
n-C_{16}	16.02	14.17	16.52	18.25	19.22
2,2-Me_2C_4	5.56	4.60	5.47	6.48	6.62
2,2,-Me_2C_5	6.53	5.51	6.44	7.52	7.75
2,2,4-Me_3C_5	7.04	6.02	6.97	8.20	8.40
2-MeC_8	9.01	7.66	9.09	10.32	10.67
Et_4C	8.90	7.48	8.81	10.17	10.36
2,2,4,4-Me_4C_5	7.72	6.47[d]	7.65	9.04[d]	9.11
2,2,5,5-Me_4C_6	8.20	6.89[d]	8.23	9.62[d]	9.81
c-C_5	5.88	5.34	6.13	6.79	6.82
c-C_6	6.72	6.06	7.03	7.90	7.90
c-C_8	8.82	8.14	9.40	10.39	10.36
Me-c-C_6	7.24	6.47	7.48	8.46	8.45
1,1-Me_2-c-C_6	7.82	6.91	8.01	9.05	9.10
cis-1,2-Me_2-c-C_6	8.21	7.37	8.43	9.47	9.52
Alkenes					
1-C_5	5.60	5.10	5.48	5.82	6.10
1-C_6	6.65	6.10	6.67	7.12	7.43
1-C_7	7.61	7.04	7.76	8.27	8.62
1-C_8	8.56	7.83	8.77	9.35	9.74
1-C_9	9.55	8.76	9.85	10.53	10.96
1-C_{10}	10.58	9.71[c]	10.89	11.67[c]	12.16
1-C_{12}	12.57	11.51	13.03	13.99	14.60
1-C_{13}	13.52	12.43	14.12	15.15	15.82
1-C_{14}	14.55	13.41	15.16	16.27	17.02
1-C_{15}	15.58	14.29	16.21	17.44	18.22
cis-2-C_8	8.36	7.66	8.60	9.12	9.55
trans-2-C_8	8.35	7.63	8.56	9.14	9.53
cis-4-C_8	8.29	7.52	8.47	9.08	9.41
trans-4-C_8	8.22	7.46	8.44	9.01	9.40
3-Et-2-Me-2-pentene	8.31	7.78	8.47	9.10	9.42
trans-*t*-BuCH=CH-*t*-Bu	(8.54)[e]	(6.94)[d,e]	(8.40)[e]	(9.62)[e]	(9.80)[f]
Cyclopentene	6.08	5.87	6.28	6.46	6.67
Cyclohexene	7.06	6.72	7.53	7.85	7.95
1-Me-cyclohexene	8.09	7.66	8.35	9.03	9.02
Benzene	7.76	8.08	8.12	7.18	8.09
1,5-Hexadiene	6.96	6.67	6.98	7.02	7.41

[a] kcal/mol ± 0.10. [b] The corresponding values in *n*-heptane are *n*-C_5, 6.28; *n*-C_6, 7.51; *n*-C_7, 8.71; *n*-C_8, 9.87; *n*-C_9, 11.06; *n*-C_{16}, 19.06; 2,2-Me_2C_4, 6.57; 2,2,4,4-Me_4C_5, 9.06; c-C_5, 6.74; c-C_6, 7.71; 1-C_6, 7.34; and 1-C_8, 9.66. [c] Heats of evaporation at 25 °C (ref 18, 19). [d] Reference 6. [e] Based on estimated value of ΔH_v (ref 6, 24). [f] Estimated value (ref 6, 24).

interactions in the solutions of hydrocarbons in cyclohexane are mainly dispersion interactions. The stronger interactions of DMF with alkenes, compared to the corresponding alkanes, results mainly from dipole–induced dipole interactions of solvent dipoles with the polarizable π-bond of alkenes. Similar interactions involving methanol are smaller because of the lower solvent dipole moment. The smaller molar volume of alkenes (5.5–6 ml/mol) also favors more exothermic values of $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$, especially in the more highly associated solvents. No special interactions of the π-bond with cyclohexane are possible, and the solvation of alkenes and alkanes (by dispersion interactions) are quite similar.

A comparison of values for *n*-hexane, 1-hexene, and 1,5-hexadiene indicates that the introduction of a second π-bond has an additive effect similar to that of the first.

The effect of introducing a double bond into the cyclohexane ring is similar to that in the open chain alkanes. The effect is somewhat smaller in the cyclopentane–cyclopentene comparison, and somewhat larger in the methycyclohexane–1-methylcyclohexene pair. A comparison of cyclohexene and benzene indicates that the solvation enthalpy of the latter is close to that which would have been predicted by additivity for a six-membered hydrocarbon ring with the equivalent of three double bonds, particularly in the solvents MeOH and DMF.

Cyclic Hydrocarbons. By comparison with $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ for straight-chain hydrocarbons, values for cyclic hydrocarbons are more exothermic by about 0.5 (MeOH), 0.7 (DMF), 0.8 (Bz), and 0.7 kcal/mol (c-C_6). This probably results in part from somewhat more efficient packing of cyclic solutes into the solvent structure, especially in the cyclic solvents benzene and cyclohexane. Several measurements were made with *n*-heptane as solvent. $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ for the straight-chain alkanes and alkenes (C_5–C_{16}) are 0.2–0.8 kcal/mol (0.04–0.05 kcal/mol/CH_2) more exothermic than transfers to the solvent cyclohexane. With cyclohexane as solute the transfer to *n*-heptane is about 0.2 kcal/mol *less* exothermic than transfer to cyclohexane solvent. Thus, the efficiency of solute packing in the solvent structure is of some importance. The smaller molar volume of cycloalkenes and cycloalkanes (> 20 mL/mol) may also make a contribution to more exothermic $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ values, but this may be largely a reflection of the greater packing efficiency in the pure liquid, rather than an independent property.

Steric Effects. Several branched alkanes and alkenes have been investigated in an attempt to explore steric effects on solvation of hydrocarbon molecules and of the π-bond functional group. Steric effects on keto group solvation are also under investigation in this laboratory. Before considering steric factors in the solvation of the π bond, we will first examine saturated molecules. A comparison of data for each of several *n*-alkanes with the corresponding branched isomers indicates that the latter are less strongly solvated in all four solvents than the straight-chain isomer. If the branching occurs twice at the same carbon atom (i.e., a quaternary carbon) the solvation enthalpy is lowered to a value approaching that of the *n*-alkane with one carbon atom fewer. $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ for 2,2-dimethylbutane is less favorable by 0.85 (±0.07 standard deviation) kcal/mol in the four solvents compared to the isomer, *n*-hexane. The mean value for 2,2-dimethylpentane is 0.91 ± 0.10, and for both solutes in all solvents, 0.88 ± 0.09. Similarly, the solvation enthalpy of *n*-nonane is more exothermic by 1.70 ± 0.14 kcal/mol in the four solvents than that of the isomer, 2,2,4,4-tetramethylpentane, which has two quaternary carbon atoms. However, in the case of 2,2,5,5-tetramethylhexane, which also has two quaternary carbons, $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ is 2.14 ± 0.09 kcal/mol less than that of *n*-decane.[24]

It is our hypothesis that differences between $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ values in cyclohexane and polar solvents for nonpolar solutes of low polarizability (alkanes) arise primarily from the greater cavity formation energy (solvent–solvent forces) of the polar solvents, but that *solvent–solute dispersion interactions do not vary greatly from solvent to solvent.* It has recently been concluded[23] from scaled particle theory that, considering molecules as hard spheres, the free energy of the dispersion interactions of a nonpolar solute with water or with a larger nonpolar solvent are of quite similar magnitude. An *n*-alkane and a branched isomer having one or more quaternary carbon atoms have similar molar volumes, and should, therefore, as solutes, cause the expenditure of about the same amount of cavity formation energy in a given solvent (provided that the surface area of the cavity is not a critical factor, as might be true for water). On the other hand, the molecular surface area, and opportunities for effective short range dispersion interactions with solvents, are considerably greater for the *n*-alkane. The observation that the quantity $\Delta H(\mathrm{v} \rightarrow \mathrm{S})_{\mathrm{normal}} - \Delta H(\mathrm{v} \rightarrow \mathrm{S})_{\mathrm{branched}}$ has nearly the same value for MeOH, DMF, benzene, and cyclohexane would be likely only if the dispersion interactions between the *n*-alkane and each of the four solvents were also similar. The intrinsic abilities of isomeric alkanes to undergo dispersion interactions with neighboring molecules are similar. Differences arise according to the effective distances at which the interactions are operative. Clearly, a quaternary carbon atom must be at a considerable distance from surrounding solvent molecules, and its contribution will be small. 2,2,5,5-Tetramethylhexane loses 0.44 ± 0.06 kcal/mol of solvation enthalpy in addition to that attributed to two quaternary carbons; this probably results from the difficulty of close solvent approach to the two interior methylene groups. Values for 2-methyloctane and 2,2,4-trimethylpentane indicate that one methyl group branch (a tertiary carbon atom) results in the loss of about 0.4–0.5 kcal/mol in solvation enthalpy. 3,3-Diethylpentane (tetraethylmethane) is different in having a somewhat smaller loss in $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ (compared to the isomer, *n*-nonane) in spite of the quaternary carbon atom present. This branched isomer is unique among those studied in having a molar volume nearly 5% smaller than the *n*-isomer. This suggests that the symmetrical 3,3-diethylpentane molecules pack very efficiently into the pure liquid, with minimal void spaces, and this should also be true in solution, where the effective solvent–solute distance should be particularly short. Dispersion interactions with this molecule (excepting the quaternary carbon) should be especially favorable. It is noteworthy that the structural features which affect solvation enthalpies also influence enthalpies of evaporation and boiling points of the hydrocarbons. ΔH_{v} is reduced by about 1 kcal/mol for each quaternary carbon atom, relative to the *n*-isomer.

1,1-Dimethylcyclohexane and *cis*-1,2-dimethylcyclohexane show reduced solvation enthalpies (1.24 ± 0.17 and 0.82 ± 0.16 kcal/mol, respectively) relative to cyclooctane. This increment for a quaternary carbon atom is somewhat larger than that for the open chain alkanes, but may be a result of different conformations of the cyclohexane and cyclooctane ring systems, and the molar volume, which is 7% larger for 1,1-dimethylcyclohexane than for cyclooctane.

Steric Hindrance to π-Bond Solvation. It was pointed out previously[6] that solvation enthalpies of alkenes are more exothermic than the values for alkanes by 0.3 (MeOH), 0.6 (DMF), and 0.3 kcal/mol (Bz), but are similar in cyclohexane. This is also valid in the cyclohexane–cyclohexene solute series, but the alkane–alkene differences are smaller in the cyclopentane compounds, and larger in the methylcyclohexane series (0.85, 1.19, 0.87, 0.57 kcal/mol). Thus 1-methylcyclohexene is more strongly solvated than its saturated analog even in the solvent cyclohexane. This possibly results from the flattened configuration at C-1, the resulting closer solvent approach to this atom, and dispersion interactions enhanced by about 0.57 kcal/mol. If this is operative in all four solvents, then the additional solvation of the π bond of 1-methycyclohexene is 0.28 (MeOH), 0.62 (DMF), and 0.30 kcal/mol (Bz), the usual values.

Suppose, though, that bulky groups are attached to an ethylenic double bond, as in *trans*-1,2-di-*tert*-butylethylene (2,2,4,4-tetramethyl-3-hexene). Can these hinder the approach of polar solvent molecules such as DMF, and reduce interactions with the polarizable π bond? ΔH_{v} for this compound has not been reported, but we can instead consider the quantity $\Delta H(\mathrm{v} \rightarrow \mathrm{DMF}) - \Delta H(\mathrm{v} \rightarrow \text{c-C}_6)$ [$= \Delta H_{\mathrm{s}}(\mathrm{DMF}) - \Delta H_{\mathrm{s}}(\text{c-C}_6)$], which is independent of ΔH_{v}. The values[6] are *n*-decane, 2.7; 1-decene, 2.0; *t*-BuCH$_2$CH$_2$-*t*-Bu, 2.7, and *t*-BuCH=CH-*t*-Bu, 2.7 kcal/mol. The relatively unhindered compound 1-decene undergoes 0.7 kcal/mol of π-bond interaction with DMF, which the isomeric hindered alkene cannot. To the combined accuracy of the measurements steric hindrance has completely precluded solvation of the π bond of di-*tert*-butylethylene. For the same four solutes $\Delta H(\mathrm{v} \rightarrow \mathrm{MeOH}) - \Delta H(\mathrm{v} \rightarrow \text{c-C}_6)$ has values of 1.49, 1.09, 1.42, and 1.08 kcal/mol, respectively, which suggests that the small hydroxyl group of methanol is able to approach the π bond of di-*tert*-butylethylene and interact as effectively as with 1-decene.

The octene series (1-octene, *cis*- and *trans*-2-octene, *cis*- and *trans*-4-octene) offers the opportunity to observe the influence of geometric isomerism and terminal vs. nonterminal π bonds on solvation enthalpies. The measured heats of solution are remarkably similar in each solvent, but because of differences in ΔH_{v}, there is some variation in $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$. Values of $\Delta H(\mathrm{v} \rightarrow \text{c-C}_6)$ are less exothermic than that of *n*-octane by 0.13, 0.36, 0.34, 0.40, and 0.47 kcal/mol, respectively. These quantities are reasonably attributed to smaller dispersion interactions of cyclohexane with the alkenes. If, as suggested above, dispersion interactions with a solute are similar in all four solvents, then the alkene dispersion interactions are weaker than the alkane dispersion interactions by about this same amount in each solvent. In MeOH and DMF polar interactions with the π bond may also be operative. If we again consider the quantity $[\Delta H(\mathrm{v} \rightarrow \mathrm{S}) - \Delta H(\mathrm{v} \rightarrow \mathrm{c\text{-}C_6})]$, we find for the solvent MeOH that all five alkenes are more solvated by *polar* interactions than is *n*-octane by 0.32 ± 0.01 kcal/mol, and DMF by 0.56 ± 0.04 kcal/mol. Additional interactions in benzene amount to 0.42 to 0.03 kcal/mol. This implies that π-bond solvation in the octenes is unaffected by *cis* or *trans* geometry, or by location of the bond at a terminal or nonterminal position in the linear carbon chain.

Acknowledgment. This research was supported by the Robert A. Welch Foundation (Grant E-136).

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(24) NOTE ADDED IN PROOF. ΔH_v of 2,2,5,5-tetramethyl-3-hexene (10.22 kcal/mol) and of 2,2,5,5-tetramethylhexane (10.15 kcal/mol) have recently been measured in this laboratory by Dr. L. A. Peacock. $\Delta H(\mathrm{v} \rightarrow \mathrm{S})$ in the four solvents are −8.93, −7.33, −8.79, −10.01, and −8.51, −7.20, −8.54, −9.93 kcal/mol, respectively. This confirms the discussion of steric hindrance to π-bond solvation of the alkene. In the discussion of steric effects on dispersion interactions with quaternary carbons, the solvation enthalpy of 2,2,5,5-tetramethylhexane is now found to be 1.80 ± 0.09 kcal/mol (0.90/C) less than that of *n*-decane, in close agreement with the results for the other three compounds having quaternary carbon atoms.

Rotational Isomerism in Leucine: Proton Magnetic Resonance Study of [γ-^{2}H]Leucine and Thermodynamic Analysis

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Abstract: The ^{1}H NMR spectrum of [γ-^{2}H]leucine has been measured and analyzed to yield values of $^3J(\mathrm{H}^\alpha\text{-}\mathrm{H}^\beta)$ over a range of pD and temperature. These coupling constants have been used to calculate populations of rotational isomers at a range of temperatures in the cationic, anionic, and zwitterionic states. The rotamer in which the side chain is gauche to the amino group and trans to the carboxylate group has the highest population under all conditions. The thermodynamic values for the interconversion of rotational isomers of leucine (about the $\mathrm{C}^\alpha\text{-}\mathrm{C}^\beta$ bond) are derived from experimental populations. For example, the interconversion between the most stable rotamer, in which the isopropyl group is trans to the carboxylate group and gauche to the amino group, and the least stable rotamer, in which the isopropyl group is gauche to both C^α substituents, is described for the leucine anion by $\Delta G°_{25} = 1000$ cal/mol, $\Delta H° = 2200$ cal/mol, and $\Delta S° = 4$ eu. The enthalpies and entropies are discussed in terms of possible intramolecular forces and solvation effects. The free energies are compared to those observed in other amino acids. Differences in rotameric state entropies ranging from 1 to 5 eu are found. These results indicate that isotopic substitution of the γ position for the purposes of spectral simplification is effective and suggests that the technique may be generally useful for other amino acids with γ protons.

The side chains of most amino acid residues of peptides and those of the corresponding free amino acids are considered to rotate freely in solution. The determination of the distributions of rotational isomers of free amino acids is important in identifying some of the forces responsible for the distribution of rotational isomers of amino acids incorporated as residues in peptides. An understanding of these forces should aid our knowledge of the factors that control peptide conformations, for the role of side chains in determining peptide conformation is well recognized.[2]

For the common natural amino acids the predominant values of χ^1, the dihedral angle about the $\mathrm{C}^\alpha\text{-}\mathrm{C}^\beta$ bond, have been assumed to be the three staggered conformations designated rotamers, I, II, and III, as shown in figure 1. An analysis of relative rotamer populations p about the $\mathrm{C}^\alpha\text{-}\mathrm{C}^\beta$ bond can be based upon the values of the coupling constants between vicinal α and β protons, $^3J(\mathrm{H}^\alpha\text{-}\mathrm{H}^\beta)$, inasmuch as the measured values of $^3J(\mathrm{H}^\alpha\text{-}\mathrm{H}^\beta)$ are weighted averages of the coupling manifest in the three rotamers.[3-9] Of the factors contributing to the energies of rotational isomerism in amino acids, steric hinderance and Coulombic interactions have been regarded as the most significant,[4-13] although ion-induced dipole contributions, hydrogen bonding, and interactions with solvent may also be important in particular cases. The energetics of rotational

Cowburn et al. / Rotational Isomerism in Leucine