Molecular Complexes of Water in Organic Solvents and in the Vapour Phase

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1 Introduction

Much of the literature of aqueous solution chemistry concerns solvation, or hydration, phenomena and the interpretation of these phenomena in terms of water-dissolved solute interactions. Previous reviews have treated the interaction of water molecules with ionic salts, polar organic molecules, and compounds of biological importance.[1] The interaction of water molecules with each other, in liquid water and in aqueous solutions, has been investigated by virtually all the applicable methods of physical chemistry.[1b,2] That formidable problems arise in interpreting physical and chemical information about concentrated solutions is evident from the variety of mutually contradictory models which have been proposed for liquid water and aqueous solutions.[3]

It has long been recognised that solute–solute molecular interactions can more readily be elucidated than solvent–solvent or solvent–solute interactions. Fortunately, non-ionic solute species ordinarily follow Henry's law at concentrations up to at least several mole percent.[4] When spectral methods are employed Beer's law generally applies to the individual solute species in a comparable concentration range. Similarly, other specific properties of a dissolved solute are constant in the dilute region: *e.g.*, partial molar volume, energy, and heat capacity; dipole moment; chemical shift in proton magnetic resonance

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[1] *a* R. A. Robinson and R. H. Stokes, 'Electrolyte Solutions', 2nd edn., Butterworths Scientific Publications, London, 1959; *b* J. L. Kavanau, 'Water and Solute-Water Interactions', Holden-Day, Inc., San Francisco, 1964; *c* F. Franks, ed. 'Physico-Chemical Processes in Mixed Aqueous Solvents', Heinemann Educational Books, Ltd., London, 1967; *d* F. Franks and D. J. G. Ives, *Quart. Rev.*, 1966, **20**, 1; *e* R. W. Gurney, 'Ionic Processes in Solution', McGraw-Hill Book Co., Inc., New York, 1953; *f* H. E. Whipple, 'Forms of Water in Biologic Systems', *Ann. New York Acad. Sci.*, 1965, **125**, (2), 249.

[2] C. Pimentel and A. L. McClellan, 'The Hydrogen Bond', W. H. Freeman and Co., San Francisco, 1960.

[3] *a* H. S. Frank and Wen-Yang Wen, *Discuss. Faraday Soc.*, 1957, **24**, 133; *b* H. S. Frank, *Proc. Roy. Soc.*, 1958, *A*, **247**, 481; *c* L. Pauling and R. E. Marsh, *Proc. Nat. Acad. Sci. U.S.A.*, 1952, **38**, 112; *d* W. F. Claussen, *J. Chem. Phys.*, 1951, **19**, 259, 1425; *e* G. Némethy and H. A. Scheraga, *ibid.*, 1962, **36**, 3382, 3401; *f* G. Némethy and H. A. Scheraga, *J. Phys. Chem.*, 1962, **66**, 1773; *g* H. S. Frank and A. S. Quist, *J. Chem. Phys.*, 1961, **34**, 604; *h* M. D. Danford and H. A. Levy, *J. Amer. Chem. Soc.*, 1962, **84**, 3965; *i* J. A. Pople, *Proc. Roy. Soc.*, 1951, *A*, **205**, 163; *j* B. E. Conway, *Ann. Rev. Phys. Chem.*, 1966, **17**, 481; *k* W. A. P. Luck in 'Physico-Chemical Processes in Mixed Aqueous Solvents', ed. F. Franks, Heinemann Educational Books Ltd., London, 1967.

[4] J. H. Hildebrand and R. L. Scott, 'The Solubility of Non-electrolytes', 3rd edn., Dover Publications Inc., New York, 1964, 29.

(^{1}H n.m.r.) studies and many others. Advantage has been taken of the constancy of parameters such as these in numerous studies of molecular interactions in dilute non-electrolyte solutions and in the vapour phase.[1b,2] As early as 1890, Beckmann inferred from cryoscopic data that benzoic acid dissolves in benzene primarily as the dimer;[5] in contrast, there is still considerable discussion regarding the size and nature of the kinetic units in pure associated liquids.

It should be emphasised that investigations of molecular complex formation in dilute solution yield information about the interactions of solute molecules which are solvated, not free as they would be in the vapour phase. For example, the equilibrium $2PhCO_2H = (PhCO_2H)_2$ is strongly affected by solvents;[6] as the medium is changed from vapour to cyclohexane to chloroform, solvation more and more effectively opposes the formation of the dimer. In comparing hydrogen-bonding interactions in dilute organic solutions with those in either aqueous solutions or the vapour state, methods are needed for predicting the influence of solvents on hydrogen-bonding equilibria.

The present Review will summarise what is known about the molecular complexity of water in the vapour state and in dilute solution in non-polar and polar organic solvents; both the behaviour of water as an individual solute and its interactions with other polar solutes will be considered. In order to relate information about the solute properties of water to results for aqueous solutions, a discussion will be given of methods for correlating effects of solvation on molecular complex formation reactions.

2 The Molecular Complexity of Water in Dilute Solution

Water vapour is acknowledged to be virtually ideal, at least at pressures less than *ca.* 90% of saturation in the vicinity of room temperature.[7] High-temperature PVT[8] and i.r. spectral[9] data show, however, that water vapour deviates considerably from ideality and that aggregates exist in significant concentrations at temperatures above 100°, near saturation. Statistical mechanical calculations, used to correlate the second virial coefficient of water vapour and the lattice energy of ice, indicate that the distance of contact between water molecules in the vapour is comparable to the intermolecular distance of water molecules in ice.[10]

Two other methods have provided information related to the association of water in vapour or in an inert matrix. Mass spectrometric measurements of irradiated water vapour have led to the determination of relative concentrations of the hydrates $H^+(H_2O)_n$ and thermodynamic constants for the reactions

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[5] E. Beckmann, *Z. phys. Chem. (Leipzig)*, 1890, **6**, 437.

[6] *a* Y. I'Haya and T. Shibuya, *Bull. Chem. Soc. Japan*, 1965, **38**, 1144; *b* G. Allen, J. G. Watkinson, and K. H. Webb, *Spectrochim. Acta*, 1966, **22**, 807.

[7] E. N. Dorsey, ed. 'Properties of Ordinary Water-Substance', Rheinhold Publishing Corp., New York, 1950, p. 54.

[8] *a* G. S. Kell, G. E. McLaurin, and E. Whalley, *J. Chem. Phys.*, 1968, **48**, 3805; *b* G. S. Kell G. E. McLaurin, and E. Whalley, *ibid.*, 1968, **49**, 2839.

[9] W. A. P. Luck and W. Ditter, *Ber. Bunsengesellschaft Phys. Chem.*, 1966, **70**, 1113.

[10] J. S. Rowlinson, *Trans. Faraday Soc.*, 1949, **45**, 974; J. S. Rowlinson, *ibid.*, 1951, **47**, 120.

$H_2O + H^+(H_2O)_{n-1} = H^+(H_2O)_n$, in which $1 \leqslant n \leqslant 8$.[11] The matrix isolation method (in which an associating vapour is diluted with an inert gas, quickly frozen, and examined spectrally) indicates that water dimers (probably cyclic) and higher polymers exist in a nitrogen matrix at 20K.[12]

In spite of some reports to the contrary,[13] it is generally believed that water dissolves in the aliphatic and aromatic hydrocarbons and CCl_4 primarily as the monomer. Cryoscopic,[14] spectral,[15] vapour pressure,[16] dielectric constant,[17] tracer,[18] and partial molar volume[19] studies indicate that only small concentrations of associated species are present, even at concentrations approaching saturation.

Activity data for water dissolved in several hydrocarbons and chlorinated hydrocarbons are shown in Figure 1. The activity, a_W, is the ratio of water partial pressure to the saturation pressure of pure water at 25°; f_W represents the total or analytical concentration of water. Data are displayed in a log–log plot; the fact that the points for the non-polar solvent systems fall nearly on straight lines drawn with unit slopes implies that water is essentially monomeric in these systems. There is evidence for positive curvature in the plots for aromatic solvents near saturation, but only a few percent of the water molecules appear to be in associated forms.

In slightly polar organic solvents, such as the partially chlorinated hydrocarbons, water is somewhat polymerised. Vapour pressure,[16b,20] i.r.,[21] and ^{1}H n.m.r.[22] data permit calculation of self-association constants for water in 1,2-dichloroethane (DCE), 1,1,2,2-tetrachloroethane (TCE), and 1,2,3-trichloropropane. Spectral evidence indicates that trimeric species are favoured over dimers, although one of the two vapour pressure investigations of water in DCE yielded data which can equally well be interpreted as indicating the presence of dimers or trimers,[20a] and the other vapour pressure results indicate that the most likely associated species are trimers or tetramers.[16b] *Ca.* 8—15% of the dissolved water is associated at saturation at 25°; trimer formation constants vary in the range 3—5 l^2 mole^{-2}. Water activity data for DCE and TCE are included in Figure 1.

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[11] *a* P. Kebarle, S. K. Searles, A. Zolla, J. Scarborough, and M. Arshadi, *J. Amer. Chem. Soc.*, 1967, **89**, 6393; *b* M. G. Inghram and R. Gomer, *Z. Naturforsch.*, 1955, **10a**, 863.

[12] M. Van Thiel, E. D. Becker, and G. C. Pimentel, *J. Chem. Phys.*, 1957, **27**, 486.

[13] *a* M. Gordon, C. S. Hope, L. D. Loan, and Kyong-Joan Roe, *Proc. Roy. Soc.*, 1960, *A*, **258**, 215; *b* T. Ackerman, *Z. phys. Chem.* (*Frankfurt*), 1964, **42**, 119; *c* A. Risbourg and R. Liebaert, *Compt. rend.*, 1967, **264**, *C*, 237.

[14] J. M. Peterson and W. H. Rodebush, *J. Phys. Chem.*, 1928, **32**, 709.

[15] *a* E. Greinacher, W. Luttke, and R. Mecke, *Z. Elektrochem.*, 1955, **59**, 23; *b* S. D. Christian, H. E. Affsprung, J. R. Johnson, and J. D. Worley, *J. Chem. Educ.*, 1963, **40**, 419.

[16] S. D. Christian, H. E. Affsprung, and J. R. Johnson, *J. Chem. Soc.*, 1963, 1896; *b* J. R. Johnson, S. D. Christian, and H. E. Affsprung, *ibid.*, (*A*), 1966, 77.

[17] M. D. Gregory, H. E. Affsprung, and S. D. Christian, *J. Phys. Chem.*, 1968, **72**, 1748.

[18] J. W. Roddy and C. F. Coleman, *Talanta*, 1968, **15**, 1281.

[19] W. L. Masterton and H. K. Seiler, *J. Phys. Chem.*, 1968, **72**, 4257.

[20] *a* W. L. Masterton and M. C. Gendrano, *J. Phys. Chem.*, 1966, **70**, 2895; *b* R. L. Lynch, Ph.D. Dissertation, University of Oklahoma, in preparation.

[21] C. Jolicoeur and A. Cabana, *Canad. J. Chem.*, 1968, **46**, 567.

[22] T. F. Lin, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1965, **69**, 2980.

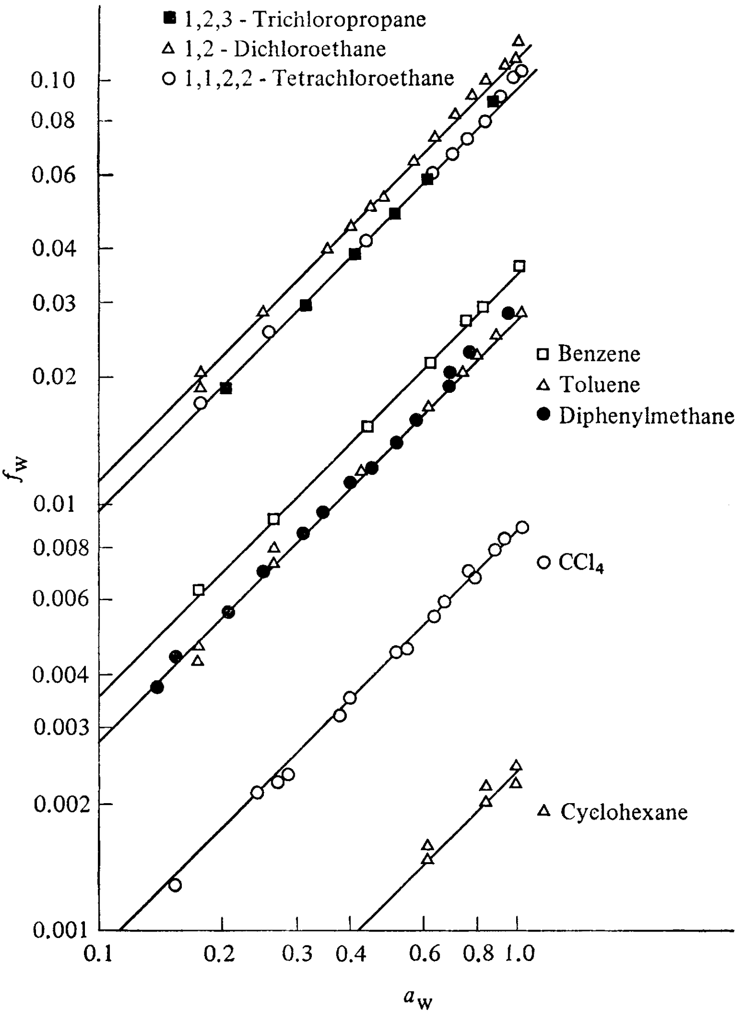

Figure 1 *Formal solubility of water at 25° in* 1,2-*dichloroethane,* 1,1,2,2-*tetrachloroethane, benzene, toluene, cyclohexane—J. R. Johnson, Ph.D. Dissertation, University of Oklahoma,* 1966; *diphenylmethane,* 1,2,3-*trichloropropane—R. L. Lynch, Ph.D. Dissertation, in preparation; carbon tetrachloride—R. D. Grigsby, Ph.D. Dissertation, University of Oklahoma,* 1966

Activity–concentration data have been obtained for water in numerous slightly volatile, polar organic solvents for which the saturation concentration of water exceeds 1 mole percent.[23] The solvents fall into two classes: those for which an assumed monomer–dimer equilibrium is adequate to correlate activity data and those for which monomer–trimer equilibrium provides a better fit. Introducing equilibrium constants for larger aggregates does not improve the correlation. Table 1 summarises formation constants at 25° for the dimer and trimer in several solvents. Included are Henry's Law constants for the water monomer, K_H; the approximate formal concentration of water at saturation in each solvent, f_W^0; and the approximate percent of the dissolved water in associated forms at saturation, r. As K_H decreases and f_W^0 increases, the association constants tend to decrease, but in spite of this, r is generally greater in the more reactive solvents. Thus, mass action is more effective in promoting association than is solvation effective in preventing it.

Table 1 *Solubility and association of* H_2O *in organic solvents*

Solvent	K_H (torr l mole^{-1})	f_W^0 (mole l^{-1})	r	K_{W_2}* (l mole^{-1})	K_{W_3}* (l^2 mole^{-2})	*Ref.*
Tributylamine	504·0	0·0522	10		16·11	23
NN-Dimethylaniline	435·1	0·0668	19		26·39	23
1,1,2,2-Tetra-chloroethane	256·2	0·1010	8·4		3·60	16*b*
1,2-Dichloroethane	218·8	0·1262	14·2		4·60	16*b*
Nitrobenzene	182·3	0·1666	21	1·054		23
Dibutylphthalate	128·3	0·2440	20	0·68		23
NN-Dimethylcyclo-hexylamine	51·9	1·5841	75		4·92	23
Methylphenylketone	43·6	0·9840	41		0·776	23
Cyclohexanone	21·8	2·4590	50		0·289	23
Aniline	18·1	2·9240	50		0·194	23
Benzyl alcohol	11·9	3·812	35·5		0·046	23

*Formation constant for associated species from monomeric water.

Little is known about water association in polar solvents which are completely, or nearly completely, miscible with water. Approximate values of dimerisation constants for water in several proton-accepting solvents have been inferred from limiting slopes of plots of proton chemical shift *vs.* water concentration.[24] To explain proton-exchange rate data for water in proton-accepting solvents, it has been proposed that solvated cyclic trimers are present at higher water concentrations, and that protons transfer within the ring by a concerted mechanism.[24b] ^{1}H N.m.r. and molar volume data are available for the system H_2O–ethylene oxide (EO) at 0° over the entire concentration range.[25] Water protons tend to bond to lone-pair electrons of other water molecules, in preference to the

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[23] J. R. Johnson, S. D. Christian, and H. E. Affsprung, *J. Chem. Soc.* (*A*), 1967, 1924.

[24] *a* G. Mavel, *Compt. rend.*, 1959, **248**, 1505; *b* J. R. Holmes, D. Kivelson, and W. C. Drinkard, *J. Amer. Chem. Soc.*, 1962, **84**, 4677.

[25] D. N. Glew, H. D. Mak, and N. S. Rath, *Canad. J. Chem.*, 1967, **45**, 3059.

EO oxygen lone-pair electrons. The linear water trimer appears to be an important associated species at mole fractions of water less than *ca.* 0·11. At higher water concentrations, chain-branching, cross-linking, and formation of cyclic aggregates predominate, until ultimately the three-dimensional water networks characteristic of pure liquid water are formed. Density and Raman spectral data for H_2O–acetic acid have been interpreted by assuming the presence of water dimers as well as hydrated acid species.[26]

3 Water–Polar Solute Interactions in Dilute Solution

Early evidence for polar solute–water interactions was provided by Bödtker, who noted that the solubility of water in ether increased as oxalic acid was added;[27] he surmised that a compound forms between water and the acid. Before 1940 there were several reports of the effect of dissolved water on results of cryoscopic,[28] partition,[29] and solubility[29b,30] studies of polar organic molecules.[31] In particular, Szyszkowski[29b] calculated dimer formation constants for several aromatic carboxylic acids in benzene from partition and acid solubility data; he showed that the apparent association constant, calculated by ignoring the effect of dissolved water, was ordinarily considerably smaller than the corrected constant, in one case by a factor of nearly 1:50. Later investigations[28b,32] indicated that hydrates of both the acid dimer and the monomer are probably present in wet benzene solutions of chlorinated aliphatic carboxylic acids. Lassettre[31] reviewed the early partition results and cautioned that the technique should not be employed for determining association constants unless corrections are made for water–polar solute interactions in the organic phase. In spite of this warning there have been numerous recent attempts to use the uncorrected partition method to study self-association; for certain systems, however, reasonable formation constants are obtained with the method,[33] presumably owing to a fortunate cancellation of errors.

The techniques employed to investigate molecular complexes of water in dilute solution resemble those usually employed in hydrogen bond studies.[2] However, complications arise in using solvents such as benzene, carbon tetrachloride, and the aliphatic hydrocarbons, which dissolve less than 0·5 mole

[26] J. J. Kipling, *J. Chem. Soc.*, 1952, 2858.

[27] E. Bödtker, *Z. phys. Chem. (Leipzig)*, 1897, **22**, 505.

[28] *a* M. Rozsa, *Z. Elektrochem.*, 1911, **17**, 934; *b* R. P. Bell and M. H. M. Arnold, *J. Chem. Soc.*, 1935, 1432.

[29] *a* S. Horiba, *Mem. Coll. Sci. Univ. Kyoto*, 1914, **1**, 49; *b* B. U. Szyszkowski, *Z. phys. Chem. (Leipzig)*, 1928, **131**, 175.

[30] *a* G. N. Lewis and G. H. Burrows, *J. Amer. Chem. Soc.*, 1912, **34**, 1515; *b* L. A. K. Staveley, J. H. E. Jeffes, and J. A. E. Moy, *Trans. Faraday Soc.*, 1943, **39**, 5; *c* E. Cohen and W. D. J. van Dobbenburgh, *Z. phys. Chem. (Leipzig)*, 1925, **118**, 37; *d* E. Cohen and S. Miyake, *Verlag Akad. Wetenschapen Amsterdam*, 1925, **34**, 933; *e* J. H. Hildebrand, *Science*, 1936, **83**, 21.

[31] E. N. Lassettre, *Chem. Rev.*, 1937, **20**, 259.

[32] R. P. Bell, *Z. phys. Chem. (Leipzig)*, 1930, **150**, *A*, 20.

[33] *a* M. Davies and H. E. Hallam, *J. Chem. Educ.*, 1956, **33**, 322; *b* M. Davies and D. M. L. Griffiths, *Z. phys. Chem. (Frankfurt)*, 1954, **2**, 352; *c* M. Davies, P. Jones, D. Patnaik, and E. A. Moelwyn-Hughes, *J. Chem. Soc.*, 1951, 1249.

percent water. It is difficult to prepare homogeneous solutions of water in a non-polar solvent by direct mixing of the liquid components; even when considerably less water is added than is required to saturate the solvent, vigorous stirring for 24 h will not promote dissolution of all the drops.[17] This problem may account for some of the discrepancies among literature reports on the properties of water in dilute solution.[17,19] The solute isopiestic method has been employed to prepare homogeneous solutions of water in organic solvents at known water activities; vapour phase equilibration of benzene or CCl_4 with an aqueous solution of known water activity occurs within only a few hours, even without stirring and without removing air from the equilibrator vessel.[15b,16a] Manometric techniques are convenient for measuring activities of water in the very dilute solution range (formal water concentrations 0·00005 to 0·00300); one method incorporates a mercury-covered, sintered-glass disc inlet valve formerly used in mass spectrometer systems.[23,34]

A. Treatment of Data.—Simultaneous partition, water solubility, and isopiestic measurements for a given organic solvent–polar solute system have been used to investigate hydration equilibria. To illustrate the rationale of the method, we shall outline the computations involved in studying a hypothetical compound which is present in the organic phase as both monomers and dimers (hydrated and unhydrated).

Assume that the solute, A, is distributed between water and the organic solvent and that the concentration of A in the aqueous phase, $c_A{}^w$, is directly proportional to the activity of A. The analytical or formal concentration of A in the organic phase may be written:

$$f_A = c_A(1 + K_{11}c_W + K_{12}c_W{}^2 + \ldots) + 2c_A{}^2(K_{20} + K_{21}c_W + K_{22}c_W{}^2 + \ldots) \qquad (1)$$

or

$$f_A = K_D c_A{}^w(1 + K_{11}c_W + K_{12}c_W{}^2 + \ldots) + 2K_D{}^2 c_A{}^{w2}(K_{20} + K_{21}c_W + K_{22}c_W{}^2 + \ldots) \qquad (2)$$

where c_A and c_W are concentrations of monomeric A and water (W), respectively, in the organic phase; K_{20} is the equilibrium constant for formation of the unhydrated dimer from the unhydrated monomers; $K_{11}, K_{12}, \ldots, K_{21}, K_{22}, \ldots$ are formation constants for the hydrates AW, $AW_2, \ldots$, A_2W, A_2W_2, *etc.* from monomers of A and W; and $K_D = c_A / c_A{}^w$ is the distribution constant for monomeric A between water and the organic phase. The concentration of water monomer, c_W, may be calculated from the expression $c_W = a_W c_W{}^0$, where $c_W{}^0$ is the limiting value of c_W at unit water activity. Henry's Law is assumed to apply to each hydrated and unhydrated species individually. The measured concentration of water in the organic phase is greater in the presence of A than it would be if only water and the solvent were present; the excess water solubility, Δf_W, is due to the formation of hydrated species and may be expressed as:

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[34] *a* A. A. Taha, R. D. Grigsby, J. R. Johnson, S. D. Christian, and H. E. Affsprung, *J. Chem. Educ.*, 1966, **43**, 432; *b* E. E. Tucker, Ph.D. Dissertation, University of Oklahoma, 1969.

$$\Delta f_{\mathrm{W}} = c_{\mathrm{A}}(K_{11}c_{\mathrm{W}} + 2K_{12}c_{\mathrm{W}}^2 + \dots) + c_{\mathrm{A}}^2(K_{21}c_{\mathrm{W}} + 2K_{22}c_{\mathrm{W}}^2 + \dots) \qquad (3)$$

Equations (1) and (3) apply equally to the partition systems or to solutions of A in solvent equilibrated at known reduced water activities. If sets of distribution data (f_{A}, $c_{\mathrm{A}}^{\mathrm{w}}$, c_{W}) and water solubility data (f_{A}, c_{W}, Δf_{W}) are available, K_{20} and the hydration constants may be inferred by non-linear least-squares analysis.[35] Initial values of K_{D}, K_{20} and the hydration constants are chosen and equation (1) is solved for each point to yield trial values of c_{A}. The mean square deviation, E, given by:

$$E = \sum_{\text{all points}} W_{\mathrm{S}}(\Delta f_{\mathrm{W}}^{\text{exptl.}} - \Delta f_{\mathrm{W}}^{\text{calc.}})^2 + \sum_{\substack{\text{partition}\\ \text{data points}}} W_{\mathrm{P}}(f_{\mathrm{A}}^{\text{exptl.}} - f_{\mathrm{A}}^{\text{calc.}})^2 \qquad (4)$$

is then calculated, using expressions (3) and (2) to obtain $\Delta f_{\mathrm{W}}^{\text{calc.}}$ and $f_{\mathrm{A}}^{\text{calc.}}$, respectively. (The weight factors, W_{S} and W_{P}, are determined from the precision of the water solubility and partition data, respectively.) After each calculation of E, the procedure is repeated with a set of modified values of all the constants, generated with a numerical optimum-seeking technique. The process is continued until the absolute minimum in E and the corresponding least-squares values of the constants are located. Standard errors in the constants may be calculated by the method of Sillén.[36] Ordinarily no more than 2 or 3 hydration constants are needed to fit a given collection of data.

Equations similar to (1—4) may be developed to fit i.r. and ^{1}H n.m.r. data for systems in which hydrates of polar solutes are present.[22,37] It is commonly assumed that both Henry's Law and Beer's Law apply individually to solute species and that the proton chemical shift of each species is concentration independent.[38] Important qualitative information about hydrated species can be obtained from spectral data, even when it is not possible to calculate hydrate formation constants. Bands in the 3 μm region have been attributed to 1:1 and 2:1 hydrates (B · · · H—O—H and B · · · H—O—H · · · B) of organic bases (B) in CCl_4.[39] The relative strengths of water–polar solute interactions have been inferred from proton chemical shifts[24] and OH-stretching frequencies[15a,40] of water dissolved in hydrogen-bonding solvents.

B. Thermodynamic Results for Water–Polar Solute Complexes.—Before the past decade, not much was known about specific molecular complexes of water.

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[35] R. Van Duyne, S. A. Taylor, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1967, **71**, 3427.

[36] L. G. Sillén, *Acta Chem. Scand.*, 1964, **18**, 1085.

[37] *a* F. Takahashi and N. C. Li, *J. Amer. Chem. Soc.*, 1966, **88**, 1117; *b* J. D. Worley, Ph.D. Dissertation, University of Oklahoma, 1964; *c* A. Fratiello and D. C. Douglas, *J. Mol. Spectroscopy*, 1963, **11**, 465; *d* J. M. Sorensen, Ph.D. Dissertation, Case-Western Reserve University, 1967.

[38] N. Muller and P. Simon, *J. Phys. Chem.*, 1967, **71**, 568.

[39] *a* S. C. Mohr, W. D. Wilk, and G. M. Barrow, *J. Amer. Chem. Soc.*, 1965, **87**, 3048; *b* P. Saumagne and M. L. Josien, *Bull. Soc. chim. France*, 1958, 813.

[40] E. D. Becker, *J. Chem. Phys.*, 1959, **31**, 269.

This section summarises recent research which has yielded thermodynamic parameters for specific complexes of water with carboxylic acids, phenols and alcohols, ketones and ethers, amines, amides, and various poly-functional polar solutes.

Since 1960, several specific complexes of water (W) with monocarboxylic acids (A) have been reported, based on spectral,[41] vapour density,[42] vapour pressure,[43] and partition–water solubility[44] measurements. I.r. spectra of moist CCl_4 solutions of three highly acidic halogenated acetic acids and three weaker aliphatic acids have been obtained in the acid C=O stretching and the water hydroxyl-stretching regions.[41] At very low water activities new bands, attributed to AW and A_2W, appear in both spectral regions. Formation constants for AW and A_2W, as well as AW_2, have been reported.[43,44] The 1:2 and 2:1 complexes of benzoic acid and salicyclic acid in several non-polar and slightly polar solvents have formation constants (from the monomers) of the order of several hundred in l^2 $mole^{-2}$ units. A vapour phase constant at 25° of 0·010 $torr^{-2}$ (equivalent to $3{\cdot}5 \times 10^6$ l^2 $mole^{-2}$) has been obtained for the formation of the dihydrate of trifluoroacetic acid.[42] Values of the formation constant for AW for aromatic acids, in CCl_4, benzene, and 1,2-dichloroethane, are in the range 1—12 l $mole^{-2}$. The large magnitudes of the A_2W and AW_2 formation constants, as well as the acid self-association constants, indicate the difficulty of determining accurate parameters for the AW formation constant; *i.e.*, it is difficult to reach a concentration region in which a major portion of the hydrated acid is in the form of the 1:1 complex. Similar but less severe complications arise in studies of hydrates of other polar solutes having both donor and acceptor sites.

Relatively little is known about specific hydrates of phenols and alcohols, in spite of the fact that numerous studies have been made of aqueous and moist organic solutions of these compounds.[1d,2] Hydration studies are complicated by a lack of reliable values of self-association constants. Early partition[29a,45] and cryoscopic[28a] investigations showed that strong interactions occur between dissolved water and phenols or alcohols in organic solvents. Partition–water solubility data at unit water activity,[46] and water solubility data at reduced water activities,[46b–d] have indicated the presence of PW, P_2W, and PW_2 in solutions of phenol (P) and water (W) in several organic solvents. In benzene, 1,1,2,2-tetrachloroethane, and 1,2-dichloroethane, the trimeric hydrates are more

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[41] J. de Villepin, A. Lautié, and M. L. Josien, *Ann. Chim.*, 1966, **1**, 365.

[42] S. D. Christian, H. E. Affsprung, and C. Ling, *J. Chem. Soc.*, 1965, 2378.

[43] G. O. Wood, D. D. Mueller, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1966, **70**, 2691.

[44] *a* R. Van Duyne, S. A. Taylor, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1967, **71**, 3427; *b* R. Van Duyne, Ph.D. Dissertation, University of Oklahoma, 1969.

[45] *a* R. D. Vold and E. R. Washburn, *J. Amer. Chem. Soc.*, 1932, **54**, 4217; *b* E. R. Washburn, V. Hnizda, and R. D. Vold, *ibid.*, 1931, **53**, 3237; *c* E. R. Washburn and H. C. Spencer, *ibid.*, 1934, **56**, 361.

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important than PW and account for most of the bound water. In CCl_4 solutions of phenol, Δf_W rises rapidly with increase in phenol concentration and a_W, indicating the presence of hydrates involving at least 2 water molecules and several phenol molecules.[46d] Partition and water solubility data for alcohols and l-naphthol in toluene indicate that 2:1 alcohol–water and 1:1 naphthol–water species are probably present,[47] although the stoicheiometry of water in these complexes has not been confirmed by solubility measurements at reduced water activities. Similar data for l-decanol (D) provide evidence for the existence of DW in 1,2-dichloroethane and D_2W in iso-octane.[48] Conductance measurements for solutions of several nitro-substituted phenols and water ($f_W > 0{\cdot}3$ mole l^{-1}) dissolved in acetonitrile indicate that the phenols form hydrates PW_n, in which n is one greater than the number of nitro-groups in the substituted phenol.[49]

Ethers and ketones are not highly associated in organic solvents at concentrations less than several tenths molar. Therefore, the analysis of hydration data for these compounds does not involve large corrections for solute self-association. The acetone monohydrate in 1,2-dichloroethane has been investigated by dielectric,[50] ^{1}H n.m.r.,[22] and partition–water solubility[22] methods; a formation constant of 0·85 l mole^{-1} at 25° and a dipole moment of 3·4 D have been reported for the complex. The dipole moment of the hydrate nearly equals the vector sum of the moments of water and acetone. Table 2 lists formation constants for the ketone (B)–water complexes BW and B_2W in CCl_4 and cyclohexane at 25°.[37a,51] Hydration equilibria of ethers have been investigated by ^{1}H n.m.r. measurements;[37a,38] constants for dioxan and tetrahydrofuran complexes are

Table 2 *Equilibrium constants for formation of hydrates of ketones, ethers, and dimethylsulphoxide at* 25°

Ketone or Ether	*Solvent*	K_{BW} (l mole^{-1})	K_{B_2W}* (l mole^{-1})	*Ref.*
Cyclopentanone	CCl_4	2·2	1·4	51*b*
Acetone	CCl_4	2·4	1·6	51*a*
2,3-Butanedione	CCl_4	1·2	2·0	51*a*
Acetylacetone	CCl_4	1·5	0·86	51*a*
Acetone	Acetone–cyclohexane		0·08†	37*a*
Tetrahydrofuran	THF–cyclohexane		0·13†	37*a*
Dioxan	Cyclohexane	1·3†	0·85†	38
Dimethylsulphoxide	Dimethylsulphoxide–CCl_4		0·26†	52

*For the reaction $BW + B = B_2W$.
†Converted from mole fraction units and interpolated to 25°

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[48] D. J. Turner and R. M. Diamond, *J. Phys. Chem.*, 1968, **72**, 3504.

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[50] T. F. Lin, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1967, **71**, 1133.

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included in Table 2 as well as the B_2W formation constant for dimethyl sulphoxide.[52] ΔH for the addition of the first water to acetone is *ca.* $-3{\cdot}5$ kcal/mole,[38] whereas values in the range $-1{\cdot}6$ to $-2{\cdot}8$ kcal/mole have been obtained for addition of a second base unit to the monohydrates of ketones and ethers[37a,38] and dimethyl sulphoxide.[52] Presumably B_2W has the bridged structure, $B \cdots HOH \cdots B$, proposed earlier on the basis of i.r. results.[39] It is obvious that weaker hydrogen bonding is involved in ketone– or ether–water complexes than in hydrates of phenols and carboxylic acids. The 1:1 complex formation constants reported in Table 2 are comparable in magnitude to those for complexes of similar bases with alcohols.[53]

Considerable effort has been expended in studying the hydration of amines, amides, and biologically important compounds containing the amide group and peptide linkages. The fact that amines are known to form strong hydrogen bonds with proton donors of many types and the tendency of amines to obey Henry's Law throughout considerable ranges of concentration have made these compounds logical choices as solutes in hydration studies. There is spectral evidence for the aromatic amine (D) hydrates DW and D_2W in organic solvents,[39,54] although some aliphatic amines apparently form DW, but not D_2W, owing to the strong inductive effect of the polar $D \cdots W$ hydrogen bond.[55] Table 3 summarizes formation constants for DW, DW_2, and D_2W (from the monomers) as well as selected values of these parameters for methanol–amine complexes. Dipole enhancements for 1:1 complexes of hydroxylic proton donors and amines have been reported; the excess of the complex dipole moment over the vector sum of the donor and acceptor moments ranges from less than 0·3 D for aliphatic alcohol complexes with pyridine and triethylamine[56] to 0·4—1·3 D for aliphatic amine complexes with water[17,57] and phenols.[58] I.r. spectral evidence for the separation of charge in various n-propylamine–proton donor complexes[59] complements the dielectric results; the amine $N-H$ and the $OH \cdot\cdot N$ stretching bands are significantly altered in complexes with proton donors having pK_a values $<$ *ca.* 10. Nevertheless, the charge separations indicated by the dipole moment enhancement for water–amine complexes are unexpectedly large compared to those reported for the aliphatic alcohol–amine complexes.

Table 4 gives formation constants for the amide (M) hydrates MW and M_2W. Difficulties again arise in treating hydration data for lactams[60] and acetamides[61] in that accurate self-association constants are required in calculating hydrate

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[52] S. F. Ting, S. Wang, and N. C. Li, *Canad. J. Chem.*, 1967, **45**, 425.

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[59] T. Zeegers-Huyskens, *Spectrochim. Acta*, 1965, **21**, 221.

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[61] *a* R. D. Grigsby, Ph.D. Dissertation, University of Oklahoma, 1966; *b* R. D. Grigsby, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1968, **72**, 2465.

Table 3 *Equilibrium constants for formation of amine–water and amine–methanol complexes at 25°*

Amine	*Solvent*	K_{DW} (l mole^{-1})	K_{DW_2}* (l^2 mole^{-2})	K_{D_2W}* (l^2 mole^{-2})	*Ref.*
	Water Complexes				
Diethylamine	Hexadecane	11·0			*a*
Diethylamine	Diphenylmethane	8·5			*a*
Diethylamine	Benzyl Ether	2·9	9·7		*a*
Pyridine	Cyclohexane	5·3	6·8		*b*
Pyridine	CCl_4	2·6	18·3	3·0	*b*
Pyridine	Toluene	1·5	7·3	1·6	*b*
Pyridine	Benzene	1·2	6·3	1·5	*b*
Pyridine	1-2 Dichloroethane	1·0	1·0	1·0	*b*
Cyclohexylamine	Benzene	6·7			*c*
N-Methylcyclohexylamine	Benzene	5·3			*c*
NN-Dimethylcyclohexylamine	Benzene	3·9			*c*
Triethylamine	Benzene	3·5			*c*
Triethylamine	Cyclohexane	7·0			*c*
Triethylamine	Toluene	3·7			*c*
	Methanol Complexes†				
Diethylamine	Hexadecane	8·7	300		*a*
Diethylamine	Diphenylmethane	4·8	22·5		*a*
Diethylamine	Benzyl Ether	2·7	8·5		*a*
Pyridine	CCl_4	2·3			*d*
2-Picoline	CCl_4	3·1			*d*
2-Ethylpyridine	CCl_4	2·6			*d*
2-Isopropylpyridine	CCl_4	1·7			*d*
2-t-Butylpyridine	CCl_4	1·4			*d*
2,6-Lutidine	CCl_4	4·4			*d*
2-Ethyl-6-methylpyridine	CCl_4	4·1			*d*
2,6-Diethylpyridine	CCl_4	3·0			*d*
2,6-Di-isopropylpyridine	CCl_4	0·8			*d*

[a] ref. 34*b*; [b] ref. 74*b*; [c] ref. 17; [d] T. Kitao and C. H. Jarboe, *J. Org. Chem.*, 1967, **32**, 407.
*For the reactions from the monomers
†In the case of the methanol systems, K_{DW} and K_{DW_2} indicate formation constants for the 1 : 1 and 1 : 2 amine – methanol complexes respectively.

Table 4 *Equilibrium constants for formation of amide hydrates at 25°*

Amide	*Solvent*	K_{MW} (l mole^{-1})	K_{M_2W}* (l mole^{-1})	*Ref.*
2-Pyrrolidone	CCl_4	9·0	9	*a*
N-Methyl-2-pyrrolidone	CCl_4	9·2	39	*b*
N-Methyl-2-pyrrolidone	Benzene	6·0	4·3	*b*
N-Methyl-2-pyrrolidone	1,2-Dichloroethane	2·4		*b*
NN-Dimethylacetamide	1,2-Dichloroethane	2·2		*b*
N-Methylacetamide	CCl_4	11		*c*
NN-Dimethylacetamide	*NN*-Dimethylacetamide–cyclohexane		0·3	*d*

[a] ref. 37*b*; [b] ref. 60*b*; [c] ref. 61*b*; [d] ref. 37*a*
*For the reaction $MW + M = M_2W$

formation constants. *N*-Methylacetamide (M) in CCl_4, for example, associates extensively in dilute solution to form chain polymers. The formation constant for the reactions $M + W = MW$, $M_2 + W = M_2W, \ldots M_n + W = M_nW$ appears to be nearly constant (*ca.* 11 l mole^{-1}) and independent of chain length, which suggests that each unit in the chain contributes an equal number of nearly equivalent hydration sites.[61b] 2-Pyrrolidone is extensively dimerized in CCl_4 and the equilibrium constant for the reaction $M + W = MW$ nearly equals that for $M_2 + W = M_2W$. The species M_2W probably consists of water attached to the cyclic 2-pyrrolidone dimer; whereas, *N*-methyl-2-pyrrolidone, which associates to a lesser degree, probably forms the bridged hydrate $M \cdots W \cdots M$.[60b]

Several reviews have treated the extensive literature on the energetics, dynamics, and structure of water molecules bound to proteins and other macromolecules;[1b,1f,62] no attempt will be made to summarise research in this area. In the context of the present Review, a series of studies of the interactions of water with various polar groups of deoxyribonucleic acid (DNA)[63] and its molecular constituents[64] is of interest. I.r. spectra of thin films of Na and Li salts of DNA have been measured at various water activities (relative humidities).[63b] As a_W increases from 0 to 0·60, first the PO_2^- group and then the POC and COC oxygens become hydrated. At a water activity of *ca.* 0·65, a hydration shell of 5 or 6 water molecules probably surrounds the phosphate group; above $a_W = 0{\cdot}65$ the C=O and ring nitrogen atoms become hydrated. Polarised i.r. and u.v. spectra indicate that between $a_W = 0{\cdot}55$ and 0·75 the structure of DNA changes from the disordered lower humidity form to the ordered *B* configuration, in which the base pairs become stacked one above another and oriented perpendicular to the axis of the helix.[63c] Gravimetric results for numerous constituents of DNA show that most of the solid purines and pyrimidines and their nucleosides and nucleotides are little hydrated at water activities $<0{\cdot}93$, whereas several salts of the nucleotides and other compounds containing the ionic phosphate group absorb >20 molecules of water per molecule of compound at $a_W = 0{\cdot}93$. These studies underline the importance of the ionic phosphate groups as sites for hydration in nucleotides and polynucleotides.[64]

Tributylphosphate (TBP) and other organophosphorus compounds have been widely used as extracting agents for metal ions.[65] There have been several reports of the hydration of these compounds,[66] although in most studies the

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[62] J. Steinhardt and S. Beychok, in 'The Proteins', ed. H. Neurath, Academic Press, New York and London, 1964, vol. 2, p. 276.

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water activity was not varied systematically so that the stoicheiometry of water in the hydrate formation reactions could be inferred. A ^{1}H n.m.r. investigation[67] yielded equilibrium constants equal to 6·1 and 0·73 l mole^{-1} (converted from mole fractions to molarity units; interpolated to 25°) for the reactions TBP + W = TBP·W and TBP·W + TBP = (TBP)$_2$·W, respectively. Enthalpy values of $-4{\cdot}1$ and $-2{\cdot}0$ kcal/mole, respectively, were given for the two hydration reactions. Water solubility–partition data for TBP yield 1:1 hydrate formation constants of 34 in iso-octane,[66d] 18 in Cl$_4$,[66b] 9·1 in benzene,[66e] and 2·6 l mole^{-1} in CHCl$_3$.[66e] ^{1}H N.m.r.[68] and partition–water solubility[69] studies indicate that HNO$_3$, H$_2$O and (HNO$_3$)$_2$, H$_2$O exist in benzene and toluene. Recent reviews provide numerous references to investigations of the hydration of ions in organic media[35] and the extraction of metal salts and chelates.[65] There is ample evidence for the formation of strong complexes between water and extracting agents (including ions, ion pairs, and polar molecules) but accurate thermodynamic data for well-characterised hydrate species are generally lacking.

4 Solvent Effects

Variation in polarity and reactivity of solvents strongly affects values of the thermodynamic functions characteristic of hydrogen bond formation reactions. Even the so-called 'inert' solvents influence molecular complex formation equilibria, in general reducing the tendency for aggregates to form, as compared to the vapour phase.[6,70] More polarisable and more polar solvents ordinarily lead to a further decrease in the percentage of associated species present at fixed total concentration of donor and acceptor. Henry's Law constants,[71] X—H stretching frequencies,[70a] solubility parameters,[70c] and ^{1}H n.m.r. chemical shifts of dissolved solutes[24] have been related to the reactivity of solvents; in a general way, changes in these parameters parallel variations in thermodynamic constants for association reactions of the solutes in the various media.

Quantitative treatments of the effect of solvent on complex formation reactions have fallen into two main classes:[70c,72] (*i*) methods in which corrections are made for presumed competitive equilibria involving the formation of complexes between solvent molecules and one or more of the dissolved solute species; and (*ii*) techniques in which solvation of the donor, acceptor, and complex molecules is considered to occur in a non-specific way. Neither approach by itself has

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[70] *a* A. R. H. Cole and A. J. Michell, *Austral. J. Chem.*, 1965, **18**, 102; *b* T. Gramstad, *Spectrochim. Acta*, 1963, **19**, 1363; *c* H. Buchowski, J. Devaure, P. V. Huong, and J. Lascombe, *Bull. Soc. chim. France*, 1966, 2532.

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proven to be entirely satisfactory. Methods of type (*i*) require the assumption that equilibrium constants for specific molecular complexes involving the solvent and other species are truly constant over wide ranges of concentration and in different media. On the other hand, type (*ii*) methods tend to underestimate the effect that formation of specific molecular complexes between solute and solvent has on the energetics of the dissolved donor, acceptor, or complex.

We shall limit our discussion of solvent effects to some recent applications of non-specific solvation theories. In treating data for reactions such as A(solvated acceptor) + D(solvated donor) = AD(solvated complex) it is convenient to define the dimensionless parameter $\alpha = \underset{\mathrm{V\to S}}{\Delta E_{\mathrm{AD}}{}^{0}} / (\underset{\mathrm{V\to S}}{\Delta E_{\mathrm{A}}{}^{0}} + \underset{\mathrm{V\to S}}{\Delta E_{\mathrm{D}}{}^{0}})$ in which $\underset{\mathrm{V\to S}}{\Delta E_{\mathrm{AD}}{}^{0}}$, $\underset{\mathrm{V\to S}}{\Delta E_{\mathrm{A}}{}^{0}}$, and $\underset{\mathrm{V\to S}}{\Delta E_{\mathrm{D}}{}^{0}}$ represent internal energies of transfer of the individual species from vapour to the solvent, s, at infinite dilution.[72e,73] The parameter α may be interpreted as the fraction of the energy of solvation of the free donor plus acceptor that is retained by the complex AD; in the absence of strong dipole enhancement in the complex, it is expected that $\alpha < 1$, since the process of bringing the reactive groups of donor and acceptor together to form the complex should 'squeeze out' solvent molecules. Apparently α is not strongly dependent upon choice of medium; a similar parameter, $\alpha' = \underset{\mathrm{V\to S}}{\Delta G_{\mathrm{AD}}{}^{0}} / (\underset{\mathrm{V\to S}}{\Delta G_{\mathrm{A}}{}^{0}} + \underset{\mathrm{V\to S}}{\Delta G_{\mathrm{D}}{}^{0}})$ (where the ΔG^0 terms represent solvation Gibbs free energies for 1M ideal dilute solution standard states) is also nearly constant from solvent to solvent. Procedures have been developed for calculating α from a lattice theory, utilising group interaction energies,[74] and for predicting α' from α.[73a] Several useful relations involving α and α' are:

$$\Delta E_{\mathrm{II}}{}^{0} - \Delta E_{\mathrm{I}}{}^{0} = (\alpha - 1)\,(\underset{\mathrm{I\to II}}{\Delta E_{\mathrm{A}}{}^{0}} + \underset{\mathrm{I\to II}}{\Delta E_{\mathrm{D}}{}^{0}}) \tag{5}$$

$$\Delta G_{\mathrm{II}}{}^{0} - \Delta G_{\mathrm{I}}{}^{0} = (\alpha' - 1)\,(\underset{\mathrm{I\to II}}{\Delta G_{\mathrm{A}}{}^{0}} + \underset{\mathrm{I\to II}}{\Delta G_{\mathrm{D}}{}^{0}}) \tag{6}$$

$$K_{\mathrm{AD}}{}^{\mathrm{II}} / K_{\mathrm{AD}}{}^{\mathrm{I}} = (K_{\mathrm{D,A}} K_{\mathrm{D,D}})^{\alpha' - 1} \tag{7}$$

in which $\Delta E_{\mathrm{II}}{}^{0}$ and $\Delta E_{\mathrm{I}}{}^{0}$ are standard internal energy changes for the reaction A + D = AD in solvent II and solvent I, respectively; $\Delta G_{\mathrm{II}}{}^{0}$ and $\Delta G_{\mathrm{I}}{}^{0}$ are the corresponding standard free energy changes; $K_{\mathrm{AD}}{}^{\mathrm{II}}$ and $K_{\mathrm{AD}}{}^{\mathrm{I}}$ are equilibrium constants for the formation reaction in the two solvents, in l mole^{-2} units; and $K_{\mathrm{D,A}}$ and $K_{\mathrm{D,D}}$ are distribution constants between solvent I and solvent II for the acceptor and donor, respectively. ($K_{\mathrm{D,A}}$, for example, may be equated to the ratio of the Henry's Law constant of the acceptor in solvent I to the Henry's Law constant of acceptor in solvent II; a large value of K_{D} for a given solute indicates that solvent II is much more effective than solvent I in solvating the species.)

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[74] *a* T. L. Stevens, Ph.D. Dissertation, University of Oklahoma, 1968; *b* J. R. Johnson, P. J. Kilpatrick, S. D. Christian, and H. E. Affsprung, *J. Phys. Chem.*, 1968, **72**, 3223.

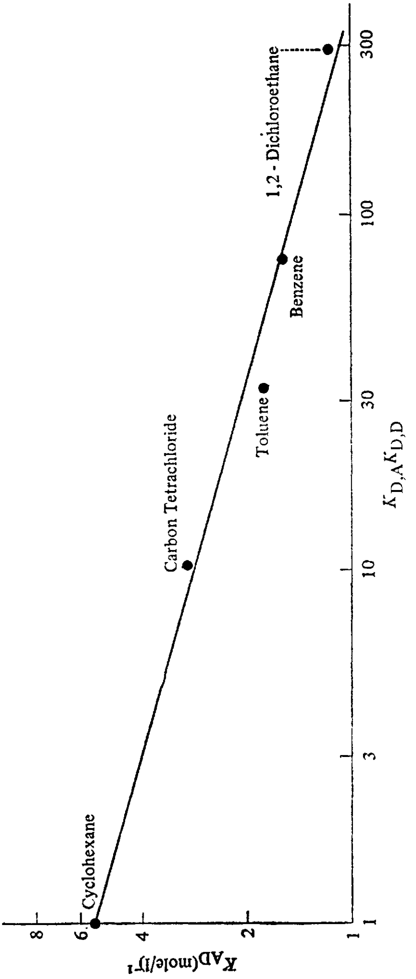

Figure 2 *Correlation of association and distribution constants for the reaction: water + pyridine = water–pyridine at 25°*

If α' is constant, equation (7) requires that for a given complex formation reaction in a series of solvents a plot of log K_{AD} *vs.* log ($K_{D,A}K_{D,D}$) will be linear with slope $\alpha' - 1$. Such a plot is shown in Figure 2 for the reaction water (A) + pyridine (D) = pyridine monohydrate (AD).[72e] The points lie approximately on a straight line having a slope of — 0·29, corresponding to $\alpha' = 0{\cdot}71$. For the solvents employed, lattice calculations lead to predicted values of α and α' of *ca.* 0·75.

Strong complexes, for which the dipole moment is considerably greater than the vector sum of the moments of the unreacted donor and acceptor, exhibit anomalous solvation effects.[75] To account for the large values of α and α' obtained for amine–water complexes, it has been proposed that the solvation energy of the complex includes an extra contribution due to excess dipole-induced dipole interactions not properly accounted for in lattice energy calculations.[57] Table 5 lists values of α and α' for the 1:1 hydrate of diethylamine (DEA) in several solvents, as well as solvation energies and free energies for the individual species involved in the reaction.[34b] Predicted values of α and α' for n-hexadecane and diphenylmethane, obtained from lattice calculations, again lie in the range 0·7—0·8. The difference between the energy of solvation of the complex in each solvent, and that predicted from the solvation energies of the unreacted components using $\alpha = 0{\cdot}75$, is *ca.* -1 to -2 kcal/mole. A major part of this extra solvation energy is probably due to the interaction of the excess dipole of the complex with the polarisable solvent. It is clear that non-polar and slightly polar solvents are less effective in preventing the formation of highly polar complexes than they are in opposing complex formation between weakly-interacting donor and acceptor molecules.

Table 5 *Effect of solvent on donor, acceptor, and complex for the reaction* DEA + W = DEA·W

	n-Hexadecane	*Diphenylmethane*	*Benzyl ether*
α(experimental)	1·01	0·85	0·80
α'(experimental)	1·05	1·01	0·90
$\underset{\mathrm{V\to S}}{\Delta E^0_{\mathrm{DEA\cdot H_2O}}}$ (kcal/mole)	−7·80	−8·92	−10·18
$\underset{\mathrm{V\to S}}{\Delta E^0_{\mathrm{DEA}}}$ (kcal/mole)	−5·60	−6·29	−6·43
$\underset{\mathrm{V\to S}}{\Delta E^0_{\mathrm{H_2O}}}$ (kcal/mole)	−2·13	−4·17	−6·33
$\underset{\mathrm{V\to S}}{\Delta G^0_{\mathrm{DEA\cdot H_2O}}}$ (kcal/mole)	−3·80	−5·20	−5·26
$\underset{\mathrm{V\to S}}{\Delta G^0_{\mathrm{DEA}}}$ (kcal/mole)	−3·16	−3·36	−3·33
$\underset{\mathrm{V\to S}}{\Delta G^0_{\mathrm{H_2O}}}$ (kcal/mole)	−0·45	−1·79	−2·54

The properties of aqueous solutions and of dilute solutions of water in non-polar media are apparently little related. However, an understanding of the progressive changes which occur in the solvation of water and other polar molecules in various media, including dipole–solvent interactions[72f] and inductive effects,[3a] should form the basis for building a bridge between these two interesting areas of chemistry.

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[75] J. Grundnes and S. D. Christian, *J. Amer. Chem. Soc.*, 1968, **90**, 2239.