The Use of Chemical Probes for the Characterization of Solvent Mixtures. Part 2.[1] Aqueous Mixtures

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The question of whether chemical properties, such as polarity or hydrogen bond donation or acceptance, can be measured in aqueous solvent mixtures by means of indicator probes, or whether their use is obviated because of preferential solvation, is examined. In some cases, such as the Kamlet–Taft π^* or β parameters, the use of several probes yielding convergent results provides acceptable values of the properties. In another case—*i.e.* the Kamlet–Taft α parameter—this question must remain open because of the relatively large spread of values obtained with different probes, which is not necessarily related to preferential solvation. A single probe, such as the betaine used for the $E_T(30)$ polarity parameter, cannot provide an answer.

In a previous paper,[1] the use of chemical probes (indicator molecules and the signals generated by them) for the determination of chemical properties of binary non-aqueous solvent mixtures was examined. Such probes, used in neat solvents, provide empirical measures of properties such as 'polarity',[2,3] hydrogen bonding abilities, *etc.*,[3] for use in linear solvation energy relationships (LSERs) involving arbitrary solutes. However, their application to mixtures, although common practice, is not really justifiable in general, in view of the preferential solvation of the probes and other solutes in mixtures. The employment of several judiciously chosen probes for a given property may, however, overcome this problem, and provide an approximate value for this property that should be valid for such LSERs.[4] This should be the case if probes of diverse structure and chemical nature yield convergent results in mixtures, which do not differ by more than the differences noted for these probes when used in neat solvents.[1]

These abstract ideas are examined in this paper for solvent mixtures containing water as one component. Aqueous mixtures differ from completely non-aqueous ones in that the constraints of miscibility limit the range of co-solvents used to polar ones, and in the highly structured and hydrogen bonding nature of water as the one component. The solvent properties examined are the Kamlet–Taft[5] dipolarity–polarizability parameter, π^*, electron pair donicity (or hydrogen bond accepting parameter), β, and hydrogen bond donation ability, α. These properties, at least for neat solvents, are orthogonal to each other, whereas other 'polarity' indices[2] are generally differently weighted composites of these properties.[3] Thus, the Dimroth–Reichardt $E_T(30)$,[6] the Kosower Z[7] and the Mayer–Gutmann AN,[8] available for many aqueous mixtures, are composites of the dipolarity–polarizability and the hydrogen bond donation ability.[3] Because of this, probes used to determine π^* are examined before those used to determine α by means of $E_T(30)$, Z or AN, for example.

In order to compare the data from various authors, they were recalculated on the common mole fraction composition scale x_w, being the mole fraction of the water in the mixtures (ignoring the probe, assumed to be present at negligible concentrations). A fourth degree power series in x_w in the form of eqn. (1) was

$$Y = x_wY_w + (1 - x_w)Y_s + \Delta Y \qquad (1)$$

$$\Delta Y = x_w(1 - x_w)[a + bx_w + cx_w^2] \qquad (2)$$

used to fit the data, where Y is the generalized symbol for the property reported and subscript 's' denotes the cosolvent. The excess (or deficiency) of Y above (or below) the linear dependence of Y on the composition, ΔY [given by eqn. (2)], is a non-unique measure of the preferential solvation.[1]

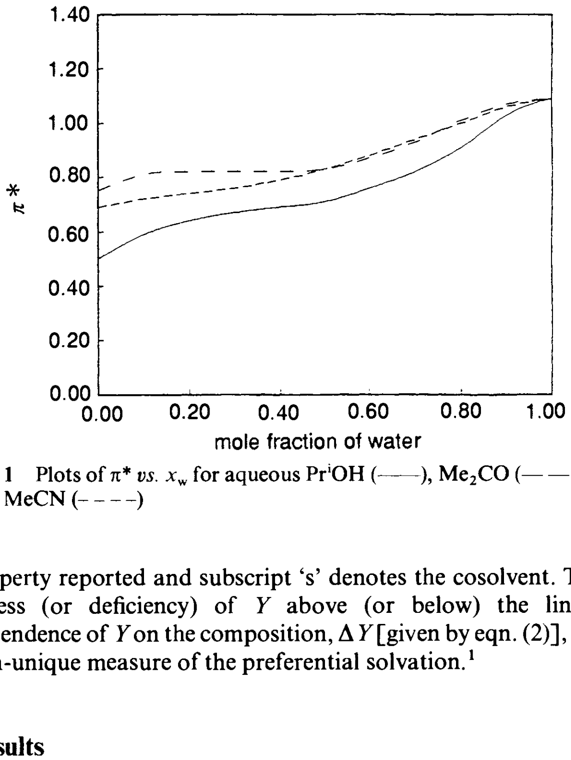

Fig. 1 Plots of π^* *vs.* x_w for aqueous Pr^{i}OH (——), Me_2CO (— — —) and MeCN (- - - -)

Results

The Dipolarity–Polarizability of Aqueous Mixtures.—The Kamlet–Taft π^* parameter has been determined over the entire composition range for aqueous methanol (MeOH),[4,9] ethanol (EtOH),[10] propan-2-ol (Pr^{i}OH),[9] tetrahydrofuran (THF),[4,9] acetone (Me_2CO),[11] formic acid (HCO_2H),[12] acetic acid ($MeCO_2H$),[4] formamide (FA),[12] *N*,*N*-dimethylformamide (DMF),[12] acetonitrile (MeCN),[9,13] pyridine (Py),[12] and dimethyl sulfoxide (Me_2SO).[12] The results are shown in Table 1 and for some representative systems in Fig. 1. The probes that have been used include *N*,*N*-diethyl-4-nitroaniline, *N*,*N*-diethyl-3-nitroaniline, 4-nitroanisole and 4-ethylnitrobenzene. These are all dipolar nitro-aromatic compounds, which vary in their other functional group, but are not very different; not surprisingly, they do give convergent results.

Several probes in addition to the above-mentioned have been used for the determination of π^* for neat solvents, including 2-nitroanisole,[9] *N*-methyl-2-nitroaniline,[9] 4-(2-nitroethenyl)-anisole[4] (these have also been used for the aqueous mixtures), *N*,*N*-diethyl-3,4-dinitroaniline, 4-(*N*,*N*-dimethylamino)benzo-

Table 1 π^* values of aqueous mixtures; numerical column headings denote x_w values; for water ($x_w = 1$), $\pi^* = 1.09$

Solvent	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9	Ref.
MeOH	0.60	0.66	0.72	0.78	0.85	0.91	0.98	1.04	1.10	1.15	9
	0.59	0.67	0.73	0.79	0.85	0.91	0.96	1.02	1.07	1.10	4
EtOH	0.54	0.62	0.65	0.68	0.73	0.79	0.88	0.99	1.10	1.17	10
Pr^iOH	0.50	0.59	0.64	0.67	0.69	0.71	0.76	0.82	0.91	1.03	9
THF	0.60	0.67	0.70	0.71	0.72	0.73	0.76	0.81	0.90	1.01	9
	0.56	0.61	0.65	0.68	0.70	0.72	0.75	0.79	0.85	0.95	12
Me_2CO	0.69	0.72	0.74	0.76	0.79	0.83	0.88	0.94	1.00	1.06	11
HCO_2H	0.99	1.02	1.04	1.06	1.08	1.10	1.11	1.12	1.13	1.12	12
$MeCO_2H$	0.58	0.69	0.76	0.81	0.85	0.91	0.97	1.04	1.10	1.13	4
FA	1.15	1.16	1.16	1.16	1.16	1.16	1.16	1.16	1.15	1.13	12
DMF	0.89	0.91	0.92	0.94	0.96	1.00	1.04	1.08	1.12	1.12	12
MeCN	0.75	0.81	0.82	0.82	0.82	0.83	0.87	0.93	1.01	1.10	9
	0.73	0.75	0.77	0.80	0.82	0.85	0.87	0.92	0.97	1.08	13
Py	0.89	0.93	0.96	0.99	1.02	1.06	1.11	1.15	1.15	1.13	12
Me_2SO	0.98	0.99	1.00	1.01	1.03	1.05	1.07	1.10	1.12	1.12	12

Table 2 β values of aqueous mixtures; numerical column headings denote x_w values; for water ($x_w = 1$), $\beta = 0.47$

Solvent	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9	Ref.
MeOH	0.70	0.70	0.68	0.67	0.67	0.68	0.69	0.68	0.65	0.59	4
	0.68	0.70	0.65	0.68	0.70	0.73	0.70	0.64	0.56	0.46	15
EtOH	0.71	0.72	0.68	0.66	0.66	0.65	0.65	0.66	0.63	0.46	15
Pr^iOH	0.81	0.84	0.80	0.77	0.74	0.71	0.70	0.68	0.66	0.59	15
THF	0.64	0.69	0.69	0.69	0.69	0.69	0.69	0.67	0.69	0.67	13
	0.50	0.47	0.44	0.42	0.43	0.41	0.44	0.50	0.58	0.64	15
Diox	0.42	0.54	0.61	0.63	0.63	0.61	0.59	0.56	0.54	0.51	16
	0.39	0.45	0.50	0.55	0.55	0.52	0.51	0.50	0.47	0.39	15
Me_2CO	0.56	0.61	0.62	0.63	0.64	0.65	0.67	0.68	0.67	0.61	11
	0.53	0.54	0.54	0.56	0.58	0.60	0.59	0.58	0.56	0.53	15
$MeCO_2H$	0.48	0.53	0.54	0.54	0.55	0.56	0.58	0.60	0.60	0.57	4
FA	0.55	0.55	0.54	0.54	0.54	0.54	0.53	0.53	0.52	0.50	12
DMF	0.77	0.78	0.76	0.76	0.73	0.71	0.68	0.66	0.62	0.57	12
	0.73	0.73	0.72	0.74	0.72	0.69	0.68	0.65	0.63	0.59	15
MeCN	0.47	0.53	0.57	0.59	0.59	0.59	0.60	0.61	0.61	0.61	13
	0.44	0.49	0.53	0.55	0.54	0.53	0.54	0.56	0.56	0.55	15
Py	0.66	0.69	0.70	0.69	0.68	0.66	0.63	0.60	0.57	0.53	12
Me_2SO	0.80	0.75	0.78	0.71	0.73	0.69	0.68	0.63	0.62	0.56	12
	0.76	0.79	0.77	0.69	0.66	0.66	0.61	0.59	0.59	0.57	15

nitrile, 4-(*N*,*N*-dimethylamino)benzophenone and methyl 4-(*N*,*N*-dimethylamino)benzoate, among a few others.[14] Some of the latter have different functional groups (nitrile, ketone, ester) to the nitro groups of the former, but all are aromatic molecules. Still, it should be instructive to see whether they give similar values of π^* in aqueous mixtures as they do in neat solvents. This would strengthen the argument that convergent results in mixtures lead to properties (in this case, combined dipolarity–polarizability) that can characterize solvent mixtures as well as they do neat solvents.

The Electron Pair Donicity of Aqueous Mixtures.—The Kamlet–Taft β parameter has been determined over the entire composition range in aqueous MeOH,[4,15] EtOH,[15] propan-1-ol (PrOH),[15] Pr^iOH,[15] THF,[4,15] 1,4-dioxane (Diox),[15,16] Me_2CO,[11,15] HCO_2H,[4] $MeCO_2H$,[12] FA,[4] DMF,[4,15], MeCN,[13,15] Py[4] and Me_2SO.[4,15] The probes used are 4-nitroaniline, 4-nitrophenol, and acetylacetonato (*N*,*N*,*N*,*N*-tetramethylethylenediamine)copper(II) perchlorate. For the former two, the dependence of the wavenumber of light absorption on π^* (as well as on β) has to be taken into account, and this was done in two ways. In ref. 12, the mean π^* (from all probes measured) was used in eqn. (3), where a, $\bar{\nu}_o$ and b have

$$\beta = a(\bar{\nu}_o - \bar{\nu}) - b\pi^* \qquad (3)$$

the values appropriate for each probe in neat solvents[3] ($\bar{\nu}_o$ is the wavenumber for a solvent with $\beta = \pi^* = 0$; for the copper complex, $b = 0$). In ref. 15, the only probe pair employed was 4-nitroaniline and *N*,*N*-diethyl-4-nitroaniline (*i.e.* the terms on the r.h.s. of eqn. (3) with a, $\bar{\nu}_o$, b and π^* for this particular pair), yielding B_{KT} values instead of β. Where data from the two groups of authors are available, the agreement is good, provided B_{KT} is first transformed into a β value by means of the expression $\beta = 0.95B_{KT} + 0.055\,\pi^* - 0.007$.[12] Important exceptions are water-rich mixtures ($x_w \geq 0.8$) and the THF mixtures at all compositions, where the data from ref. 15 are conspicuously low and possibly incorrect.[12] The results are shown in Table 2 and for some representative mixtures in Fig. 2.

The 'Polarity' of Aqueous Mixtures.—Of the many 'polarity indices' that have been proposed over the years, only two have shown sufficient viability to be of widespread use nowadays. These are the Dimroth–Reichardt $E_T(30)$ index, using 2,6-diphenyl-4-(2,4,6-triphenylpyridin-1-iumyl)phenoxide, and the Kosower Z index, using 1-ethyl-4-methoxycarbonylpyridinium iodide. Both use solvatochromatic indicator probes, and hence are readily measurable, their values being expressed as the transition energies of the longest wavelength absorption peak in kcal mol^{-1} (1 cal = 4.184 J). In this context, 'polarity' means the general solvation ability, including both specific and non-specific interactions;[2] in fact, the $E_T(30)$ and Z indices measure differently weighted combinations of dipolarity–polarizability (π^*) and hydrogen bond donation ability (α) of solvents.[3] Since they have been used for the determination of the latter, it is expedient to discuss them first.

Table 3 $E_T(30)$ values of aqueous mixtures; numerical column headings denote x_w values; for water (x_w = 1), $E_T(30)$ = 63.1

Solvent[a]	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9	Ref.
MeOH	55.7	55.8	56.1	56.5	56.8	57.3	57.8	58.5	59.5	61.0	4, 10, 15, 17–19
EtOH	51.7	52.3	52.8	53.2	53.6	54.0	54.5	55.6	57.2	59.6	10, 15, 17–20
PrOH	50.6	50.8	51.7	52.6	53.3	53.5	53.4	53.6	54.6	57.6	15, 17, 20
Pr^{i}OH	48.7	48.3	49.2	50.3	51.2	51.6	51.8	52.1	53.4	56.6	10, 15, 18
Bu^{t}OH	43.9	45.6	47.5	49.0	50.0	50.5	50.9	51.6	53.2		20
Et12diol	56.3	56.5	56.8	57.1	57.5	58.0	58.6	59.2	60.1	61.4	21
Pr12diol	54.0	54.1	54.2	54.3	55.4	56.2	56.5	56.6	57.1	58.9	21
Pr13diol	54.9	55.0	55.1	55.2	55.4	56.3	56.8	57.1	57.6	59.3	21
THF	37.5	40.9	44.3	47.1	48.9	49.7	49.7	50.0	51.1	54.8	10, 15, 19
Diox	36.0	40.3	43.7	46.3	47.1	48.9	49.5	50.5	52.4	56.3	15, 16, 18, 20, 21
Me_2CO	42.2	45.6	48.3	50.3	51.6	52.4	52.7	53.4	54.9	57.8	11, 15, 17–20
DMF	43.8	44.9	46.2	47.5	48.7	49.9	51.1	52.7	54.9	58.2	15, 20
MeCN	46.0	50.8	53.6	54.9	55.4	55.6	55.8	56.5	57.6	59.8	10, 13, 15, 20
Py	40.3	42.0	44.1	46.1	47.0	48.1	49.2	51.5	53.5	57.3	15, 20
2-MePy	38.3	40.5	42.6	44.4	45.9	47.2	48.5	50.2	52.8	56.9	18
2,6-Me_2Py	36.7	39.5	41.8	43.6	45.2	46.6	48.1	50.1	53.0	57.1	18
Pip	35.5	38.0	40.4	42.5	44.2	45.6	47.0	48.8	51.6	56.0	18
Me_2SO	45.0	45.6	46.5	47.6	48.7	50.0	51.3	53.1	55.4	58.5	15,22

[a] Et12diol = ethane-1,2-diol, Pr12diol = propane-1,2-diol, Pr13diol = propane-1,3-diol, 2-MePy = 2-methylpyridine (2-picoline), 2,6-Me_2Py = 2,6-dimethylpyridine (2,6-lutidine), Pip = piperidine.

Table 4 Z values of aqueous mixtures; numerical column headings denote x_w values; for water (x_w = 1), Z = 92.5

Solvent	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9	Ref.
MeOH	83.6	84.5	85.3	86.0	86.7	87.5	88.5	89.6	91.0	92.6	23, 24
EtOH	79.5	80.5	81.7	82.8	84.0	85.2	86.3	87.6	89.3	91.4	23, 24, 25
Pr^{i}OH	76.3	78.0	79.5	80.8	82.1	83.5	85.1				25
Bu^{t}OH	71.3	73.5	75.6	77.7	79.7	81.7	83.9				25
THF[a]	58.3	67.7	74.2	78.1	80.4	81.6	82.3	83.4	85.3	88.9	26
Diox	65.4	68.0	71.3	74.6	77.5	79.9	81.8	83.6	85.7	89.0	23
Me_2CO	65.7	70.5	74.1	76.8	79.0	81.0	82.9	85.1	87.6	90.5	23, 24
HCO_2H	88.7	88.5	88.8	89.2	89.5	89.6	89.6	89.7	90.1	91.5	23
$MeCO_2H$	80.0	83.8	86.0	87.0	87.3	87.2	87.0	87.2	88.3	90.5	23
MeCN	71.7	76.3	79.7	82.2	84.1	85.5	86.7	87.8	89.1	90.6	27
Me_2SO	71.2	72.5	73.8	75.3	76.9	78.8	81.0	83.6	86.7	90.3	24

[a] Calculated as $Z = Z'/0.944$.

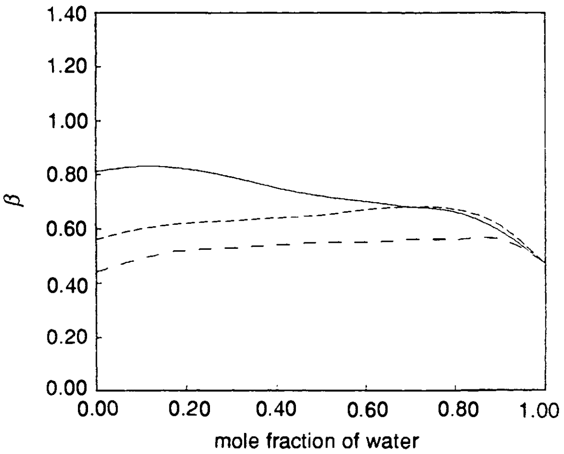

Fig. 2 Plots of β vs. x_w for aqueous Pr^{i}OH (——), Me_2CO (— — —) and MeCN (- - - -)

$E_T(30)$ has been measured for a fairly large number of aqueous mixtures (see Table 3). The precision of the measurement of the wavelength of the lowest energy absorption peak of the betaine dye is ± 1 nm, leading to ± 0.1 kcal mol^{-1} in $E_T(30)$. The unweighted average of the reported values for each solvent composition should represent the actual $E_T(30)$ within ± 0.3 kcal mol^{-1}, as reported in Table 3. These values are the basis for the calculation of α values (see below).

Z values have been measured for fewer aqueous mixtures than $E_T(30)$, but recent measurements of their variant, called here Z' (using 4-cyano-1-ethylpyridinium iodide instead of the original compound) augment their number to a considerable extent. For neat solvents, the latter is practically proportional to the former: $Z' = (-0.03 \pm 0.08) + (0.944) \pm (0.002)Z$ (ref. 3). If this is taken to hold also for the aqueous mixtures, then a unified scale can be set up. Table 4 lists the unweighted average values of Z (eventually calculated from Z') available for the aqueous mixtures. These values, again, are the basis for the calculation of α values (see below).

The Hydrogen Bond Donation Ability of Aqueous Mixtures.— The Kamlet–Taft α parameter can be determined directly, i.e. without the need for auxiliary data such as π*, from the ^{13}C NMR chemical shifts of 2- and 3-C relative to 4-C in pyridine *N*-oxide.[11] This probe was applied only to aqueous MeOH, Me_2CO and MeCN, however. Other probes that are sensitive to α are sensitive to π* too, to a greater or lesser extent. The ^{13}C

Table 5 α Values of aqueous mixtures; numerical column headings denote x_w values; for water (x_w = 1), α = 1.17

Solvent	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9	Ref.
MeOH	1.00	1.01	1.01	1.00	0.99	0.98	0.98	1.01	1.04	1.11	4, 8, 10, 11, 15, 23
EtOH	0.90	0.91	0.93	0.94	0.96	0.96	0.96	0.96	1.00	1.09	8, 10, 23
Pr^{i}OH	0.77	0.76	0.79	0.83	0.88	0.91	0.91	0.91	0.93	1.01	8, 10, 15, 18, 25
THF	0.00	0.15	0.30	0.42	0.48	0.50	0.50	0.50	0.54	0.74	12
Diox	0.02	0.20	0.38	0.52	0.60	0.73	0.77	0.82	0.92	1.06	8, 15, 16, 18, 20, 23
Me_2CO	0.08	0.33	0.48	0.60	0.68	0.73	0.76	0.83	0.90	1.03	8, 11, 15, 17, 18, 23
HCO_2H	1.27	1.24	1.23	1.23	1.22	1.21	1.18	1.17	1.16	1.21	23
$MeCO_2H$	1.12	1.08	1.07	1.09	1.08	1.05	1.01	0.98	0.99	1.05	4, 23
FA	0.63	0.65	0.68	0.71	0.74	0.77	0.80	0.83	0.89	0.99	12
DMF	0.00	0.10	0.19	0.27	0.32	0.36	0.40	0.46	0.57	0.79	12
MeCN	0.25	0.55	0.70	0.78	0.83	0.86	0.89	0.92	0.96	1.05	4, 8, 10, 11, 13, 15
Py	0.00	0.00	0.06	0.12	0.19	0.24	0.29	0.36	0.49	0.73	12
Me_2SO	0.00	0.06	0.13	0.20	0.26	0.32	0.38	0.46	0.59	0.81	12
HMPT	0.00	0.00	0.04	0.12	0.21	0.31	0.43	0.58	0.75	0.97	8

NMR chemical shifts of the ring carbons relative to the carbonyl carbon in *N*,*N*-dimethyl- (or -diethyl-) benzamide[4] are 1.9 to 4.7 times more sensitive to α than to π^* in neat solvents, and presumably also in mixtures. This probe was applied to aqueous MeOH, THF, $MeCO_2H$, FA, DMF, MeCN, Py and Me_2SO. The ^{31}P NMR chemical shifts of triethylphosphine oxide (extrapolated to zero probe concentration, corrected for bulk susceptibility and normalized to a scale from 0 for hexane to 100 for antimony pentachloride) give the *AN* scale,[8] which is 2.1 times more sensitive to α than to π^* in neat solvents. Assuming this also to hold in aqueous mixtures, the *AN* data[8] for MeOH, EtOH, Pr^{i}OH, Diox, Me_2CO, DMF, MeCN, Py, Me_2SO, and hexamethylphosphoric triamide (HMPT) are converted into α by means of eqn. (4).[3] A slightly lower relative sensitivity to α

$$\alpha = [AN - 2.9 - 14.0\pi^*]/29.7 \qquad (4)$$

and π^* in neat solvents, a factor of 2.0, is also observed for *Z* (and *Z′*). The conversion expression of the values in Table 4 to α is given by eqn. (5). For these probes (except pyridine *N*-oxide),

$$\alpha = [Z - 55.9 - 10.2\pi^*]/20.6 \qquad (5)$$

a knowledge of π^* for the aqueous mixtures is required, although approximate values obtained by estimation may suffice, because of the relatively low sensitivity of the probe data to π^*.

The relative sensitivity of $E_T(30)$ to α compared with π^* in neat solvents is smaller than that of the probes discussed above, being only a factor of 1.3, so that reliable values of π^* become more important. The expression used to convert the $E_T(30)$ values in Table 3 into α values is given by eqn. (6).[3]

$$\alpha = [E_T(30) - 31.2 - 11.5\pi^*]/15.2 \qquad (6)$$

Another group of probes that have been employed to give α values for aqueous mixtures are based on dicyanoiron(II) complexes further substituted as bis(1,10-phenanthroline)[12] and bis[α-(2-pyridylbenzylidene)-3,4-dimethylaniline] derivatives.[10] Their relative sensitivities to α and π^* are 2.2 and 1.6, respectively. However, these probes are also sensitive to β, although 3.7 times less so than towards α. The dicyanobisphenanthroline complex was applied to aqueous MeOH, THF, Me_2CO, $MeCO_2H$, FA, DMF, MeCN, Py and Me_2SO systems, whereas the other iron complex was applied to aqueous MeOH, EtOH, Pr^{i}OH, THF and MeCN.

The averaged α values of aqueous mixtures obtained from all these probes, pyridine *N*-oxide, dialkylbenzamides, triethylphosphine oxide (*AN*), 4-methoxycarbonyl- and 4-cyano-1-ethylpyridinium iodide (*Z* and *Z′*), the betaine used for $E_T(30)$, and the iron complexes, are shown in Table 5.

Other Probes and Properties of Aqueous Solvents.—4-(1-Methyl-2-ethenylpyridin-1-ium-4-yl)phenoxide, a different betaine to that used for the $E_T(30)$ measurements, was used by Dawber *et al.*[17] in aqueous MeOH, EtOH, PrOH and Me_2CO. For the former two aqueous mixtures, the results do not differ much from the $E_T(30)$ data, but for the latter two, they do differ. There are insufficient data for neat solvents for this probe to be able to relate the results to α values.

Some other properties have been measured with chemical probes in aqueous solvents, such as Winstein's *Y* and Brooker's χ_B and χ_R. The former[28] described the (log of the) rate constant at 25 °C of the unimolecular solvolysis of *tert*-butyl chloride relative to the rate in ethanol–water (80% v/v) as reference. The numerical *Y* scale runs from −3.26 for *tert*-butyl alcohol (Bu^{t}OH) to 0.00 for the reference solvent to 3.49 for water, and values are known for (only) nine neat solvents. Aqueous mixtures have been studied over the entire composition range for MeOH, EtOH, HCO_2H and $MeCO_2H$ and over a part of it for Diox and Me_2CO. Eqn. (7) holds over much of the

$$Y = a + 0.9E_T(30) \qquad (7)$$

composition range (up to $x_w \sim 0.8$) for those solvents for which $E_T(30)$ values are known (Table 3) with $a = -50.5$ for MeOH, −48.2 for EtOH, −47.4 for Me_2CO and −44.4 for Diox. In water-rich mixtures, the lines $Y = f[E_T(30)]$ bend over to meet at the common point for water. These scant data, therefore, do not provide more information than $E_T(30)$ does, and the labour involved in the kinetic measurements does not encourage their further use.

Two of Brooker's merocyanine dyes show large red- and blue-shifts, respectively.[29] The former dye yields χ_R for nearly 60 neat solvents on a scale from 50.9 for hexane to 33.6 for 3-methylphenol; the latter dye yields χ_B for only 12 neat solvents, on a scale from 41.7 for toluene to 68.9 for water (numerical values in kcal mol^{-1}). Eqn. (8) holds with a correlation

$$\chi_B = 8.6 + 0.97E_T(30) \qquad (8)$$

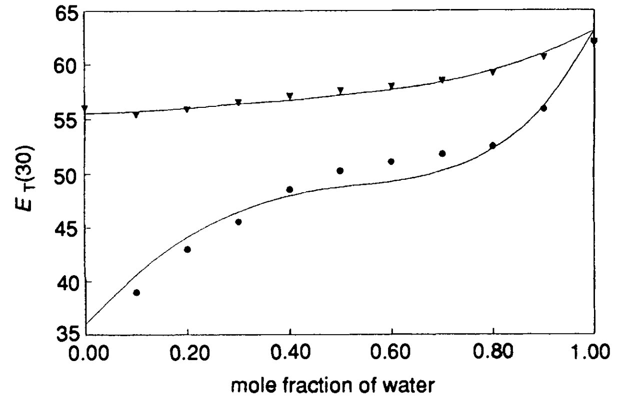

Fig. 3 Plots of $E_T(30)$ calculated *via* eqn. (8) from χ_B of Brooker's blue-shifting merocyanine in aqueous MeOH (▼) and Diox (●) and the corresponding values with the Dimroth–Reichardt betaine (——) against x_w

coefficient of > 0.99. Brooker *et al.* provided χ_B for four aqueous solvents [MeOH, Diox, Py and 2,6-dimethylpyridine (2,6Lu)] over the entire composition range,[29] and when eqn. (8) is applied to these data in the inverse form, $E_T(30)$ values are obtained for the mixtures in agreement with those obtained directly. That is, the polarity of the mixtures is sensitive in a similar manner to the merocyanine and betaine dyes (Fig. 3).

A different property of solvents in general, and aqueous mixture in particular, is their softness.[30] The probe used for measuring it consists of three ions: silver, sodium and potassium, and the quantity measured is the standard molar Gibbs free energy of their transfer from water into the solvent (mixture) in question, expressed in kJ mol^{-1}. The softness parameter, μ, is given by eqn. (9). Since μ involves the difference

$$\mu = \{\Delta_{tr}G^\circ(\mathrm{Ag}^+) - [\Delta_{tr}G^\circ(\mathrm{Na}^+) + \Delta_{tr}G^\circ(\mathrm{K}^+)]/2\}/100 \quad (9)$$

in $\Delta_{tr}G^\circ$ of univalent cations, the extrathermodynamic assumption used in its specification is immaterial. The parameter μ is known for many neat solvents (to the previous listing for 32 solvents,[30] the following should be added, from recent $\Delta_{tr}G^\circ$ references:[31] γ-Butyrolactone −0.02, ethylene sulfite −0.01, tetrahydrofuran 0.00, trimethyl phosphate 0.02, hexan-1-ol ~0.12, propanenitrile 0.36, butanenitrile 0.37, 2-phenylethanenitrile 0.38, 2-methylpropanenitrile 0.41, aniline ~0.75 and pyrrole 0.81). However, μ values are known for only a few aqueous mixtures, and these involve mainly hard solvents, for which $|\mu| < 0.1$, *i.e.*, near the defined quantity for water, 0.00, and little can be learned from these. In the case of aqueous MeCN, however, more interesting results are obtained from the available $\Delta_{tr}G^\circ$ data (see Fig. 4).[32] The deviation, $\Delta\mu = \mu - (1 - x_w)\mu(\mathrm{MeCN})$ peaks near $x_w = 0.8$, where the microheterogeneous mixture of the two components becomes microscopically homogeneous again, the MeCN molecules taking up voids in the structured water.[17]

Partly Miscible Water–Solvent Systems.—Some information is available concerning $E_T(30)$ of aqueous solvent systems which are only partly miscible.[33] One case is butan-1-ol–water: $E_T(30)$ varies from 49.8 to 52.1 as x_w increases towards the miscibility gap at $x_w = 0.512$, and this butanol-rich mixture is at equilibrium (at 25 °C) with the water-rich phase at $x_w = 0.981$ with $E_T(30) = 59.8$, whence it increases to 63.1 in pure water.

The values of π^*, β and α in water-saturated alkyl alcohols,[12] including octan-1-ol,[34] as well as tributyl phosphate,[12] are shown in Table 6.

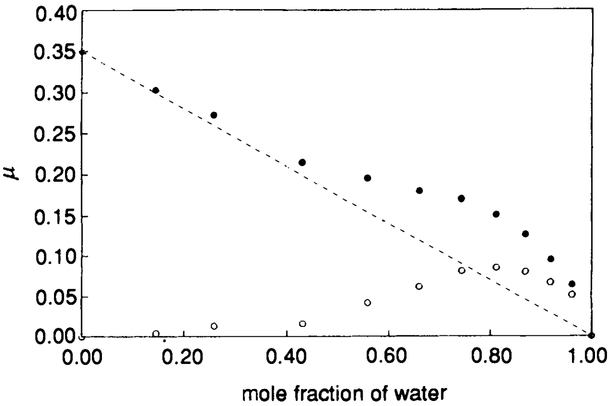

Fig. 4 Plots of the softness parameter μ of aqueous MeCN (●) and its deviation from linearity with the composition (○) against x_w

Discussion

The cosolvents present in aqueous mixtures are necessarily polar, since they would not be miscible otherwise. They include both protic (*e.g.*, alcohols, carboxylic acids) and aprotic (*e.g.*, $\mathrm{Me_2CO}$, DMF, $\mathrm{Me_2SO}$) solvents, and both very hydrophilic (ethane-1,2-diol, $\mathrm{HCO_2H}$, FA) and partially hydrophobic ($\mathrm{Pr^iOH}$, Diox, Py) ones. In spite of this variety of functional groups amongst the cosolvents, their high polarity makes some of their properties (*e.g.* π^*, β, *etc.*) not so different from those of water, as are those of apolar water-immiscible solvents. The values in aqueous mixtures, then, do not cover a large range and preferential solvation, where it exists, may not manifest itself appreciably. Those properties that strongly depend on α [Z, AN, $E_T(30)$], however, do vary considerably, from low values for aprotic cosolvents to the high value for water, and the effects of preferential solvation would be more readily discernible.

Preferential solvation has often been defined by means of eqn. (2), employing a chemical probe for the measurement of the quantity Y; this phenomenon takes place if $\Delta Y \neq 0$. Local mole fractions of the solvents around the probe are then defined[26,35] by eqn. (10), where x_w^L is the local mole fraction of component w

$$Y = x_w^L Y_w + (1 - x_w^L) Y_s \quad (10)$$

(water in our case). It is obvious, however, that such local mole fractions are specific for the particular probe employed and even for the directly measured quantities (*i.e.*, Y, Y_w and Y_s). They do not describe the local compositions in the binary solvent mixture in the absence of the probe, which may induce a certain sorting due to its particular preferences, nor those of the 'general solute', required for the specification of a 'chemical property of the mixture' to be used in LSERs.

Such problems are not encountered in the statistical thermodynamic definition of preferential solvation, given by Ben-Naim,[36] among others. He defined it quantitatively for a 'correlation region' around a particle in the mixture (a solute or one of the solvents) in a manner that is independent of measurements with chemical probes. It makes use of the Kirkwood–Buff integrals $G_{ij} = 4\pi_0\int^\infty(g_{ij} - 1)r^2\,dr$, where g_{ij} is the pair correlation function of particles i with a particle j at the origin. The integrals G_{ij} are obtainable from thermodynamic data on the binary solvent mixtures. The correlation volume V_{cor} is the volume of the region around particle j beyond which $g_{ij} \approx 1$, so that at larger distances, no increments to the integral are obtained. Then eqns. (11) hold. Eqn. (11a) gives the fraction

$$x_{ii}^L = x_i[1 + x_j(G_{ii} - G_{ji})/(x_iG_{ii} + G_{ji} + V_{cor})] \quad (11a)$$

of i molecules around a given i molecule and eqn. (11b) that

Table 6 Solvatochromic parameters π^*, β and α measured in neat ('dry') and water-saturated (at 25 °C, 'wet') water-immiscible solvents; the water content of the wet solvents is x_w[41]

		π^*		β		α		
Solvent	x_w	Dry	Wet	Dry	Wet	Dry	Wet	Ref.
Butan-1-ol	0.515	0.49	0.68	0.84	0.81	0.78	0.85	12
Isobutanol	0.456	0.51	0.69	0.83	0.82	0.73	0.80	12
Pentan-1-ol	0.340	0.50	0.63	0.81	0.78	0.73	0.80	12
Hexan-1-ol	0.313	0.47	0.59	0.81	0.78	0.74	0.80	12
Octan-1-ol	0.275	0.47	0.52	0.85	0.77	0.70	0.81	12
	0.289	0.50	0.53	0.91	0.79	0.76	0.78	34
Decan-1-ol	0.259	0.45	0.50	0.82	0.75	0.70	0.77	12
Tributyl phosphate	0.497	0.63	0.69	0.80	0.82	0.00	0.42	12

$$x^{L}_{ij} = x_i[1 + x_j(G_{ij} - G_{jj})/(x_iG_{ij} + G_{jj} + V_{cor})] \quad (11b)$$

fraction around a given j molecule.

Local mole fractions for many aqueous solvent mixtures (most of those in Tables 1–5) have been reported on the basis of the Kirkwood–Buff integrals by Matteoli and Lepori[37] and by Marcus.[38] It has been noted that self-clustering of the water ($x^L_{ww} > x_w$) takes place when the cosolvent has an appreciably hydrophobic character, *i.e.*, several methyl or methylene groups or an aromatic ring. This leads to microheterogeneity of the mixtures, and in more extreme cases (such as acetonitrile at low temperatures or butan-1-ol) to phase separation. In such cases, it has been found[41] that α and β values measured by the usual probes in the water-saturated organic solvents (Table 6) do *not* represent an average environment that a solute 'sees', but rather the results of water self-clustering. On the contrary, highly polar miscible cosolvents interact so strongly with water that $x^L_{ws} > x_w > x^L_{ww}$, this being the case for DMF and Me_2SO. These observations help to explain the results obtained for chemical probes.

As mentioned above, the dipolarity–polarizability of the aqueous mixtures, measured by π^*, is not very sensitive to preferential solvation. In the cases of MeOH, THF and MeCN, where results are available from two groups of authors, agreement between the average values of π^* (for five probes) reported by each group is within 0.03 units (on a scale that goes from -0.4 for fluorocarbons to ≈ 1.1 for water). This is better than the internal agreement among values for the different probes reported, where the spread can be up to 0.2 units in the worst cases. However, this holds for the neat solvents as it does for the mixtures, and the (unweighted) averages for the former are the accepted values.[4] This spread cannot, therefore, be ascribed to preferential solvation, but may be due to different blends of dipolarity and polarizability measured by each probe. The standard deviation of π^* for neat solvents is taken to be ± 0.06, and this appears also to hold for aqueous mixtures.

The probes used for the determination of β have different electron pair accepting sites: hydrogen atoms connected to nitrogen and to oxygen and copper ions. In this sense, they provide a more stringent test for the applicability of the probes to mixtures than do the π^* probes discussed above. The spread of values noted for given mixtures for the different probes is ≤ 0.1 units (on a scale ranging from 0 for hydrocarbons to ≈ 1 for HMPT), but the total variation of β from the values in the cosolvent, which, for the water-miscible solvents studied so far range from 0.4 to 0.8, to the β of water (0.47), is rather small. Relatively flat $\beta = f(x_w)$ curves are obtained, with a more or less abrupt decrease of β when pure water is reached. Altogether, little can be said at this stage regarding whether or not preferential solvation is of importance with respect to the donicity of solvent mixtures.

The agreement between the $E_T(30)$ values reported by various authors for a given aqueous solvent mixture (after recalculation by means of eqn. (1) to the nearest x_w in steps of 0.1) is generally within ± 0.5 kcal mol^{-1}, the same as for the neat solvents. According to eqn. (6), this can account for a ± 0.03 variation in the derived α values. The situation differs for Z data. In the cases of MeOH, EtOH and Me_2CO, there are data from two or three sources, but the agreement between the reported Z values is not so good (*e.g.*, values of 83.7[23] and 82.1[24] kcal mol^{-1} were reported for ethanol with $x_w = 0.3$). Such large discrepancies, however, are not common, and most values agree better, *i.e.* within 0.5 kcal mol^{-1}, as do most of the values for the neat solvents, when several have been reported. A notable exception is water itself, the Z value of which cannot be measured directly but is attainable through extrapolation of values for aqueous mixtures to $x_w = 1$. The reported Z values for water are 91.4[25] and 94.6,[7,25] those of Z' are 87.0[39] and 89.6,[40] translatable *via* the expression given above to $Z = 92.2$ and 94.9, respectively (all in kcal mol^{-1}). Values of α, calculated *via* eqn. (5), have an uncertainty of ± 0.03 due to the uncertainty in Z.

In many cases, α data have been obtained with several probes, and it is instructive to see what degree of agreement between the resulting α values for given aqueous solvent compositions is achieved. Figs. 5–7 show the data for aqueous MeOH, Me_2CO and MeCN. The α values for neat solvents obtained with different probes may generally differ by up to ± 0.08 from the mean,[3,5] *i.e.* have a spread of 0.16 units. There are also discrepancies among the data furnished with a given probe by different authors. This spread relates to a scale that goes from 0 for aprotic solvents to ≈ 1.2 for water and to ≈ 2 for hexafluoropropan-2-ol. A considerable part of the noted spread is due to the uncertainties inherent in the conversion expressions, eqns. (4)–(6), obtained from data for the neat solvents (typically ± 0.08 from the regression, ± 0.03 from uncertainties in π^* and ± 0.03 from uncertainties in $E_T(30)$ or Z). The spread in the aqueous MeOH mixtures is not larger than 0.16 (Fig. 5), but the spreads in aqueous Me_2CO and MeCN are seen in Figs. 6. and 7 to be larger. Still, the characteristic S-shape of the curves (due to microheterogeneity over a part of the composition range[13,38]) is retained for the latter two solvents with all the probes. Similar results are obtained for other aqueous solvents, not shown here.

The question now arises as to whether there are systematic deviations between α values obtained with different probes, ascribable to preferential solvation, or whether all the discrepancies may be ascribed to the uncertainties noted above. Park *et al.*[10] have taken the former view, at least for aqueous MeOH and EtOH. They argued that since plots of their transition energies for the betaine used in $E_T(30)$ measurements and their Fe complex are not linear, the cybotactic regions around the two probes must be different in these two aqueous mixtures. This argument is weakened by the fact that both transition energies are linear combinations of the independent

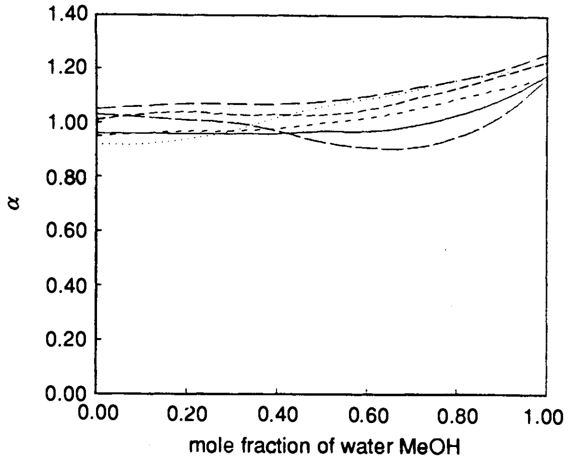

Fig. 5 Plots of α for aqueous MeOH *vs.* x_w obtained with different probes: the Fe complex (——), the $E_T(30)$ betaine (— — —), triethylphosphine oxide (*AN*,— — — —), 1-ethyl-4-methoxycarbonylpyridinium iodide (*Z*, - - - - - -), dialkylbenzamide (- - - - -) and pyridine *N*-oxide (· · · · · ·)

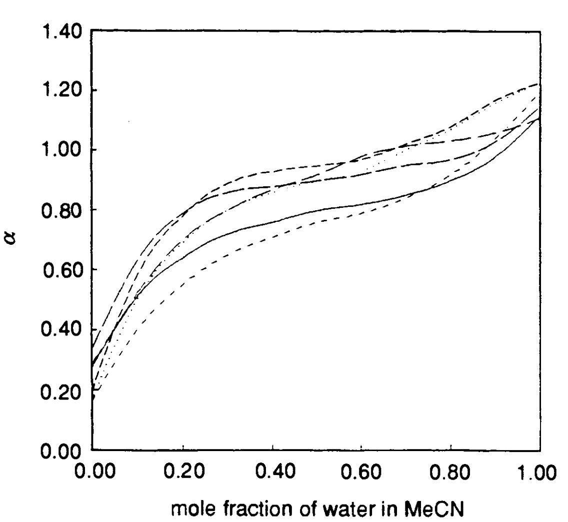

Fig. 6 Plots of α for aqueous MeCN *vs.* x_w obtained with different probes: the Fe complex (——), the $E_T(30)$ betaine (— — —), triethylphosphine oxide (*AN*, — — — —), 1-ethyl-4-methoxycarbonylpyridinium iodide (*Z*, - - - - - -), dialkylbenzamide (- - - - -) and pyridine *N*-oxide (· · · · · ·)

orthogonal variables α and π*, so that the plots need not be linear at all. The second view, that discrepancies in α obtained with different probes are due to the propagation of errors in the translation of the primary data for the probes *via* the regressions into α values was preferred by Cheong and Carr.[9]

It has been argued[4] that one particular probe, triethylphosphine oxide, yielding ^{31}P NMR chemical shifts that are translated into α values, does show such *systematic* deviations. Its very highly basic nature, more so than probes such as pyridine *N*-oxide or *N*,*N*-dialkylbenzamide, causes it preferentially to form hydrogen bonds with the water when the cosolvent is aprotic. With protic or protogenic cosolvents (methanol, acetonitrile), no such deviations are found.[4] This argument is weakened when more cases are examined and deviations with probes not having extreme properties are also noted. The tentative conclusion is that α for aqueous mixtures cannot be defined to better than ±0.12 in the general case. Competition between the cosolvent and a solute (eventually, a probe) for hydrogen bonding with the water in aqueous solvent mixtures makes the concept of 'the propensity of the mixture to donate hydrogen bonds' relatively vague. This situation does not appear to be ameliorated by the use of several probes, contrary to the cases of π* and β.

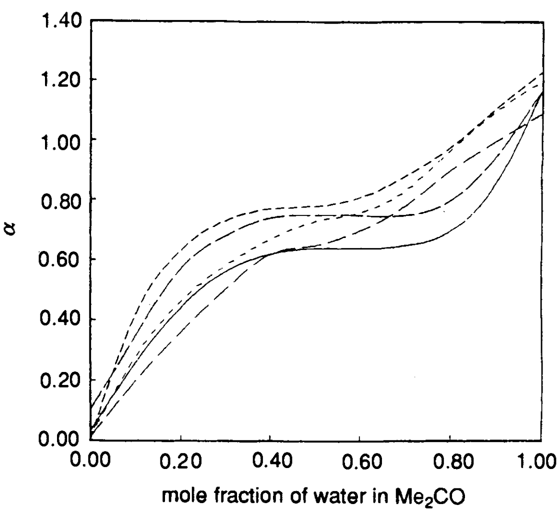

Fig. 7 Plots of α for aqueous Me_2CO *vs.* x_w obtained with different probes: the Fe complex (——), the $E_T(30)$ betaine (— — —), triethylphosphine oxide (*AN*, — — — —), 1-ethyl-4-methoxycarbonylpyridinium iodide (*Z*, - - - - - -), dialkylbenzamide (- - - - - -) and pyridine *N*-oxide (· · · · · ·)

If one sticks to a single probe, such as the betaine for the $E_T(30)$ measurements, one has a definite scale of polarity with which to characterize the aqueous mixtures. Many quantities (such as the rate constant for a given reaction) correlate well with these $E_T(30)$ values. In such cases, different mixtures having the same $E_T(30)$ should give the same results (*i.e.*, rate constants). However, their interpretation with respect to the interactions with the water and the cosolvent must elude the investigator. They may be due to the donation of hydrogen bonds by the water to the solutes (when the cosolvent does not compete too strongly for them) or to polar interactions of highly dipolar cosolvents with the solutes, or to induced dipoles in very polarizable cosolvents by the solutes, or more probably to some combination of all these effects.

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