Enthalpies of Transfer of Pyridines from the Gas Phase to Solution in Weakly and Moderately Polar Aprotic Solvents

Irmina Uruska* and Małgorzata Koschmidder
Department of Physical Chemistry, Technical University of Gdańsk, 80-952 Gdańsk, Poland

The enthalpy of solution of pyridine and of its methyl-, chloro-, and cyano- derivatives has been measured calorimetrically in some aliphatic and aromatic hydrocarbons and in carbon tetrachloride. The enthalpy of transfer from the gas phase to a given solvent has been calculated and is discussed in terms of specific and non-specific solute–solvent interactions. It is shown that dispersion forces play the dominant role in the solvation process; the contribution of dipole–dipole, dipole–induced dipole, and specific interactions seem not to exceed 20% of the overall effect. The specific interactions involve primarily the π-electron system of the pyridines. No correlation of the enthalpies of solvation with the *N*-donor ability of the pyridines has been found. The effect of substituent groups on the enthalpies of solvation is approximately additive.

In the past decade there has been considerable interest in the thermodynamics of solvation of organic non-electrolytes. Intensive studies have resulted in useful generalizations for different classes of solutes and solvent media. However, pyridine derivatives were found not to obey the generalized correlation equations, *e.g.* the LSERs, without special corrections which at present do not find any justification.[1] This behaviour, and the fact that they are prototypes of heterocyclic aromatics of biological importance, make them an especially interesting class of compounds for an investigation of solute–solvent interactions. Much attention has been paid to the thermodynamics of solvation of pyridine derivatives in strongly polar and hydroxylic solvents,[2,3] but no systematic studies on their interactions with weakly polar media have been published. In this work we report calorimetric enthalpies of solution and calculated enthalpies of solvation for eight pyridine derivatives of different basicity (pK_a 0.7–6.7) and steric properties in non-protic aliphatic and aromatic hydrocarbons and in carbon tetrachloride.

Experimental

Reagents.—Solvents were purified by standard methods as described previously;[4] Pyridine (Py), 2-methylpyridine (2-Mepy), 2,4- and 2,6-dimethylpyridine (2,4-Me_2py and 2,6-Me_2py), and 2- and 3-chloropyridine (2-Clpy and 3-Clpy), Fluka AG, were dried over KOH and distilled from molecular sieves (4A) *in vacuo*. 3- and 4-cyanopyridine (3-CNpy and 4-CNpy), Fluka AG, were recrystallized twice from hexane–benzene mixtures. All reagents were kept and manipulated in a dry-box desiccated with P_2O_5.

Calorimetric Measurements.—Calorimetric measurements were performed as outlined in ref. 4, in moisture-free conditions using a constant-temperature environment type of calorimeter with reaction ampoules made of Teflon and thin-walled glass. The temperature was measured with a 33 kΩ thermistor in a Wheatstone bridge ($\pm 7 \times 10^{-4}$ K). Electrical calibration was performed for each experiment. The device has been tested by measurement of the known enthalpy of dissolution of KCl in water (17.60 ± 0.04 kJ mol^{-1}, lit.,[5] 17.59 kJ mol^{-1}). Less than 1 g of the amine or amine solution was added to *ca.* 90 cm^3 of solvent. A correction for the free volume in the calorimeter was made; this was very small (10^{-2} kJ mol^{-1}) because of the low volatility of the dissolved species at the temperature of the measurement and the comparatively small vapour space in the ampoule (<0.5 cm^3). All calorimetric measurements were carried out at least three times with variation in the solute concentration; no concentration effect on the enthalpies of solution was observed. The ΔH_s data listed in Table 1 are the average values. The uncertainty limits were calculated as standard deviations of the mean values.

Results and Discussion

As the final concentrations of the solutions are very low (5–50 mmol dm^{-3}) and no appreciable dependence of the enthalpies of solution, ΔH_s, on the solute concentration has been observed within the range of concentrations investigated, we can take the measured enthalpies of solution (Table 1) as essentially infininte dilution values and we can thus discuss some of the observed regularities in terms of solute–solvent interactions. The enthalpies of solution in very weakly solvating media such as aliphatic hydrocarbons can also supply some information on solute–solute interactions in the pure state.

Solute-solvent interactions can be best discussed in terms of the enthalpies of transfer, ΔH_{tr}, of the pyridines from the gaseous state (g) or from a medium which solvates only by dispersion interactions (cyclohexane) to the given solvent (s) at infinite dilution. The $\Delta H_{tr(g \rightarrow s)}$ values (often called solvation enthalpies), defined as $\Delta H_s - \Delta H_{vap}$, where ΔH_{vap} is the enthalpy of vapourization, are presented in Table 2; the $\Delta H_{tr(cyclohex \rightarrow s)}$ values are given in Table 3. For pyridine, 2,6-dimethylpyridine, and 3-chloro- and 4-cyano-pyridine the $\Delta H_{tr(cyclohex \rightarrow CCl_4)}$ values obtained in this work can be compared with those calculated from the enthalpies of solution measured in refs. 2 and 6; they show good agreement within experimental error, *i.e.* within *ca.* 1 kJ mol^{-1}. Table 2 also contains the $\Delta H_{tr(g \rightarrow H_2O)}$ and $\Delta H_{tr(g \rightarrow MeOH)}$ values taken from the literature[2,3] for comparison.

From Table 2 it is clearly seen that for each particular pyridine derivative the enthalpy of solvation $[\Delta \mathrm{H}_{tr(g \rightarrow s)}]$ increases monotonically in the order: cyclohexane < hexane $\ll$ CCl_4 $\simeq$ benzene $\simeq$ chlorobenzene $\ll$ methanol $\ll$ H_2O. There are some striking features in this order.

(*a*) The solute is more strongly solvated by hexane than by cyclohexane whereas the reverse order should be expected as a result of the higher polarity of the latter (a higher value of the solvatochromic parameter,[7] π^*, or any known 'polarity function'[8]). Though the difference in the $\Delta H_{tr(g \rightarrow s)}$ values for the

Table 1. Enthalpies of solution/kJ mol^{-1} of liquid and solid pyridines at 298.2 K.

Solute	C_6H_{14}	cyclo-C_6H_{12}	CCl_4	C_6H_6	C_6H_5Cl
Pyridine	8.07 ± 0.02	8.28 ± 0.08	0.76 ± 0.01	0.04 ± 0.01	0.02 ± 0.02
2-Mepy	5.76 ± 0.09	6.59 ± 0.07	−0.50 ± 0.02	0.08 ± 0.01	−0.93 ± 0.06
2,4-Me_2py[4]	5.97 ± 0.08	7.63 ± 0.12	−1.38 ± 0.08	0.46 ± 0.01	−1.13 ± 0.01
2,6-Me_2py[4]	5.22 ± 0.03	5.49 ± 0.12	−1.30 ± 0.02	0.38 ± 0.08	−1.30 ± 0.04
2-Clpy	9.04 ± 0.09	9.1 ± 0.2	3.60 ± 0.04	0.86 ± 0.02	0.72 ± 0.02
3-Clpy	7.6 ± 0.1	7.9 ± 0.3	0.94 ± 0.07	1.02 ± 0.06	0.96 ± 0.02
3-CNpy	31.7 ± 0.5	31.6 ± 0.3	22.7 ± 0.6	18.0 ± 0.7	18.2 ± 0.3
4-CNpy	42.2 ± 0.6	34.4 ± 1.2	24.1 ± 0.7	18.2 ± 0.6	20.7 ± 0.4

Table 2. Enthalpies of solvation, $\Delta H_{tr(g \rightarrow solvent)}$, and vapourization, ΔH_{vap}, of some pyridines in various solvents at 298.2 K.

			$\Delta\Delta H_{tr(g \rightarrow s)}$/kJ mol^{-1}						
Solute	pK_a	ΔH_{vap}/kJ mol^{-1}	C_6H_{14}	cyclo-C_6H_{12}	CCl_4	C_6H_6	C_6H_5Cl	MeOH[3]	H_2O[2]
Pyridine	5.2	40.15[a]	32.08	31.87	39.39	40.11	40.13	44.21	49.85
2-Mepy	6.0	42.51[a]	36.75	35.92	43.01	42.43	43.44	48.36	54.65
2,4-Me_2py	6.6	47.50[b]	41.53	39.87	48.88	47.04	48.63	—	61.67
2,6-Me_2py	6.7	45.38[b]	40.16	39.89	46.68	45.00	46.68	52.89	60.01
2-Clpy	0.7	48.8[c]	39.7	39.7	45.2	47.9	48.0	—	—
3-Clpy	2.8	49.6[c]	42.0	41.7	48.6	48.5	48.6	49.6	54.3
3-CNpy	1.4	72.1[d]	40.4	40.5	49.4	54.1	53.9	—	—
4-CNpy	2.1	73.2[d]	39.0	38.8	49.1	55.0	52.5	52.9	58.3

[a–d] Refs. 12–15, respectively.

Table 3. ΔH_{tr}/kJ mol^{-1} from cyclohexane to the solvent(s) for some pyridine derivatives at 298.2 K.

Solute/solvent(s)	C_6H_{14}	CCl_4	C_6H_6	C_6H_5Cl
Pyridine	−0.21 ± 0.1	−7.52 ± 0.03	−8.24 ± 0.03	−8.26 ± 0.04
2-Mepy	−0.83 ± 0.16	−7.09 ± 0.11	−6.51 ± 0.1	−7.52 ± 0.15
2,4-Me_2py	−1.66 ± 0.2	−9.01 ± 0.16	−7.17 ± 0.09	−8.76 ± 0.09
2,6-Me_2py	−0.27 ± 0.15	−5.79 ± 0.05	−5.11 ± 0.11	−6.79 ± 0.07
2-Clpy	−0.03 ± 0.11	−5.47 ± 0.13	−8.21 ± 0.11	−8.35 ± 0.11
3-Clpy	−0.3 ± 0.4	−6.96 ± 0.17	−6.85 ± 0.16	−6.91 ± 0.12
3-CNpy	0.1 ± 0.8	−8.9 ± 1.1	−13.6 ± 1.2	−13.4 ± 0.8
4-CNpy	−0.2 ± 1.8	−10.3 ± 1.3	−16.2 ± 1.2	−13.7 ± 1.0
Benzene[a]	−0.5 ± 0.2[b]	−3.26 ± 0.09	−3.8 ± 0.09	
Toluene[a]	−1.16 ± 0.2[b]	−3.38 ± 0.09	−2.76 ± 0.17	
Chlorobenzene[a]	—	−3.13	−3.64	−3.76
Hexane[a]	0.79 ± 0.17[b]	0.83 ± 0.09	3.64 ± 0.21	

[a] From refs. 9 and 10. [b] $\Delta H_{tr(cyclohex \rightarrow heptane)}$.

two solvents is small (comparable to the experimental error), it is nevertheless significant; see Table 3 for the ΔH_{tr} values from cyclohexane to hexane. Such behaviour may result from a higher endoergic cavity term for cyclohexane, as suggested by the Hildebrand solubility parameter (16.8 and 14.9 MPa$^{\frac{1}{2}}$, respectively[8]) or it could be due to the poorer solvating ability of the less flexible cyclohexane molecules. It is interesting that the enthalpies of solution of benzene derivatives in cyclohexane and heptane reported by Fuchs *et al.*,[9,10] show the same effect.

(*b*) The enthalpies of solvation in carbon tetrachloride, benzene, and chlorobenzene are comparable, except for the cyanopyridines in aromatic hydrocarbons. For carbon tetrachloride, benzene, and chlorobenzene we can assume that the energies of cavity formation are comparable or that they increase slightly in the order CCl_4 < benzene < chlorobenzene, in the same way as the values of the Hildebrand solubility parameter of the solvents (17.6, 18.8, and 19.4 MPa$^{\frac{1}{2}}$, respectively). It can also be expected that the enthalpy term corresponding to restructuring of the solvent after cavity formation is small for such weakly polar non-hydrogen bonding media and would not significantly influence the overall effect. With such assumptions the comparable enthalpies of solvation seem to reflect the comparable strength of the solute–solvent interactions after cavity formation in the three solvents. These can be arbitrarily divided into Van der Waals and specific interactions. While for the former a monotonic increase in the order of increasing polarity/polarizability of the medium can be expected (CCl_4 < benzene < chlorobenzene; π^* 0.28, 0.59, and 0.71, respectively),[8] the latter may contribute in different ways to the overall solvation effect. The specific interactions may involve donor–acceptor n–π^* and π–π^* (Py-aromatics) and n–σ^* (Py-CCl_4) bonds as well as hydrogen π-H bonds. In general these should be weak as may be inferred from a comparison of the enthalpies of solvation of the pyridines with those of the benzene derivatives in the same media (Table 4). It seems reasonable to assume that cavity terms and dispersion forces are nearly the same for every pair: pyridine–benzene, methylpyridine–toluene, and chloropyridine–chlorobenzene. Thus the $\Delta\Delta H_{tr(g \rightarrow s)}$ values* should reflect dipolar and specific interactions. Data from Table 4 clearly show that they are only a small part of the overall effect. We can estimate the upper limit

---

* $\Delta\Delta H_{tr(g \rightarrow s)} = \Delta H_{tr(g \rightarrow s)}(YPy) - \Delta H_{tr(g \rightarrow s)}(YBe)$.

Table 4. The difference between the enthalpies of solvation of pyridines and those of analogous benzene derivatives[b] (model compounds), $\Delta\Delta H_{tr(g\to s)}$[a,b] at 298.2 K.

$-\Delta H_{tr(g\to s)}$/kJ mol^{-1}	cyclo-C_6H_{12}	CCl_4	C_6H_6	C_6H_5Cl
Py–benzene	1.87	6.13	6.31	
2-Mepy–toluene	1.19	4.89	4.98	
2,4-Me_2py-*p*-xylene	1.13	—	5.75	
2,6-Me_2py-*p*-xylene	1.15	—	3.71	
2-Clpy–chlorobenzene	0.95	3.3	5.51	5.49
3-Clpy–chlorobenzene	2.95	6.72	6.11	6.09

[a] See definition of $\Delta\Delta H_{tr(g\to s)}$ in the footnote on p. 1846. [b] Refs. 8 and 9.

Table 5. Group contributions to the enthalpies of solvation of some pyridine derivatives, $[\Delta H_{tr(g\to s)}(Py) - \Delta H_{tr(g\to s)}(YPy)]$/kJ mol^{-1}.

Solute/solvent(s)	C_6H_{14}	cyclo-C_6H_{12}	CCl_4	C_6H_6	C_6H_5Cl
Py	0	0	0	0	0
2-Mepy	4.7	4.1	3.6	2.3	3.3
2,4-Me_2py	9.4	8.0	9.5	6.9	8.5
2,6-Me_2py	8.1	8.0	7.3	5.6	6.6
2-Clpy	7.7	7.8	5.8	8.5	7.8
3-Clpy	9.9	9.8	9.2	9.1	8.4
3-CNpy	8.4	8.6	10.3	14.7	13.7
4-CNpy	7.0	6.9	9.7	15.6	12.3

of the specific interactions, taking as their measure the difference between the $\Delta\Delta H_{(g\to s)}$ values obtained for CCl_4, benzene, and chlorobenzene and those for cyclohexane. This difference is about 10% or less of the total enthalpy of solvation of the pyridines. Unfortunately, such calculations cannot be made for cyanopyridines (which seem to interact more strongly with aromatic solvents) because appropriate data for the model compound, cyanobenzene, are not available.

The striking feature of the enthalpies of solvation $\Delta H_{tr(g\to s)}$ (Table 2), is that they show no regular dependence on the pK_a values of the pyridines. This would be expected if specific *n*–π or *n*–σ* interactions prevailed (compare the data for the methyl- and chloropyridines). The data seem, however, to support the concept of additivity of group contribution to the solvation.[9] This effect is clearly seen if we compare group increments, calculated as $\Delta H_{tr(g\to s)}(Py) - \Delta H_{tr(g\to s)}(YPy)$, Table 5, and particularly those for pyridine (taken as zero) and its methyl derivatives. Two points deserve here special attention. (*a*) The increments seem not to be significantly influenced by the medium (except for cyanopyridines in aromatic solvents): they are of nearly the same value for aliphatic 'inert' solvents and polar chlorobenzene with a reactive π-electron system. (*b*) The increments calculated for pyridine derivatives are comparable in magnitude to those reported by Fuchs *et al.*[9] for methyl- and chlorobenzenes: *ca.* 4.5 and 3.7 kJ mol^{-1} per methyl group; and 8.8 and 8.6 kJ mol^{-1} per chlorine substituent in cyclohexane and benzene, respectively.

These facts suggest that the differences in the enthalpies of solvation of the pyridine derivatives (at least those with the alkyl and chlorine groups) are determined by local substituent–solvent interactions. This conclusion as well as the lack of dependence of $\Delta H_{tr(g\to s)}$ on pK_a values shows that the presence of the lone-pair on nitrogen does not play a significant part in solvation in the systems investigated. For the same reasons, only minor interactions are likely to result from hydrogen bonding between the acidic protons of pyridines and the π-electronic system of aromatic hydrocarbons. Thus the dominant interactions should involve the π-electron system of the pyridines. This specific effect cannot easily be separated from the dipole–induced dipole interactions. We can, however, anticipate that it is comparable to that observed for enthalpies of solvation of benzene derivatives in the same solvents. For benzene, toluene, xylene, and chlorobenzene the difference between the enthalpies of solvation in benzene and toluene and those in cyclohexane is nearly constant at *ca.* 3 kJ mol^{-1}.[9,10] Such a difference for the pyridine derivatives is *ca.* 7 kJ mol^{-1} and is also nearly constant. It is quite reasonable to assume that half of this value (*ca.* 3 kJ mol^{-1}) is due to the presence of the aromatic ring. It is interesting that CCl_4 gives similar enthalpic effects as aromatic solvents in the solvation of both pyridines and benzenes. It is possible that this behaviour reflects comparable energies of the LUMO orbitals for CCl_4 and aromatics: for benzene, toluene, pyridine, chlorobenzene, and carbon tetrachloride they are *ca.* −8.3, −8.3, −9.3, −9.8, and −10.61 eV, respectively;[11] for comparison, the LUMO of cyclohexane is 0.09.[11]

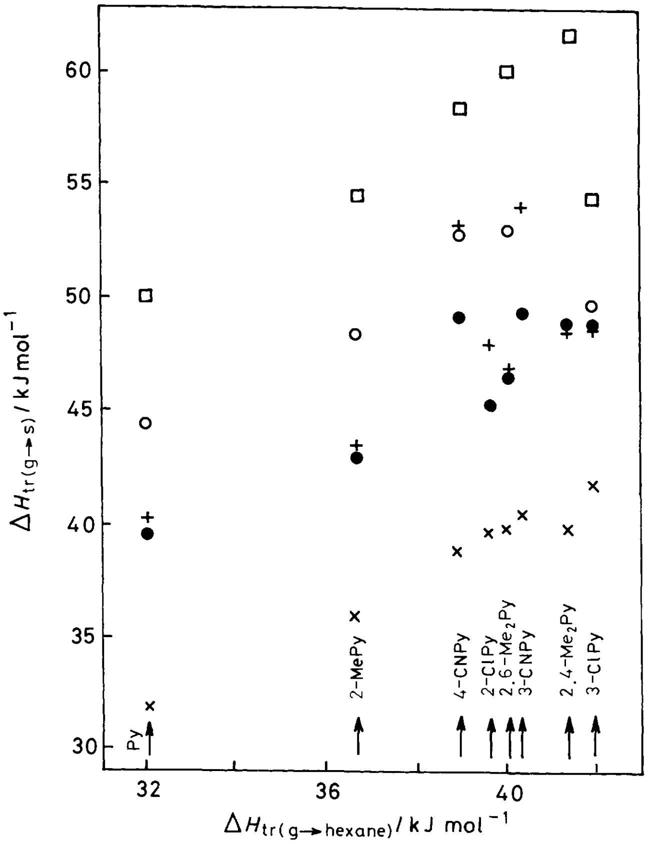

Figure. Enthalpies of transfer of some pyridine derivatives from the gas phase to cyclohexane, carbon tetrachloride, chlorobenzene, methanol,[3] and water[2] *vs.* enthalpies of transfer from the gas phase to hexane; ×, cyclohexane; ●, CCl_4; +, chlorobenzene; ○, methanol; and □, water.

From the above reasoning it follows that dipole–dipole and dipole–induced dipole forces also play only a small part in the overall effect. This can be supported by the fact that the enthalpies of transfer of the pyridines from the gaseous state to CCl_4 and benzene ($\mu = 0$) and to the pure solute (heats of vapourization of the pyridines) are nearly equal. They also show no regular dependence on the dipole moment of the amines (μ = 1.6 to 3.65 D). All these results lead to the conclusion that even in pure, fairly polar, solutes (the relative permittivity of pyridine is 12.3) the most significant interactions originate from dispersion forces. One can argue that this could be a result of a partial cancellation of the endoergic cavity and exoergic solute–solvent interaction terms in the overall solvation effect, the two terms being proportional to each other. However the enthalpies of transfer of the solutes from aliphatics (cyclohexane) to other non-protic media, which should contain only a small contribution of the cavity term (see discussion above), substantiate this conclusion.

In an attempt to correlate the enthalpies of solvation of pyridines in non-protic solvents of weak and medium polarity with those in protic polar media [2,3] we plotted the $\Delta H_{tr(g \rightarrow s)}$ values against ΔH_{tr} from the gas phase to hexane. As could be expected, a smooth correlation was found for ΔH_{tr} to cyclohexane for all pyridines investigated, with the scatter of points within experimental error. A similar correlation shows the data for carbon tetrachloride, benzene, chlorobenzene, and pure solutes as solvating media (only the plots for carbon tetrachloride and chlorobenzene are shown here for the sake of simplicity). A strong deviation of the points for cyanopyridines can be observed in this case. The intermolecular forces responsible for these deviations seem to increase in different media in the order: CCl_4 < benzene $\simeq$ chlorobenzene < pure solute, the last being a solid at normal temperature and pressure. These deviations do not however appear in the plots for cyclohexane, methanol, and water as solvents. Moreover, the Figure reveals a fairly good linear relationship between the values of ΔH_{tr} from the gas phase to hexane and those to water and methanol for all pyridines investigated except for 3-chloropyridine. Even more interesting is the fact that the slopes of the lines which could be drawn in the Figure are approximately equal to 1 (from *ca.* 0.9 for CCl_4, benzene, and chlorobenzene to *ca.* 1.05 for methanol and *ca.* 1.2 for water). This suggests that the substituent effects are similar for solvation in the different solvents in question and the hydrogen bonding between pyridines and protic solvents is not likely to be the dominant factor which determines the substituent effect on the enthalpies of solvation. Thus, our results support a similar conclusion based on another set of data, drawn by Rochester and Waters [3] for alkylpyridines in methanol.

Acknowledgements

The study was supported by Polish Academy of Sciences Grant No. CPBP 01.12-8.12.

References

1 M. J. Kamlet, R. M. Doherty, M. H. Abraham, P. W. Carr, R. F. Doherty, and R. W. Taft, *J. Phys. Chem.*, 1987, **91**, 1996.

2 E. M. Arnett, B. Chawla, L. Bell, M. Taagepera, W. J. Hehre, and R. W. Taft, *J. Am. Chem. Soc.*, 1977, **99**, 5729, and references therein.

3 C. H. Rochester and J. A. Waters, *J. Chem. Soc., Faraday Trans. 1*, 1982, **78**, 631.

4 I. Uruska and H. Karaczewska, *J. Solution Chem.*, 1979, **8**, 105.

5 E. D. Rossini, 'Selected Values of Thermodynamic Properties,' Natl. Bur. Stand (US) Circ. 500, Washington, 1952.

6 A. D. Sherry and K. F. Purcell, *J. Am. Chem. Soc.*, 1970, **92**, 6387.

7 J. L. Abboud, R. W. Taft, and M. J. Kamlet, *J. Chem. Res.*, (*S*), 1984, 98.

8 M. J. Kamlet, J. M. Abboud, M. H. Abraham, and R. W. Taft, *J. Org. Chem.*, 1983, **48**, 2877.

9 R. Fuchs, T. M. Young, and R. F. Rodewald, *J. Am. Chem. Soc.*, 1974, **96**, 4705.

10 W. K. Stephenson and R. Fuchs, *Can. J. Chem.*, 1985, **63**, 2529.

11 M. Chast-ette, M. Rajzman, M. Chanon, and K. F. Purcell, *J. Am. Chem. Soc.*, 1985, **107**, 1.

12 V. Majer, V. Svoboda, and M. Lencka, *J. Chem. Thermodyn.*, 1984, **16**, 1019.

13 B. Wiśniewska, M. Lencka, and M. Rogalski, *J. Chem. Thermodyn.*, 1986, **18**, 703.

14 Ts. B. Konunova and M. F. Franze, *Russ. J. Inorg. Chem.* (*Engl. Transl.*), 1973, **18**, 95.

15 J. Bickerton, G. Pilcher, and G. Al-Takhin, *J. Chem. Thermodyn.*, 1984, **16**, 375.

Received 11*th July* 1988; *Paper* 8/02769B