Enthalpies and Entropies of Solution of Tris(3-hydroxy-1,2-dimethylpyridin-4-onato)aluminium(III) in Methanol–Water Mixtures from Temperature Dependences of Solubilities

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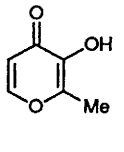
Solubilities of tris(3-hydroxy-1,2-dimethylpyridin-4-onato)aluminium(III), AlL_3, have been determined in a series of methanol–water solvent mixtures (0, 20, 40, 60, 80 and 100% methanol) over the temperature range 278.2 to 310.2 K. Solubility maxima are independent of temperature. Enthalpies and entropies of solution exhibit compensation. Maxima in thermodynamic solution parameters do not correspond to maximal solvent structural effects and are therefore assignable to solute–solvent rather than to solvent–solvent interactions.

We recently reported solubility trends for several tris(4-pyrone) (*e.g.* compound **1**) and tris(pyridin-4-one) (*e.g.* compound **2**)



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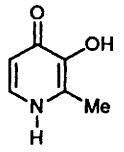

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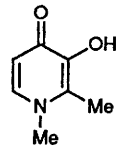

L

metal(III) complexes in methanol–water mixtures. In the case of the 4-pyrone complexes there were solubility maxima at around 70 vol. % methanol, suggesting favourable solvation by both solvent components in methanol–water solvent mixtures.[1] These results were obtained at 298.2 K. It seemed sensible, in view of the potential physiological importance of such ligands and complexes, to see whether a similar pattern existed at 310.2 K. Having decided to carry out the required experiments, it seemed a useful extension of the investigation to establish the temperature dependences of the solubilities sufficiently precisely to derive enthalpies and entropies of solution as a function of solvent composition. We have therefore measured appropriate solubilities for the aluminium complex of the pharmacologically important ligand L, 3-hydroxy-1,2-dimethylpyridin-4-one, which has successfully completed clinical trials as an orally effective chelator for the removal of iron[2] and may well prove valuable in controlling aluminium levels in the body.

Experimental

The ligand L, and its aluminium(III) complex, were prepared from compound **1** by published methods.[3] Solubilities were determined by our usual method, with spectrophotometric determination of complex concentrations.[1] We checked that the molar absorption coefficient of 30 000 $dm^3\ mol^{-1}\ cm^{-1}$ at 290 nm, originally determined in aqueous solution at 298.2 K,[1] did not vary within experimental uncertainty over the temperature and solvent composition ranges employed in this investigation. Samples consisting of a generous excess of complex plus appropriate solvent (mixture) were thermostatted for approximately 12 h, before the removal of the first aliquot for solubility determination. Mixed solvents were prepared by pipetting (at the laboratory temperature of 19–21 °C) appropriate volumes of the components. Mixed-solvent compositions are expressed by volume before mixing; 20 vol. % methanol is the result of mixing methanol (20 cm^3) with water (80 cm^3). Methanol was dried by treatment with magnesium and a trace of iodine followed by distillation.

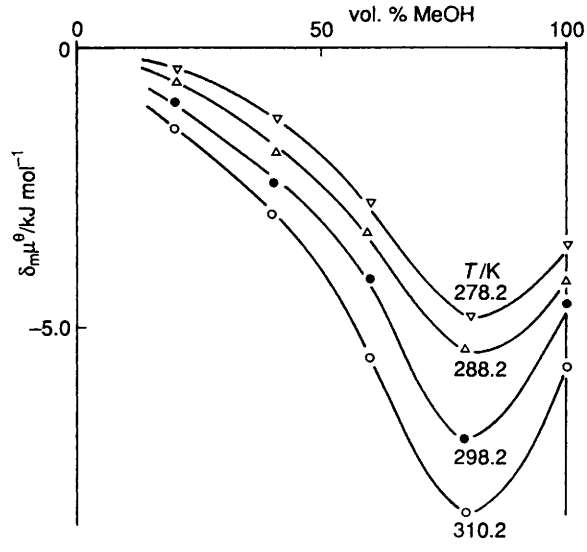

Fig. 1 Transfer chemical potentials for AlL_3 into methanol–water mixtures at temperatures in the range 278.2–310.2 K

Results

Measured solubilities are reported in Table 1. This Table also contains calculated values for transfer chemical potentials (from water), enthalpies of solution with their standard deviations, and enthalpies and temperature–entropy products of transfer (from water).

Discussion

The pattern of transfer chemical potentials for tris(3-hydroxy-1,2-dimethylpyridin-4-onato)aluminium(III), AlL_3, at the various temperatures is depicted in Fig. 1. It is clear that the trends at the various temperatures are very similar. In particular the solvent composition corresponding to maximum solubility does not change significantly with temperature. This behaviour parallels that established many years ago for picric acid in aqueous ethanol, propan-1-ol and propan-2-ol.[4]

Table 1 Solubilities (S/mol dm^{-3}) of AlL_3 in methanol–water mixtures, corresponding transfer chemical potentials (Δ/kJ mol^{-1}), and derived enthalpy (ΔH/kJ mol^{-1}, ΔH_{tr}/kJ mol^{-1}) and entropy ($T\Delta S_{tr}$/kJ mol^{-1}, T = 298.2 K) quantities

Vol. % MeOH	0	20	40	60	80	100
T = 278.2 K						
10^3 S	1.35	1.63	2.51	4.2	9.3	5.8
Δ	0	−0.5	−1.6	−2.8	−4.8	−3.6
T = 288.2 K						
10^3 S	1.98	2.39	4.6	7.4	20.6	10.7
Δ	0	−0.5	−2.0	3.3	−5.8	−4.2
T = 298.2 K						
10^3 S	2.9	4.2	7.7	15.1	49	18.1
Δ	0	−0.9	−2.4	−4.1	−7.0	−4.5
T = 310.2 K						
10^3 S	3.8	6.7	12.6	36	108	38
Δ	0	−1.4	−3.0	−5.6	−8.3	−5.7
ΔH	23.2	32.4	36.4	45.8	55.7	41.3
σ	±1.0	±1.2	±1.4	±1.6	±1.5	±1.5
ΔH_{tr}		+9.2	+13.2	+22.6	+32.5	+18.1
$T\Delta S_{tr}$		+10	+16	+27	+40	+23

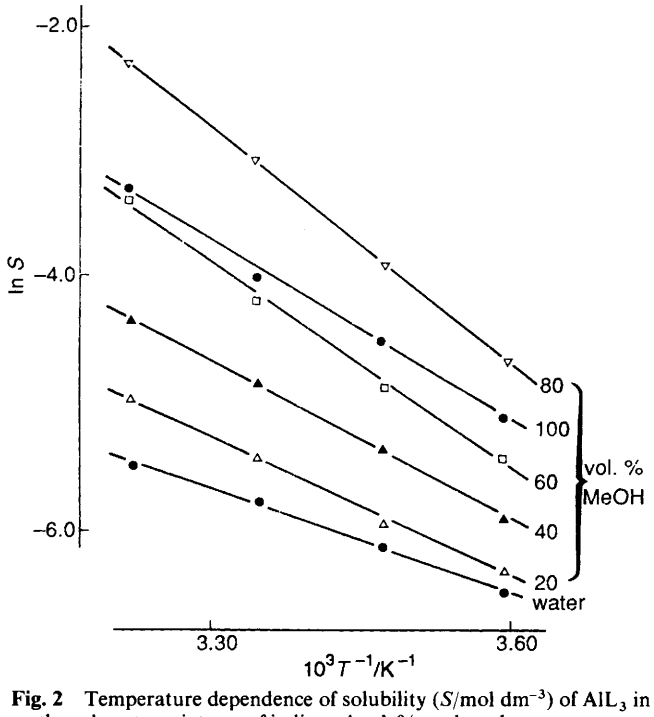

Fig. 2 Temperature dependence of solubility (S/mol dm^{-3}) of AlL_3 in methanol–water mixtures of indicated vol. % methanol

The van't Hoff plots for the aluminium complex are reproduced in Fig. 2, to show that they are linear within experimental error at every solvent composition. There is no evidence of solvent-imposed complications, *e.g.* at the lower temperatures where the methanol cosolvent effects on water structure are greater.[5] We therefore feel justified in deriving enthalpies and entropies of solution from our results (Table 1). The enthalpy and entropy results are displayed in the form of transfer functions in Fig. 3. This representation clearly shows compensation of the ΔH and $T\Delta S$ parameters resulting in much smaller changes in ΔG—as is so often observed. It is interesting to note from Fig. 3 that the extrema in the ΔG, ΔH and $T\Delta S$ values all occur in the region of 60 to 80 vol. % methanol; there

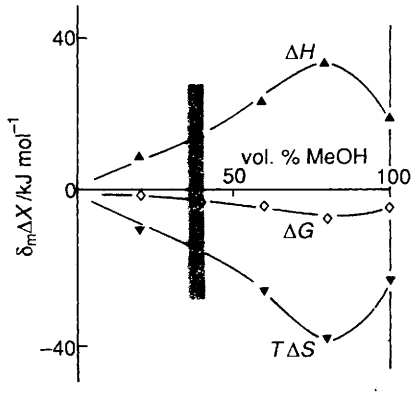

Fig. 3 Dependence of thermodynamic transfer parameters (298.2 K) for AlL_3 on solvent composition in methanol–water mixtures. The shaded bar represents the region of maximum structuredness of the solvent mixtures

are no signs of irregularities in the region of maximum structuredness (*ca.* 40% methanol)[5] of these solvent mixtures.

Figs. 1 and 3 both indicate that the solution behaviour of our aluminium complex in methanol–water mixtures is dominated by solute–solvent, *i.e.* complex solvation, rather than by solvent–solvent, interactions. There is a marked difference between the present aluminium complex and, say, $[Fe(CN)_2(bipy)_2]$ (bipy = 2,2′-bipyridine), whose solubility trends do reflect alcohol–water interactions.[6]

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