Osmotic and activity coefficient studies on the aqueous solutions of tetramethylurea at 298.15 K

K J Patil, A M Sargar & D H Dagade
Department of Chemistry, Shivaji University, Kolhapur 416 004, India

Received 10 April 2002; revised 3 July 2002

The osmotic coefficients and densities of good precision of aqueous solutions of tetramethylurea (TMU) in the concentration range 0.1 to 8.0 M at 298.15 K have been experimentally determined and the data used to calculate apparent and partial molar volumes, activities and activity coefficients of water and TMU as a function of concentration of TMU. It has been observed that partial molar volumes of solvent and solute go through a maximum and a minimum, respectively, at about 0.055 mole fraction of TMU. The activity coefficient values for H₂O increase, while those of TMU decrease with increase in concentration of TMU. These results are compared with those of aqueous urea solutions and are in agreement with literature values indicating the presence of hydrophobic interactions in H₂O-TMU solutions. Excess free energy of mixing of TMU in water has been calculated using activity data at 298.15 K and is found to decrease as a function of TMU concentration in the studied concentration range. The values for the attractive and repulsive contributions to the solute-solvent interaction have been estimated and data is compared with other non-electrolytes. Further, applying McMillan-Mayer theory of solutions to the data, osmotic second and third virial coefficients for TMU have been obtained. The values of virial coefficients, on decomposition into attractive and repulsive contributions to solute-solute interactions, indicate that pairwise interaction between TMU molecules and water structure making effect is more prominent than triplet and higher order interactions in the studied concentration range of aqueous-TMU solutions.

The thermodynamic properties of aqueous urea solutions are well-known. Considering the importance of urea as denaturing agent for aqueous bio-polymer solutions, many basic studies were conducted to obtain information about molecular interaction in solutions. It has been postulated that urea mixes ideally with a postulated 'dense' (non-hydrogen bonded) species of water with a slight structure breaking effect via H-bond formation i.e. solute-solvent interactions. Stokes has explained the concentration dependence of the thermodynamic properties in terms of association of urea molecules. There is much controversy regarding the interpretation of urea acting as structure-making or structure-breaking solute and a lot depends upon the unique structural properties of water. It is known that the alkyl substitution in a molecule alters the thermodynamic properties of solutions and the effects can be explained in terms of hydrophobic hydration and interactions. Methyl substitution on amino group in urea has three effects: (i) contribution from methyl group itself (intrinsic and hydrophobic), (ii) hindrance of H-bonding ability of the substituted amino group and (iii) a decrease of H-bonding ability of the carbonyl group as indicated by the dipole moments. Therefore, it is interesting to compare thermodynamic properties of urea-water solutions with that of aqueous solutions of tetramethylurea so that useful information about solute-solvent and solute-solute, and hence hydrophobic interaction in aqueous solutions can be derived.

Tetramethylurea is a typical liquid at ordinary conditions, also is an excellent solvent for aromatic organic compounds. It is soluble in water at all proportions. It has been reported that ultrasonic absorption, density and viscosity properties show a maximum at about 20 mole % of TMU in water and attributed to the formation of molecular complexes of stoichiometry of 4:1 (water-TMU). Recently, the small angle neutron scattering studies in aqueous solutions of TMU have indicated the presence of hydrophobic interaction, which increase with temperature. Computer simulation of liquid TMU and its aqueous solution studies reveal that free energy of hydration is much less negative than urea-water solutions and the influence of TMU on energetics of water-water H-bonding is negligible. We have earlier detected hydrophobic association using NMR relaxation and self-diffusion coefficient data in aqueous solutions of TMU. Thus, TMU acts as hydrophobic solute in aqueous solutions.

The nature of solute-solute interactions in mixtures can be investigated by applying lattice and other statistical mechanical theories. The McMillan-
Mayer formalism appears at first sight to be an attractive one, as the pair and multiplet interaction parameters which can be extracted from thermodynamic data can, under favourable conditions, provide direct estimates of the different types of solute interactions. The theory has been successfully applied to solutions of alcohols\textsuperscript{16}, carbohydrates\textsuperscript{17} and of electrolytes\textsuperscript{18}. Recently, we have made such an analysis for aqueous solutions of 18-crown-6\textsuperscript{19}. To understand the thermodynamic behaviour of a typical hydrophobic solute like TMU in water, from which solute-solvent as well as solute-solute interactions can be investigated, precise and accurate activity coefficient data along with partial molar volume data are required. In this context, we are reporting in the present paper the osmotic, activity coefficient and volumetric studies of aqueous TMU solutions at 298.15 K. The data is used to estimate the attractive and repulsive contributions to the solute-solvent interaction, the magnitude of which can be used to interpret the solute-solvent interactions in terms of H-bonding and hydrophobic hydration. The data is also subjected to the analysis using McMillan-Mayer theory of solutions\textsuperscript{20} and the approach evolved by Kozak \textit{et al.}\textsuperscript{11}. The information about pairwise and triplet interactions is derived and compared with other non-electrolytes like urea, sucrose, alcohols, etc.

\section*{Materials and Methods}

Tetramethylurea (>99\% pure) procured from Merk-Schuchardt was used without further purification. The salt NaCl of AR grade (BDH) was dried under vacuum for 24 hours before use. All the solutions were prepared in fresh doubly glass-distilled water on molality basis and converted to molarity scale when required using the density data at 298.15 K.

The density measurement of aqueous TMU solutions were made using Anton Paar digital densitometer (model DMA 60/602) at 298.15 \pm 0.02 K. The reproducibility of the values was found to be better than \pm 1 \times 10^{-2} \text{ kg m}^{-3}.

The osmotic coefficients ($\phi$) of aqueous TMU solutions were determined using KNAUER K-7000 vapour pressure osmometer at 298.15 \pm 0.001 K. The instrument was kept in specially fabricated refrigerated thermostat working between 0 to 40°C. The osmometer was calibrated using aqueous NaCl solutions for the measurement of aqueous TMU solutions. The accuracy in $\phi$ measurements was found to be better than \pm 1 \times 10^{-3} at the lowest concentration studied.

\section*{Results}

The apparent molar volumes ($\phi_0$) of TMU in water were obtained using the density data in the concentration range \sim 0.1 to 8.0 m using the usual equation. These data are in good agreement with those reported earlier\textsuperscript{21}. The partial molar volume ($\bar{V}_2$) of TMU at various concentrations were estimated using Eq. (1)

$$\bar{V}_2 = \phi_0 + m \left( \frac{d\phi_0}{dm} \right)$$  \hspace{1cm} ... (1)

The parameter ($d\phi_0/dm$) was estimated by differentiating the appropriate polynomial of the plot of $\phi_0$ against $m$. The partial molar volumes of water ($V_i$) in aqueous TMU solutions at different concentrations were calculated using Eq. (2) given below

$$\bar{V}_i = \frac{1}{n_i} \left[ n_i V_i^0 - m^2 \left( \frac{d\phi_0}{dm} \right) \right]$$  \hspace{1cm} ... (2)

where $n_i$ is the number of moles of water and $V_i^0$ is the molar volume of pure water. The data for $\bar{V}_2$ and $\bar{V}_i$ along with the density data are collected in Table I and their variation as a function of concentration of tetramethylurea is shown in Fig. 1. The smooth extrapolation of the curve $\bar{V}_2$ against concentration of TMU in Fig. 1 to infinite dilution yielded the value of limiting partial molar volume ($\bar{V}_2^0$) of TMU. The errors involved in $\phi_0$ and $\bar{V}_2$ values are of the order of \pm 0.1 \times 10^3 and \pm 0.5 \times 10^3 mm^3 mol^{-1}, respectively at the lowest concentration.

The osmotic coefficients ($\phi$) of aqueous tetramethylurea solutions were determined over the range 0.1 to 8.0 mol kg^{-1} at 298.15 K. Since the TMU is high boiling liquid (bp=449.65 K), its contribution to total vapour pressure over the solution at 298.15 K is assumed to be negligible. The data are well represented by Eq. (3)

$$\phi = 1-0.5267 m + 0.4356 m^2 - 0.1811 m^3$$  
$$+ 0.0382 m^4-0.0039 m^5 + 0.0002 m^6$$  \hspace{1cm} ... (3)

and are collected in Table 2.
The solvent activity coefficients were calculated from the experimental osmotic coefficient data using expression

$$\phi = -\left\{ \ln a_1/(x_1/x_2) \right\} \quad \ldots \quad (4)$$

where \(x_1\) and \(x_2\) are the mole fractions of solvent and solute respectively, and \(a_1\) is solvent activity. The activity coefficient of solute (\(\gamma_2\)) has been calculated using Eq. (5)

$$\ln \gamma_2 = (\phi - 1) + \frac{\mu}{(\phi - 1)} \, d \ln m \quad \ldots \quad (5)$$

Since the osmotic coefficient is expressed as a power

| Table 1—Volume properties of TMU + H₂O system at 298.15 K |
|-----------------|-----------------|-----------------|-----------------|-----------------|-----------------|
| \(m\) mol kg⁻¹  | \(c\) mol dm⁻³  | \(d\) kg m⁻³    | \(10^{-3} \times \phi\) mm³ mol⁻¹ | \(10^{-3} \times \bar{V}_2\) mm³ mol⁻¹ | \(10^{-3} \times \bar{V}_1\) mm³ mol⁻¹ |
| 0.0000          | 0.0000          | 997.047         | 114.9            | 114.9           | 18.07           |
| 0.1018          | 0.1003          | 997.218         | 114.8            | 114.7           | 18.07           |
| 0.2098          | 0.2043          | 997.410         | 114.7            | 114.5           | 18.07           |
| 0.4854          | 0.4586          | 997.943         | 114.5            | 114.2           | 18.07           |
| 0.9993          | 0.8945          | 999.068         | 114.2            | 113.7           | 18.08           |
| 1.5592          | 1.3207          | 1000.416        | 113.9            | 113.2           | 18.09           |
| 2.9486          | 2.2048          | 1003.859        | 113.4            | 112.6           | 18.13           |
| 4.4758          | 2.9655          | 1007.049        | 113.1            | 112.7           | 18.12           |
| 5.7053          | 3.4612          | 1008.718        | 113.1            | 113.2           | 18.05           |
| 6.7719          | 3.8273          | 1009.758        | 113.2            | 113.7           | 17.94           |
| 7.7924          | 4.1326          | 1010.385        | 113.3            | 114.3           | 17.80           |

*Values at infinite dilution.*

| Table 2—Osmotic and activity coefficients of TMU + H₂O system at 298.15 K |
|-----------------|-----------------|-----------------|-----------------|-----------------|-----------------|
| \(m\) mol kg⁻¹  | \(\phi\)         | \(a_1\)         | \(\gamma_1\)    | \(\gamma_2\)    | \(\Delta G^F\) J mol⁻¹ |
| 0.1018          | 0.9846          | 0.9982          | 1.0000          | 0.9042          | -0.39          |
| 0.2098          | 0.8836          | 0.9967          | 1.0004          | 0.8233          | -0.75          |
| 0.4854          | 0.8181          | 0.9929          | 1.0016          | 0.6822          | -4.40          |
| 0.9993          | 0.7735          | 0.9862          | 1.0039          | 0.5498          | -16.69         |
| 1.5592          | 0.7439          | 0.9793          | 1.0068          | 0.4828          | -32.9          |
| 2.9486          | 0.6993          | 0.9635          | 1.0147          | 0.3920          | -82.7          |
| 4.4758          | 0.6838          | 0.9463          | 1.0227          | 0.3322          | -152.4         |
| 5.7053          | 0.6945          | 0.9311          | 1.0268          | 0.3144          | -207.9         |
| 6.7719          | 0.7203          | 0.9159          | 1.0276          | 0.2996          | -264.7         |
| 7.7924          | 0.7592          | 0.8989          | 1.0251          | 0.2930          | -320.8         |
series on the molality scale of the solute by Eq.(6)

$$\phi = 1 + \sum_{i=1}^{n} A_i m^i$$ ... (6)

where the coefficient $A_i$ can be obtained by the method of least squares, Eq. (5) takes the form, after solving the right hand side integral, as

$$\ln \gamma_2 = \sum_{i=1}^{n} \left( \frac{i+1}{i} \right) A_i m^i$$ ... (7)

The data for the activity coefficients of solvent and solute at various concentrations are collected in Table 2. Figures 2a and 2b respectively represent the variation of activity coefficient of solvent and solute as a function of concentration of solute at 298.15 K. It is observed that the activity coefficient of solvent increases, while that of solute decreases as the solute concentration is increased.

The activity data, which have been converted into the mole fraction scale, were used to calculate the excess free energy change of mixing ($\Delta G^E$) of TMU with water using Eq. (8)

$$\Delta G^E = RT \left( x_1 \ln \gamma_1 + x_2 \ln \gamma_2 \right)$$ ... (8)

where $x_1$ and $x_2$ are the mole fractions of solvent and solute respectively. It is observed that the values are negative and decreases with increase in concentration of TMU as shown in Fig. 3. The data are also included in Table 2.

**Discussion**

Examination of data in Table 1 and Fig. 1 reveals that $\gamma_1$ and $\gamma_2$ goes through a maximum and a minimum respectively at -0.055 mole fraction of TMU. Similar observations are well known for monofunctional solutes like alcohols, amines and ethers as well as for the aqueous solutions of tetraalkylammonium salts. The partial molar volume study of aqueous 18-crown-6 solutions at 298.15 K also shows such behaviour. These observations indicate that the hydrophobic hydration is predominant in the studied concentration range. The negative slope in the lower concentration region can be attributed to the reduction in hydrophobic hydration per mole of solute through an overlap of the co-spheres (solute-solvent-solute interactions). The increase of $\gamma_2$ at high concentration range gives an indication of TMU-TMU interactions in solution. The limiting excess partial molar volume ($\overline{V}_2^{ex} = \overline{V}_2^o - V_2^o$, where $V_2^o$ is the molar volume of pure TMU at 298.15 K) of TMU is found to be negative and of magnitude $5.6 \times 10^3$ mm$^3$mol$^{-1}$. The negative
magnitude is comparable with monofunctional non-electrolytes like t-butanol, higher amines etc. and indicates the extent of solute-solvent interaction, which occur with a volume loss for the solute.\(^27\)

The solute-solvent cluster integral, \( b_1 \), is related to the partial molecular volume of the solute at infinite dilution by\(^28-30\)

\[
b_1^0 = -v_2^0 + kT \kappa
\]

where \( k \) is the Boltzmann constant, \( T \) is the absolute temperature and \( \kappa \) is the isothermal compressibility coefficient of the pure solvent. The values for solute-solvent interaction \( NB_1^{0*} \) (where \( B_1^{0*} = -b_1^0 \)) for aqueous TMU solutions calculated using Eq. (9) are given in Table 3a. The values for other non-electrolytes in aqueous solutions at 298.15 K are also included in Table 3a for comparison.

The solute-solvent cluster integral in the above equation is related to the potential of mean force (\( \omega \)) between one molecule of solute and one of solvent in the pure solvent by the expression (10)

\[
b_1^0 = -4\pi \int_0^R [1 - \exp(-\omega/\kappa T)] r^2 dr
\]

where \( r \) is the distance apart of the centres of the molecules. It was shown that this integral could be split into attractive and repulsive parts as

\[
B_1^{0*} = 4\pi \int_0^R [1 - \exp(-\omega/\kappa T)] r^2 dr
+ \int_0^R [1 - \exp(-\omega/\kappa T)] r^2 dr
= S + \Phi^A
\]

where \( R \) is the distance of closest approach of the two molecules, \( S \) is the repulsive and \( \Phi^A \) is the attractive contribution.

If the form of potential \( \omega \) (ref.11) is known, then the integration could be performed to yield \( B_1^{0*} \). The simplest potential function regards the molecules as rigid spheres. For two hard spheres of diameters \( R_1 \) and \( R_2 \),

\[
S = \frac{\pi}{6} (R_1 + R_2)^3
\]

The water molecule can be considered to be a sphere of diameter 0.304 nm (although one may use the diameter of water as 0.276 nm, we retained the value equal to 0.304 nm as the comparison of the data for other solutes can be made) and taking the diameter of TMU molecule as 0.725 nm (calculated on the basis of molar volume and assuming a spherical shape), the repulsive contribution to the solute-solvent interaction comes out as \( NS = 344.2 \text{ cm}^3\text{ mol}^{-1} \) and the attractive contribution at 298.15 K can be obtained as

\[
N\Phi^A = NB_1^{0*} - NS = -230.2 \text{ cm}^3\text{ mol}^{-1}
\]

The data of attractive and repulsive contributions to the solute-solvent interaction in water for other H-bonding non-electrolytes at 298.15 K are given in Table 3a for comparison. The attractive contribution increases with increase of H-bonding sites. The value of \( \Phi^A \) for TMU is much larger than that for urea indicating the stabilization of the structure of solvent in aqueous TMU solutions. Urea-water hydrogen bond interactions exist but may be short-lived. The

<table>
<thead>
<tr>
<th>Solvent Combination</th>
<th>(10^3 \times \overline{V}_2^0) mm mol(^{-1})</th>
<th>(10^3 \times kT \kappa) mm mol(^{-1})</th>
<th>(10^3 \times NB_1^{0*}) mm mol(^{-1})</th>
<th>(10^3 \times NS) mm mol(^{-1})</th>
<th>(10^3 \times (-N\Phi^A)) mm mol(^{-1})</th>
</tr>
</thead>
<tbody>
<tr>
<td>TMU + H(_2)O</td>
<td>114.9</td>
<td>1.11</td>
<td>113.8</td>
<td>344</td>
<td>230.2</td>
</tr>
<tr>
<td>18C6 + H(_2)O*</td>
<td>223.2</td>
<td>1.11</td>
<td>222.1</td>
<td>699</td>
<td>477</td>
</tr>
<tr>
<td>Sucrose + H(_2)O*</td>
<td>211.5</td>
<td>1.11</td>
<td>210.4</td>
<td>476</td>
<td>266</td>
</tr>
<tr>
<td>Glucose + H(_2)O*</td>
<td>112.2</td>
<td>1.11</td>
<td>111.1</td>
<td>358</td>
<td>246</td>
</tr>
<tr>
<td>Hexamethylenetetramine + H(_2)O*</td>
<td>110.6</td>
<td>1.11</td>
<td>109.5</td>
<td>300</td>
<td>191</td>
</tr>
<tr>
<td>Urea*</td>
<td>44.2</td>
<td>1.11</td>
<td>43.1</td>
<td>176</td>
<td>143</td>
</tr>
</tbody>
</table>

\# Data from reference 19; \* Data from reference 30
magnitude of $\Phi^1$ is small compared to other non-electrolytes like sucrose but is comparable to hexamethylenetetramine and glucose having more H-bonding sites than TMU. Thus along with H-bonding with oxygen atom of the carbonyl group of TMU, there is large hydrophobic hydration. The minimum and maximum observed in partial molar volume of TMU at $\sim 0.055$ mole fraction indicate that in aqueous TMU solutions $1:17$ type clathrate hydrates equilibria may exist supporting the estimated $\Phi^1$ value which is larger than the expected one for the available H-bonding sites.

In aqueous tetramethylurea solutions, the logarithms of activity coefficients of water at 298.15 K are positive in the studied concentration range indicating the positive deviation from ideality. Such type of positive deviation from ideality was also reported in case of alcohol-water mixtures. The negative $\Delta G^F$ obtained in aqueous solutions may be indicative of H-bonding of water molecules with carbonyl oxygen. However, we restrained to discuss it in detail, since the data of excess enthalpy and excess entropy for solutions are not available.

According to the theory of McMillan and Mayer, for a solution of a solute in a solvent, the osmotic pressure, $\pi$, is given by

$$\frac{\pi}{kT} = n + B_2^0 n^2 + B_3^0 n^3 + \ldots$$  \hspace{1cm} \ldots (13)

where $n$ is the number density. It can be shown that the osmotic second and third virial coefficients, $B_2^*$ and $B_3^*$ respectively, for the osmotic pressure of non-electrolyte solutions can be calculated from experimental activity data and the partial molar volume of solute and solvent as

$$B_2^* = 4N \left[ \frac{\bar{V}_2^0}{V_1^0} - V_1^0 \left( \frac{1 - B}{2} \right) \right]$$  \hspace{1cm} \ldots (14)

$$B_3^* = \frac{1}{N^2} \left[ \frac{b}{V_1^0} + g + (V_1^0 - \bar{V}_2^0)^2 - V_1^0 (1 - 2B)(V_1^0 - \bar{V}_2^0) + (V_1^0)^2 \left( \frac{1}{3} - C \right) \right]$$  \hspace{1cm} \ldots (15)

where $V_1^0$ and $\bar{V}_2^0$ are the partial molar volumes of solvent and solute respectively at infinite dilution, $b$ and $g$ are the coefficients in equation of expansion of partial molar volume of solvent and solute at concentration $C$ mol dm$^{-3}$ given as

$$\bar{V}_1 = V_1^0 + aC + bC^2$$  \hspace{1cm} \ldots (16)

$$\bar{V}_2 = V_2^0 + gC + hC^2$$  \hspace{1cm} \ldots (17)

The $B$ and $C$ coefficients appearing in Eq. (18) are the coefficients in the following Eq. (18) in which the solvent activity coefficient is expressed as a power series in the mole fraction scale of solute as,

$$\ln \gamma_1 = Bx_1^2 + Cx_1^3 + \ldots$$  \hspace{1cm} \ldots (18)

The $B$ and $C$ coefficients appearing in Eq. (18) with their sign and magnitude are of special importance in understanding the thermodynamic behaviour since they are related to solute-solvent association, solute-solute interactions and solute size. Kozak et al. have shown that the osmotic second virial coefficient can be decomposed into attractive and repulsive components. The minimum attractive ($A_{2\text{min}}$) and minimum repulsive ($R_{2\text{min}}$) contributions to the osmotic second virial coefficient have been calculated by following equations

$$R_{2\text{min}} = \frac{f}{2} (4 \bar{V}_2^0)$$  \hspace{1cm} \ldots (19)

$$A_{2\text{min}} = R_{2\text{min}} - NB_2^*$$  \hspace{1cm} \ldots (20)

where $f$ is the factor which is measure of the ellipticity of the molecule. For spherical molecule $f$ is unity. In much the same way as the second virial coefficient, the third virial coefficient can be decomposed into attractive and repulsive components. The minimum attractive ($A_{3\text{min}}$) and minimum repulsive ($R_{3\text{min}}$) contributions of these two components have been calculated using Eqs (21) and (22).

$$R_{3\text{min}} = 10 \bar{V}_2^0$$  \hspace{1cm} \ldots (21)

$$A_{3\text{min}} = R_{3\text{min}} - N^2 B_3^*$$  \hspace{1cm} \ldots (22)

The values of osmotic second and third virial coefficients as well as the minimum attractive and repulsive contributions to the solute-solute interactions are given in Table 3b. The value for osmotic second virial coefficient of TMU is estimated to be $+35.1$ cm$^3$ mol$^{-1}$ and is large positive than the
value reported earlier by Bezzabotno et al. as small negative using small angle neutron scattering data. The data for other non-electrolytes in aqueous medium, collected from the literature, are also included in Table 3 for comparison. The discrepancy in $B_2^*$ value obtained in the present work with that obtained from neutron scattering data may be attributed to the limitation of spectroscopic techniques to obtain infinitely dilute solution properties.

The minimum attractive contribution to the solute-solute interactions for TMU in aqueous solutions is found to be $-424.5\ \text{cm}^3\text{mol}^{-1}$, which is a measure of pairwise interactions between two TMU molecules in water at 298.15 K. Since the magnitude of the attractive contribution increases with the number of functional groups capable of H-bonding as well as hydrophobic interactions with another solute molecules, it reveals that one TMU molecule may interact strongly with another TMU molecule by means of hydrophobic interaction. Comparing the magnitude of minimum attractive contributions to solute-solute interactions for aqueous TMU solutions to that of other aqueous non-electrolytes at 298.15 K, it seems that the solute-solute pairwise interactions in aqueous TMU solutions are larger than those in urea, amino acids and amides but are of smaller magnitude than the polyfunctional solutes like sucrose.

A relative importance of pairwise verses triplet interactions can be derived by writing

$$\Gamma = 12(NB_2^* - 3N^2B_3^*)$$

where $\Gamma > 0$ signifies that triplet clustering is more pronounced than that would be predicted from the summation of three pairwise interactions. For tetramethyleurea $\Gamma$ is negative indicating that the pairwise interactions are predominant than the triplet and higher order clustering. Comparing the magnitude of $A_{2\min}$ and $A_{3\min}$ which have reflected in the value of $\Gamma$ for urea and TMU in their aqueous solutions at 298.15 K reveals that favourable pairwise interactions in aqueous TMU solutions are due to the increased number of hydrophobic alkyl groups.

### References


21 Patil K J, Unpublished data.


