Available online at www.sciencedirect.com

SCIENCE DIRECT®

Thermochimica Acta 411 (2004) 91–94

thermochimica acta

www.elsevier.com/locate/tca

Calorimetric determination of enthalpies for the proton ionization of *N*,*N*-bis[2-hydroxyethyl]-2-aminoethanesulfonic acid (BES) and *N*-tris[hydroxymethyl]methyl-2-aminoethanesulfonic acid (TES) in water–methanol mixtures

B.N. Bulos [b], F.H. Jumean [a,*]

[a] *Department of Biology, Chemistry and Environmental Sciences, American University of Sharjah, P.O. Box 26666, Sharjah, United Arab Emirates*
[b] *Department of Physics, American University of Sharjah, P.O. Box 26666, Sharjah, United Arab Emirates*

Received 19 June 2003; received in revised form 2 August 2003; accepted 4 August 2003

Abstract

The enthalpies of proton ionization of the biochemical buffers *N*,*N*-bis[2-hydroxyethyl]-2-aminoethanesulfonic acid (BES) and *N*-tris[hydroxymethyl]methyl-2-aminoethanesulfonic acid (TES) were obtained in water–methanol mixtures in which the methanol mole fraction (X_m) varied in the range 0–0.36. For both buffers, ionization enthalpy for the first proton (ΔH_1) was small in all solvent media. However, upon addition of methanol, ΔH_2 increased steadily from 22.2 to a maximum of 27.2 kJ mol^{-1} for BES, whereas for TES it varied from 30.0 to 32.4, with a minimum of 28.6 kJ mol^{-1} at $X_m = 0.123$. It is noteworthy that this solvent composition lies within the region of maximum structure enhancement of water by methanol. The results were interpreted in terms of methanol–water interactions.

Keywords: BES; TES; Proton ionization; Calorimetry

1. Introduction

This paper deals with the enthalpies of proton ionization of *N*,*N*-bis[2-hydroxyethyl]-2-aminoethanesulfonic acid (BES) and *N*-tris[hydroxymethyl]methyl-2-aminoethansulfonic acid (TES), zwitterionic biochemical buffers that ionize in two steps. Measurements were made in water–methanol mixtures containing up to 50% (w/w) methanol, corresponding to a methanol mole fraction (X_m) in the range 0–0.36. To date, most thermodynamic measurements in mixed solvents have had the aim of assessing the influence of solvent composition on pK and obtaining free energies of proton transfer from water to the solvent system. Associated ΔH values were obtained by potentiometric or spectrophotometric methods. Examples are found in [1–6]. There are as yet no reports of studies primarily concerned with calorimetric determination of enthalpies of proton ionization of biochemical buffers in mixed solvents.

2. Experimental

Calorimetric determinations were made using a Parr 1455 solution calorimeter. BES and TES were Sigma chemicals. Spectro grade methanol was used. The recrystallized, carefully dried and finely ground solid was placed in a Teflon dish enclosed within a sealed glass rotating cell. When a push rod was depressed, the entire contents of the dish dropped into a glass Dewar. The signal from a thermistor immersed in the Dewar was fed via a hyperlink to a computer that monitored the temperature as a function of time. The temperature–time curve was analyzed by extrapolating pre-reaction and post-reaction baselines, followed by drawing a vertical intercept through the point on the temperature–time curve at which 1/*e* of the total change remained. ΔT was the length of the intercept connecting the baselines. The heat capacity (C_p) of the calorimeter was determined under the same conditions as those used for the runs. The initial solution temperature was ca. 21 °C. ΔT varied in the range 0–0.2 °C for the addition of BES and TES to the various solutions but reached 1.4 °C in neutralization runs. ΔT was read to an estimated precision of ±0.002 °C.

---

* Corresponding author. Tel.: +971-6-515-2407; fax: +971-6-515-3444.
E-mail address: fjumean@ausharjah.edu (F.H. Jumean).

0040-6031/$ – see front matter © 2003 Elsevier B.V. All rights reserved.
doi:10.1016/j.tca.2003.08.003

For all runs, including those for the determination of the enthalpies of dilution and neutralization, the ionic strength was adjusted to 0.300 M by the addition of NaCl. The experimental procedure followed was similar to that described by Ramette [7] for enthalpy determinations of glycine proton ionizations in water.

3. Results and discussion

The starting materials in each case were the zwitterionic forms of BES and TES:

$(HOCH_2CH_2)_2N^+HCH_2CH_2SO_3{}^-$ and
$(HOCH_2CH_2)_3C^+NH_2CH_2CH_2SO_3{}^-$

For simplicity, these species are represented by $HB^{\pm}$. A weighed amount of solid $HB^{\pm}$ (ca. 0.5 g) is dropped from the sample compartment into 100 g of 0.300 M HCl solution contained in the Dewar. The observed temperature change is due to the reaction:

$$HB^{\pm}(c) + H^+ \rightarrow H_2B^+(S), \qquad \Delta H_A \tag{1}$$

Here, (S) represents the mixed solvent. Since the ionization steps occur in the mixed solvent, it was necessary to obtain the enthalpy of solution of the zwitterionic $HB^{\pm}$ at each solvent composition. This was done by dropping solid $HB^{\pm}$ into a Dewar containing 100 g of 0.300 M NaCl. Salt was added in order to eliminate ionic strength effects when comparing enthalpies of solution in the various media. The process can be represented by

$$HB^{\pm}(c) \rightarrow HB^{\pm}(S), \qquad \Delta H_B \tag{2}$$

Since the ionization of the first proton is given by

$$H_2B^+(S) \rightarrow H^+ + HB^{\pm}(S), \qquad \Delta H_1 \tag{3}$$

It is apparent from Eqs. (1)–(3) that

$$\Delta H_1 = \Delta H_B - \Delta H_A \tag{4}$$

In order to obtain the enthalpy change for the ionization of the second proton, the following two reactions (5a and b) were studied:

$$HB^{\pm}(c) + OH^- \rightarrow B^-(S) + H_2O, \qquad \Delta H_C \tag{5}$$

Here, ΔH_C was determined, following a similar procedure to that previously outlined, by dropping $HB^{\pm}$ into 100 g of 0.300 M NaOH.

Finally, enthalpy values need to be determined for the neutralization reaction at each solvent composition:

$$H^+ + OH^- \rightarrow H_2O(S), \qquad \Delta H_N \tag{6}$$

Values for ΔH_N are cited as a function of temperature in water [8], and as a function of solvent composition in ethanol–water [9] and dioxane–water [10]. These and other values at various ionic strengths have been tabulated [11]. Measurements of reaction (6) were performed in water–methanol mixtures spanning the composition range of interest. Initially, 1.0 ml of carefully standardized 10.0 M (12.6 m) HCl was dropped into 100 g of each solvent composition at 21 °C, and the enthalpy of dilution ($\Delta_{dil}H$) obtained. Following that, the same amount of HCl was added to 100 g of solvent containing 0.105 M NaOH, yielding an observed value for the enthalpy of neutralization ($\Delta_{obs}H$). The slight excess of base was necessary in order to correct for traces of carbonate impurity. In each dilution or neutralization run, the ionic strength after mixing was maintained at 0.300 M by NaCl addition. The true enthalpy of neutralization, ΔH_N, was calculated by subtracting $\Delta_{dil}H$ from $\Delta_{obs}H$. Corrections were made for the small volume changes that accompanied HCl addition. The dilution and neutralization results are summarized in Table 1. Reported values for $\Delta_{dil}H$ of HCl at 25 °C, from molality m to infinite dilution, are 10.24, 15.40 and 0.556 kJ mol^{-1} for 10, 15 and 0.1 M solutions, respectively [12]. Interpolation shows that the expected $\Delta_{dil}H$ for HCl (12.6 m → 0.1 m) is ca. 12.9 kJ mol^{-1}, a value that is fairly comparable to the 12.0 kJ mol^{-1} obtained in this work, considering that temperature and ionic strength conditions differ somewhat. It is noteworthy that $\Delta_{dil}H$ exhibits a minimum at $X_m = 0.2$, a composition at which there is significant enhancement of the structure of water by methanol [13]. The value of 57.7 kJ mol^{-1} for ΔH_N in water at 21 °C can be compared to cited values. Some of these are, in kJ mol^{-1}: 55.8 (25 °C, $I = 0$) [14,15], 57.3 (20 °C, $I = 0$) [8], 57.0 (20 °C, $I = 0.1$ M) [16], 56.7 (25 °C, $I = 0.5$ M) [17], 54.6 (25 °C, $I = 3$ M) [18]. The observed steady decrease in ΔH_N with methanol addition is much more pronounced than that observed with ethanol [9] and dioxane [10] addition in the composition range investigated. The estimated uncertainty in $\Delta_{dil}H$ and $\Delta_{obs}H$ is 0.2, and 0.3 kJ mol^{-1} in ΔH_N.

Table 1
$\Delta_{dil}H$, $\Delta_{obs}H$ and ΔH_N in water–methanol mixtures

X_m	$-\Delta_{dil}H$ ($\pm$0.2 kJ mol^{-1})	$-\Delta_{obs}H$ ($\pm$0.2 kJ mol^{-1})	$-\Delta H_N$ ($\pm$0.3 kJ mol^{-1})
0	12.0	69.7	57.7
0.058	9.3	61.2	51.9
0.122	8.4	55.6	47.2
0.190	9.9	53.8	43.9
0.270	12.7	54.1	41.4
0.355	15.8	54.8	39.0

From Eqs. (2), (5) and (6), it is apparent that

$$\Delta H_2 = \Delta H_C - \Delta H_B - \Delta H_N \tag{7}$$

Table 2 lists values of ΔH_A, ΔH_B, and ΔH_C for BES and TES. All enthalpies progressively increase as methanol is added. This behavior indicates that solvation of buffer species H_2B^+, $HB^{\pm}$ and B^- becomes less favorable as the dielectric constant of the medium is lowered ($\varepsilon_{H_2O} = 78.3$; $\varepsilon_{methanol} = 32.6$, at 25 °C).

Table 3 lists values for ΔH_1 and ΔH_2, as calculated from Eqs. (4) and (7), respectively. For both BES and TES, ΔH_1

Table 2
ΔH_A, ΔH_B, and ΔH_C in water–methanol mixtures

X_m	ΔH_A (± 0.2 kJ mol^{-1})	ΔH_B (± 0.2 kJ mol^{-1})	ΔH_C (± 0.2 kJ mol^{-1})
BES			
0	24.9	24.3	−11.2
0.059	29.0	27.7	−1.5
0.123	29.1	29.1	5.5
0.194	30.8	30.8	13.2
0.273	32.6	31.3	16.9
0.360	32.2	31.3	19.5
TES			
0	20.7	18.6	−9.2
0.059	22.2	21.7	−0.7
0.123	23.7	23.5	5.0
0.194	25.4	24.5	13.1
0.273	25.5	24.3	15.2
0.360	23.0	23.8	17.3

is small, ranging from -2.1 to 1.0 kJ mol^{-1}. Reported values for pK_a's of $-SO_3H$ groups in compounds that bear structural similarity to BES and TES are in the range 1.5–2.0. For example, pK_{a1} at 25 °C is 1.5 for 2-aminoethanesulfonic acid and 1.3 for DL-cysteic acid [19]. Thus, in view of the expected pK_{a1} range for BES and TES and the large excess of HCl used (0.300 M versus ca. 0.02 M for BES and TES), reaction (5) may be considered complete. By contrast, the high positive ΔH_2 values in water for BES and TES, 22.2 and 30.0 kJ mol^{-1}, fall in the range that would be expected for weakly acidic protonated amines. The lower values for BES in the solvent range studied are in agreement with observations and compilations [20,21] indicating that ΔH for the ionization of the *N*-bound proton in protonated amines rises steadily with increasing p*K*. At 25 °C and $I = 0.1$ M, pK_{a2}'s are 7.06 and 7.5 for BES and TES, respectively [22]. For BES, reported ΔH_2 values in water at 25 °C are 22.8 ($I = 0.01$ M) [23], 23.9 ($I = 0$) [24] and 25.2 ($I = 0.1$) kJ mol^{-1} [22]. For TES, they are 29.2 ($I = 0.01$ M) [23], 32.1 ($I = 0$) [25,26], and 32.7 ($I = 0$) [22] kJ mol^{-1}. Table 3 shows that ΔH_2 for BES rises with methanol addition then tends to level off at the higher X_m. For TES, however, a minimum in an otherwise slower rise in ΔH_2 is observed at $X_m = 0.123$. The rise in ΔH_2 may be explained on the basis that methanol addition weakens the "ice-like" organized structure of water [27], thereby rendering proton solvation less favorable, since such a process would necessarily enter a competition with methanol for H-bonding with water. The minimum with TES may be explained on the basis of increased basicity of the medium as ethanol is initially added [28]. With TES, the presence of an additional –OH group and its comparatively lower acid strength could have contributed to this minimum. However, interpretation of ΔH behavior is much more complex than that of the free energy, ΔG. For the latter, solvent dependence may be explained on the basis of electrostatic and non-electrostatic terms. Enthalpy changes, on the other hand, in addition to electrostatic contributions [29] involve structural and steric factors that are often difficult to assess [27]. However, accumulation of significant proton ionization enthalpies in mixed solvents on structurally similar molecules should aid in understanding the role of these factors.

Table 3
ΔH_1 and ΔH_2 for BES and TES in water–methanol mixtures

X_m	ΔH_1(± 0.2 kJ mol^{-1})	ΔH_2 (± 0.2 kJ mol^{-1})
BES		
0	−0.7	22.2
0.059	−1.3	22.8
0.123	0.0	24.6
0.194	0.0	26.4
0.273	−1.4	27.0
0.360	−1.0	27.2
TES		
0	−2.1	30.0
0.059	−0.5	29.5
0.123	−0.2	28.6
0.194	−1.0	32.5
0.273	−1.4	32.3
0.360	1.0	32.4

Acknowledgements

The authors wish to thank the American University of Sharjah for supporting this work.

References

[1] R.G. Bates, Determination of pH, second ed., Wiley, New York, 1973 (Chapter 7).

[2] R.G. Bates, R.N. Roy, R.A. Robinson, J. Solut. Chem. 3 (1974) 905.

[3] R.N. Roy, J.J. Gibbons, R. Snelling, J. Solut. Chem. 6 (1977) 475.

[4] R.C. Das, U.N. Dash, K.N. Panda, J. Chem. Soc., Faraday I 76 (1980) 2152.

[5] F.H. Jumean, Z. Abdelrahim, Ann. Chim. (Rome) 82 (1992) 49.

[6] F.H. Jumean, Z. Abdelrahim, Arab. J. Sci. Eng. 19 (1994) 77.

[7] R.W. Ramette, J. Chem. Ed. 61 (1984) 76.

[8] H.S. Harned, W.J. Hamer, J. Am. Chem. Soc. 55 (1933) 2194.

[9] G.L. Bertrand, F.J. Millero, C.H. Wu, L.G. Hepler, J. Phys. Chem. 70 (1966) 699.

[10] H. Kido, W.C. Fernelius, J. Phys. Chem. 65 (1961) 574.

[11] J.J. Christensen, R.M. Izatt, L.D. Hansen, in: G.D. Fasman (Ed.), Handbook of Biochemistry and Molecular Biology, third ed., vol. I, CRC Press, Boca Raton, FL, 1976, pp. 260–261.

[12] V.B. Parker, in: Thermal Properties of Aqueous Uni-Univalent Electrolytes, Natl. Stand. Ref. Data Ser. -Natl. Bur. Stand. (US) 2, US Government Printing Office, Washington, DC, 1965.

[13] F. Franks, in: F. Franks (Ed.), Physiochemical Processes in Mixed Aqueous Solvents, Elsevier, New York, 1967, pp. 50–79.

[14] J.D. Hales, R.M. Izatt, J.J. Christensen, J. Phys. Chem. 67 (1963) 2605.

[15] C.E. Vanderzee, J.A. Swanson, J. Phys. Chem. 67 (1963) 2608.

[16] P.J. Antikainen, Suomen Kemistilehti B29 (1956) 14.

[17] G. Olofsson, L. Hepler, J. Solut. Chem. 4 (1975) 127.

[18] R. Arnek, C.C. Patel, Acta Chem. Scand. 22 (1968) 1097.
[19] Handbook of Chemistry and Physics, 83rd ed. CRC Press, Boca Raton, FL, pp. 2002–2003 (Section 8).
[20] P. Lumme, P. Lahermo, T. Tummavuori, Acta Chem. Scand. 19 (1965) 2175.
[21] R.N. Goldberg, N. Kishore, R.M. Lennen, J. Phys. Chem. Ref. Data 31 (2002) 231.
[22] H. Fukada, K. Takahashi, Proteins: Struct. Funct. Genet. 33 (1998) 159.
[23] C.D. McGlothlin, J. Jordan, J. Anal. Lett. 9 (1976) 245.
[24] T. Roig, P. Backman, G. Olofsson, Acta Chem. Scand. 47 (1993) 899.
[25] C.A. Vega, R.G. Bates, Anal. Chem. 48 (1976) 1293.
[26] R.N. Roy, C.P. Moore, P. Lord, D. Murad, L.N. Roy, W.S. Good, J. Niederschmidt, K.M. Kuhler, J. Chem. Thermodyn. 29 (1997) 1323.
[27] D. Feakins, pp. 71–90 (in ref. 3).
[28] R.G. Bates, R.A. Robinson, in: B.E. Conway, R.G. Barradas (Eds.), Chemical Physics of Ionic Solutions, Wiley, New York, 1966, pp. 211–235.
[29] L.D. Hansen, L.G. Hepler, Can. J. Chem. 50 (1972) 1030.