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THERMODYNAMICS OF THE MIXING PROCESS OF SEVERAL SODIUM SULFONAMIDES IN ETHANOL + WATER COSOLVENT MIXTURES

ESTUDIO TERMODINÁMICO DEL PROCESO DE MEZCLA DE ALGUNAS SULFONAMIDAS SÓDICAS EN MEZCLAS COSOLVENTES ETANOL + AGUA

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ABSTRACT

Sodium sulfonamides have been extensively used for the treatment of certain infections caused by several types of microorganisms. Although sulfonamides are still widely used in therapeutics, the physicochemical information about their aqueous solutions has not been completed. In this context, the thermodynamic functions of mixing three structurally related sodium sulfonamides were evaluated: Gibbs energy, enthalpy, and entropy. The quantities of mixing were calculated based on the fusion calorimetric values obtained from differential scanning calorimetry measurements and equilibrium solubility values reported in the literature for all the drugs with ethanol + water mixtures. By means of an enthalpy-entropy compensation analysis, non-linear ΔH^0_{mix} vs. ΔG^0_{mix} plots with negative slopes from neat ethanol to a 0.60 ethanol mass fraction, and positive slopes from the latter composition to neat water were obtained. From these results, it was concluded that the dissolution process of these drugs in ethanol-rich mixtures was entropy-driven; whereas, in water-rich mixtures the process was enthalpy-driven. Nevertheless, the molecular and ionic events involved in the dissolution of these drugs in this cosolvent system remain unclear.

Keywords: Sodium sulfonamides, mixing process, cosolvency, ethanol, solution thermodynamics.

RESUMEN

Las sulfonamidas sódicas han sido ampliamente utilizadas en el tratamiento de ciertas infecciones causadas por diferentes microorganismos. Si bien las sulfonamidas siguen siendo ampliamente usadas en la terapéutica actual, la información fisicoquímica de sus soluciones acuosas aún no es completa. En este contexto, estudiamos aquí las funciones termodinámicas de mezcla de tres sulfonamidas sódicas relacionadas estructuralmente y que fueron calculadas a partir de las propiedades calorimétricas de fusión y de los valores de solubilidad en equilibrio en mezclas etanol + agua publicados en la literatura. Mediante análisis de compensación entálpica-entrópica se obtuvieron gráficos no lineales de ΔH^0_{mix} vs. ΔG^0_{mix} exhibiendo pendientes negativas desde el etanol puro hasta la mezcla cosolvente de 0,60 en fracción másica de etanol y pendientes positivas desde esta mezcla hasta el agua pura; de acuerdo a este resultado se tiene que el proceso de disolución de estos fármacos en mezclas ricas en etanol es conducido

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entrópicamente, mientras que en mezclas ricas en agua el proceso es conducido entálpicamente. Sin embargo, los eventos moleculares e iónicos involucrados en el proceso de disolución de este fármaco en este sistema cosolventes aún no son claros.

Palabras clave: sulfonamidas sódicas, proceso de mezcla, cosolvencia, etanol, termodinámica de soluciones.

INTRODUCTION

Sodium sulfonamides (Sulfadiazine: Na-SD, Sulfamerazine: Na-SMR, and Sulfamethazine: Na-SMT, which molecular structures are presented in figure 1) are drugs extensively used for the treatment of certain infections caused by several types of microorganisms (1). Although these drugs are widely used nowadays in therapeutics, the physicochemical information about their aqueous solutions is not complete at present, even though several physicochemical studies have been done. Thus, the solution thermodynamics in aqueous media for these drugs (as dissociate and non-dissociate compounds) has been presented in the literature (2, 3). Moreover, the transfer physicochemical aspects of these drugs (as non-dissociate compound) from aqueous media up to octanol and some phospholipidic vesicles have also been reported (4). In addition, the apparent molar volumes in water and ethanol have also been studied as a function of drug concentration at 298.15 K (2, 5). Finally, the solubility in ethanol + water cosolvent mixtures has been studied as a function of cosolvent composition and temperature, and the results have been correlated by means of the Jouyban-Acree solubility model (6). Similarly, the thermodynamic quantities of the Na-SD solution were reported in the literature (7).

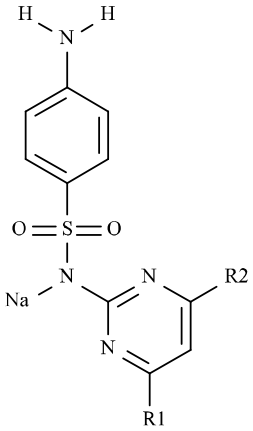

Figure 1. Molecular structure of sodium sulfonamides. Na-SD: R1 = H, R2 = H; Na-SMR: R1 = CH_3, R2 = H; Na-SMT: R1 = CH_3, R2 = CH_3.

As it has been already described, the solubility behavior of drugs in cosolvent mixtures is very important because cosolvent blends are frequently used in purification methods, preformulation studies, and pharmaceutical dosage forms designs, among other applications (8). For these reasons, it is important to systematically determine and collect data about the equilibrium solubility of pharmaceutical compounds. It is also particularly important because it is not possible yet to predict the solubility of drugs in water, organic solvents and/or mixed solvent systems with an acceptable margin of error in the prediction (9). Moreover, temperature-solubility dependence allows us to carry out the respective thermodynamic analysis, which, on the other hand, also permits inside the molecular mechanisms, involved toward the solution processes (10).

In this context, the main objective of this research is to evaluate the effect of the cosolvent composition on the thermodynamics of mixing of Na-SD, Na-SMR and Na-SMT in some ethanol (EtOH) + water cosolvent mixtures. This study is based on both the calorimetric properties of fusion obtained by differential scanning calorimetry (DSC), and the van't Hoff treatment of equilibrium solubility values reported in the literature (6, 7).

MATERIALS AND METHODS

Reagents

All the sodium sulfonamides (Na-SD: 4-Amino-*N*-2-pyrimidinylbenzenesulfonamide sodium salt, CAS: [547-32-0]; Na-SMR: 4-Amino-*N*-(4-methyl-2-pyrimidinyl)benzenesulfonamide sodium salt, CAS: [127-58-2]; and Na-SMT: 4-Amino-*N*-(4,6-dimethyl-2-pyrimidinyl)benzenesulfonamide sodium salt, CAS: [1981-58-4] (11)) used were in agreement with the quality requirements indicated in the American Pharmacopeia, USP (12), as well as in the British Pharmacopoeia, BP (13).

Calorimetric study

Melting point and enthalpy of fusion of three sodium sulfonamides were determined through

DSC studies (DSC 823E Mettler Toledo). Thermal analyses were performed at a heating rate of 10 K min^{-1} in a dynamic nitrogen atmosphere (60 mL min^{-1}). Approximately 1.5 mg of drugs were used. The equipment was calibrated using Indium as a standard (14).

RESULTS AND DISCUSSION

Before presenting the results, it is important to consider that these drugs (in similar way to procaine hydrochloride and sodium naproxen) have an electrolyte behavior (15, 16). Therefore, such drugs dissociate in aqueous solution, interacting with the cosolvent mixture by strong ion-dipole interactions, as well as by other weak non covalent interactions due to their nonpolar groups. Thus, they also could act as a Lewis acid (–NH_2 group) or as a Lewis base (–NH_2 and –SO_2– groups), in order to establish hydrogen bonds with proton-acceptor or donor functional groups in the solvents (–OH groups) (17, 18).

Ideal solubility of sodium sulfonamides

The ideal solubility of non-electrolyte crystalline solutes in a liquid solvent can be calculated by means of equation 1:

$$\ln x_2^{id} = -\frac{\Delta H_{fus}(T_{fus}-T)}{RT_{fus}T} + \left(\frac{\Delta C_p}{R}\right)\left[\frac{(T_{fus}-T)}{T} + \ln\left(\frac{T}{T_{fus}}\right)\right] \qquad \textbf{Equation 1.}$$

where x_2^{id} is the ideal solute solubility as mole fraction; ΔH_{fus} is the molar enthalpy of fusion of the pure solute (at the melting point); T_{fus} is the absolute melting point; T is the absolute solution temperature; R is the gas constant (8.314 J mol^{-1} K^{-1}); and ΔC_p is the difference between the molar heat capacity of the crystalline form and the molar heat capacity of the hypothetical supercooled liquid form, both at the solution temperature (19). Since ΔC_p cannot be easily experimentally determined, it is usual to assume that it may be approximated to the entropy of fusion, which is calculated as follows: $\Delta S_{fus} = \Delta H_{fus}/T_{fus}$. The main reasons for the last assumption have been well discussed in the literature (20). Although equation 1 was developed for non electrolyte compounds, it has also been used to estimate ideal solubilities of some electrolyte drugs (21, 22).

In this context, tables 1 and 2 summarize the thermodynamic properties of fusion and the ideal solubilities of the three sodium sulfonamides, respectively. At the same temperature, x_2^{id} diminishes in the following order: Na–SMR > Na–SD > Na–SMR; whereas for the respective molecular drugs it diminishes in the following order: SMT > SMR > SD (3). In the case of sulfamethazine, the ideal solubilities are similar for the sodium salt and the molecular form (1.068 x 10^{-2} and 1.051 x 10^{-2} for Na–SMT, and molecular SMT at 298.15 K, respectively). Now, in the case of sulfadiazine, the ideal solubilities are also similar but slightly greater for sodium salt (4.03 x 10^{-3} and 3.03 x 10^{-3} for Na–SD and molecular SD at 298.15 K, respectively). Finally, in the case of sulfamerazine, the ideal solubilities are greater for the molecular form in comparison to sodium salt (2.458 x 10^{-3} and 5.46 x 10^{-3} for Na–SMR and SMR at 298.15 K, respectively).

Moreover, the experimental equilibrium solubility values for these drugs have been reported in the literature (6).

Table 1. Thermodynamic properties of sodium sulfonamides fusion[a].

Drug	T_{fus} / K	ΔH_{fus} / kJ mol^{-1}	ΔS_{fus} / J mol^{-1} K^{-1}
Na-SD	652.0	38.2 (0.4)	58.6 (0.6)
Na-SMR	644.4	41.8 (0.4)	64.8 (0.6)
Na-SMT	582.8	32.8 (0.3)	56.3 (0.6)

[a] Values in parentheses are standard deviations.

Table 2. Ideal solubility of sodium sulfonamides at several temperatures[a].

T / K	Na-SD	Na-SMR	Na-SMT
278.15	2.469 (0.025) x 10^{-3}	1.431 (0.014) x 10^{-3}	6.67 (0.07) x 10^{-3}
283.15	2.800 (0.028) x 10^{-3}	1.644 (0.016) x 10^{-3}	7.53 (0.08) x 10^{-3}
288.15	3.17 (0.03) x 10^{-3}	1.884 (0.019) x 10^{-3}	8.47 (0.08) x 10^{-3}
293.15	3.58 (0.04) x 10^{-3}	2.154 (0.022) x 10^{-3}	9.52 (0.10) x 10^{-3}
298.15	4.03 (0.04) x 10^{-3}	2.458 (0.025) x 10^{-3}	1.068 (0.011) x 10^{-2}
303.15	4.53 (0.05) x 10^{-3}	2.798 (0.028) x 10^{-3}	1.195 (0.012) x 10^{-2}
308.15	5.08 (0.05) x 10^{-3}	3.18 (0.03) x 10^{-3}	1.335 (0.013) x 10^{-2}

[a] Values in parentheses are standard deviations.

Thermodynamic quantities of sodium sulfonamide solution

Due to the fact that the drugs considered in this study are electrolyte compounds, it is important to keep in mind that in general terms, it could be stated that a strong electrolyte dissociates according to the following expression: $C_{\nu+}A_{\nu-} \rightarrow \nu_+ C^{z+} + \nu_- A^{z-}$, where ν_+ is the number of cations (C^{z+}) of valence

$z+$, and ν_- is the number of anions (A^{z-}) of valence $z-$., The concept of mean ionic activity ($a^{\nu}_{\pm}$) is used because is not possible to determine experimentally the activity of ions separately. Thus, the thermodynamic activity for an electrolyte can be defined as follows: $a_2 = a_+^{\nu+} a_-^{\nu-} = a_{\pm}^{\nu}$ (23-25).

The three sodium sulfonamides are electrolyte solutes of type one-one, which means that they dissociate in aqueous solutions to generate two species: a monovalent anion and a monovalent cation, respectively. If the inter-ionic interactions are not considered, the ν value for these drugs could be ideally assumed to be 2, in a first approach, thus this value could be used to calculate the apparent thermodynamic solution functions (7, 15, 16).

Therefore, according to van't Hoff's analysis, the apparent standard enthalpy change of solution (ΔH^0_{soln}) for electrolytes type one-one (such as the sodium sulfonamides studied here), is obtained using the mean harmonic temperature (T_{hm} is 292.8 K in our case) according to equation 2 (7, 15, 16), that is if the inter-ionic interactions are not considered.

$$\left(\frac{\partial \ln x_{\text{Na-Sulf}}}{\partial(1/T - 1/T_{\text{hm}})}\right)_P = -\frac{\Delta H^0_{\text{soln}}}{2 \cdot R} \qquad \textbf{Equation 2.}$$

where, R is the universal gas constant (8.314 J mol^{-1} K^{-1}). In all the studied cases, linear models with good determination coefficients were obtained.

The apparent standard Gibbs energy change for the solution process (ΔG^0_{soln}) of electrolytes type one-one was calculated at the mean harmonic temperature by means of equation 3, considering the approach proposed by Krug *et al.,* 1976 (26).

$$\Delta G^0_{\text{soln}} = -2 \cdot R \cdot T_{\text{hm}} \cdot \text{intercept} \qquad \textbf{Equation 3.}$$

in which, the intercept used was the one obtained in the analysis taking $\ln x_{\text{Na-Sulf}}$ as a function of $1/T - 1/T_{\text{hm}}$. Finally, the apparent standard entropic change for the solution process (ΔS^0_{soln}) was obtained from the respective ΔH^0_{soln} and ΔG^0_{soln} values by means of the following equation:

$$\Delta S^0_{\text{soln}} = \frac{\left(\Delta H^0_{\text{soln}} - \Delta G^0_{\text{soln}}\right)}{T_{\text{hm}}} \qquad \textbf{Equation 4.}$$

Table 3 summarizes the apparent standard thermodynamic functions for the experimental solution process of sodium sulfonamides in all EtOH + water cosolvent mixtures. In order to calculate the thermodynamic quantities for the experimental solution processes, some uncertainty propagation methods were used (27). As it was expected, it was found that the standard Gibbs energy of the solution is positive in all cases because the mole fraction is always lower than the unit; thus its logarithmic term is negative, therefore, the standard Gibbs energy will be a positive quantity.

The apparent enthalpy of solution is positive in all cases, therefore, the process is always endothermic. Similarly, the entropy of solution is also positive for Na-SMT in all solvent systems and Na-SMR in compositions from neat water to 0.80 in a mass fraction of EtOH. Oppositely, this quantity is negative for Na-SD in all solvent systems and Na-SMR in the mixture of 0.90 in mass fraction of EtOH and neat EtOH. These results do not show enthalpy or entropy-driving overall in the solution process for the last indicated systems. For all sodium sulfonamides, ΔH^0_{soln} values increase from neat water to 0.60 in mass fraction of EtOH, and decrease from this EtOH proportion to neat EtOH. In contrast with enthalpy, ΔS^0_{soln} values increase from neat water to the mixtures of 0.30, 0.40 and 0.50 in mass fraction of EtOH for Na-SD, Na-SMR and Na-SMT respectively, despite the sign obtained and diminishing beyond this composition to neat EtOH. The apparent enthalpic and entropic values obtained for the dissolution process of Na-SD in neat water are in good agreement regarding those reported in the literature (11.0 ± 0.6 kJ mol^{-1} and –13.5 ± 0.7 J mol^{-1} K^{-1}, respectively) (2). However, the same quantities for Na-SMT are not in agreement with those reported in the same research (10.9 ± 0.6 kJ mol^{-1} and –19.7 ± 1.0 J mol^{-1} K^{-1}, respectively) (2). That is why, it is important to note that the studied temperature intervals and the used analytical techniques were significantly different in both investigations.

Aiming to compare the relative contributions by enthalpy (ζ_H) and by entropy (ζ_{TS}) toward the solution process, equations 5 and 6 were respectively employed (28).

Table 3. Apparent thermodynamic quantities related to the sodium sulfonamide solution process in ethanol + water cosolvent mixtures at 292.8 K[a].

μ_{EtOH} [b]	ΔG^0_{soln} / kJ mol^{-1}	ΔH^0_{soln} /kJ mol^{-1}	ΔG^0_{soln} /J mol^{-1} K^{-1}	$T\Delta S^0_{soln}$ /kJ mol^{-1}	ζ_H [c]	ζ_{TS} [c]
Na-SD [d]						
0.00	15.00 (0.02)	11.02 (0.02)	–13.59 (0.03)	–3.98 (0.01)	0.735	0.265
0.10	16.00 (0.02)	13.03 (0.03)	–10.15 (0.03)	–2.97 (0.01)	0.814	0.186
0.20	17.17 (0.04)	16.64 (0.05)	–1.80 (0.01)	–0.53 (0.00)	0.969	0.031
0.30	18.73 (0.04)	18.43 (0.05)	–1.04 (0.00)	–0.31 (0.00)	0.984	0.016
0.40	20.23 (0.07)	19.91 (0.08)	–1.11 (0.01)	–0.32 (0.00)	0.984	0.016
0.50	22.28 (0.03)	20.56 (0.07)	–5.87 (0.02)	–1.72 (0.01)	0.923	0.077
0.60	24.55 (0.06)	20.57 (0.06)	–13.60 (0.05)	–3.98 (0.02)	0.838	0.162
0.70	27.12 (0.03)	19.13 (0.04)	–27.30 (0.06)	–7.99 (0.02)	0.705	0.295
0.80	31.36 (0.10)	14.75 (0.06)	–56.73 (0.29)	–16.61 (0.09)	0.470	0.530
0.90	36.99 (0.19)	12.65 (0.10)	–83.1 (0.8)	–24.34 (0.22)	0.342	0.658
1.00	45.71 (0.05)	9.95 (0.02)	–122.1 (0.3)	–35.76 (0.10)	0.218	0.782
Ideal	14.01 (0.14)	17.15 (0.13)	10.36 (0.13)	3.14 (0.04)	0.845	0.155
Na-SMR						
0.00	20.03 (0.10)	32.17 (0.13)	41.45 (0.27)	12.14 (0.08)	0.726	0.274
0.10	20.70 (0.08)	35.27 (0.07)	49.75 (0.22)	14.57 (0.07)	0.708	0.292
0.20	21.62 (0.08)	40.02 (0.22)	62.8 (0.4)	18.39 (0.12)	0.685	0.315
0.30	22.57 (0.05)	43.86 (0.06)	72.71 (0.18)	21.29 (0.05)	0.673	0.327
0.40	23.47 (0.10)	45.51 (0.07)	75.3 (0.3)	22.04 (0.10)	0.674	0.326
0.50	24.65 (0.10)	46.11 (0.09)	73.3 (0.3)	21.46 (0.10)	0.682	0.318
0.60	26.52 (0.17)	46.65 (0.17)	68.7 (0.5)	20.12 (0.15)	0.699	0.301
0.70	28.99 (0.19)	44.61 (0.13)	53.3 (0.4)	15.62 (0.11)	0.741	0.259
0.80	32.65 (0.23)	40.61 (0.26)	27.20 (0.26)	7.96 (0.08)	0.836	0.164
0.90	37.90 (0.24)	32.09 (0.13)	–19.84 (0.15)	–5.81 (0.04)	0.847	0.153
1.00	42.95 (0.17)	23.52 (0.14)	–66.3 (0.5)	–19.43 (0.14)	0.548	0.452
Ideal	15.27 (0.15)	18.97 (0.14)	12.21 (0.15)	3.70 (0.05)	0.837	0.163
Na-SMT						
0.00	17.57 (0.12)	28.93 (0.15)	38.8 (0.3)	11.36 (0.10)	0.718	0.282
0.10	17.89 (0.13)	36.31 (0.14)	62.9 (0.5)	18.43 (0.15)	0.663	0.337
0.20	18.20 (0.08)	42.73 (0.07)	83.8 (0.4)	24.53 (0.12)	0.635	0.365
0.30	18.58 (0.07)	51.38 (0.10)	112.0 (0.5)	32.80 (0.15)	0.610	0.390
0.40	19.15 (0.09)	56.10 (0.09)	126.2 (0.6)	36.95 (0.19)	0.603	0.397
0.50	19.87 (0.12)	59.03 (0.12)	133.7 (0.8)	39.16 (0.25)	0.601	0.399
0.60	21.08 (0.12)	59.96 (0.17)	132.8 (0.8)	38.88 (0.24)	0.607	0.393
0.70	22.85 (0.12)	54.14 (0.11)	106.9 (0.6)	31.29 (0.18)	0.634	0.366
0.80	25.85 (0.12)	42.22 (0.11)	55.9 (0.3)	16.37 (0.09)	0.721	0.279
0.90	29.86 (0.07)	38.21 (0.09)	28.51 (0.10)	8.35 (0.03)	0.821	0.179
1.00	31.90 (0.10)	34.0 (0.3)	7.09 (0.07)	2.08 (0.02)	0.942	0.058
Ideal	11.57 (0.12)	16.48 (0.13)	16.18 (0.21)	4.91 (0.06)	0.771	0.229

[a] Values in parentheses are standard deviations. [b] μ_{EtOH} is the mass fraction of ethanol in the solute-free cosolvent mixture. [c] ζ_H and ζ_{TS} are the relative contributions by enthalpy and entropy toward Gibbs energy of solution. [d] Taken from Delgado and Martinez, 2010 (7).

$$\zeta_H = \frac{\left|\Delta H^0_{soln}\right|}{\left|\Delta H^0_{soln}\right| + \left|T\Delta S^0_{soln}\right|} \qquad \textbf{Equation 5.}$$

$$\zeta_{TS} = \frac{\left|T\Delta S^0_{soln}\right|}{\left|\Delta H^0_{soln}\right| + \left|T\Delta S^0_{soln}\right|} \qquad \textbf{Equation 6.}$$

From the ζ_H and ζ_{TS} values shown in table 3, it was concluded that the main contributor of the standard Gibbs energy of the solution process of SD-Na is enthalpy, except for Na-SD in the following cosolvent composition interval: $0.80 \leq \mu_{EtOH} \leq 1.00$. This result evidences the relevance of the energetic factor on the dissolution processes of these drugs.

Thermodynamic quantities of the sodium sulfonamide mixture

The solution process can be represented by the following hypothetical stages (10):

$$\text{Solute}_{(\text{Solid})} \rightarrow \text{Solute}_{(\text{Liquid})} \rightarrow \text{Solute}_{(\text{Solution})}$$

where, the respective partial processes regarding the solution are solute fusion and mixing at the same temperature (292.8 K), which allows calculating the partial thermodynamic contributions to the overall solution process by means of equations 7 and 8, respectively.

$$\Delta H^0_{\text{soln}} = \Delta H^{292.8}_{\text{fus}} + \Delta H^0_{\text{mix}} \qquad \textbf{Equation 7.}$$

$$\Delta S^0_{\text{soln}} = \Delta S^{292.8}_{\text{fus}} + \Delta S^0_{\text{mix}} \qquad \textbf{Equation 8.}$$

where, $\Delta H^{292.8}_{\text{fus}}$ and $\Delta S^{292.8}_{\text{fus}}$ represent the thermodynamic functions of the fusion process at the harmonic temperature (292.8 K). $\Delta H^{292.8}_{\text{fus}}$ was calculated by means of the following equation: $\Delta H^{T}_{\text{fus}} = \Delta H^{\text{MP}}_{\text{fus}} - \Delta C_{\text{p}}(T_{\text{fus}} - T)$, using $\Delta S^{\text{MP}}_{\text{fus}}$ instead of ΔC_{p}, for obtaining the values of 17.16, 18.98 and 16.49 kJ mol^{-1} for Na-SD, Na-SMR, and Na-SMT, respectively. These values coincide with the enthalpic changes for the ideal solution processes, as it is shown in table 3. In contrast, the entropies of fusion at 292.8 K (58.6, 64.8, and 56.3 J mol^{-1} K^{-1} for Na-SD, Na-SMR, and Na-SMT, respectively) do not coincide with the entropic changes of the ideal solution processes at this temperature, as it is also shown in table 3. For this reason and for practical purposes, $\Delta S^{0\text{id}}_{\text{soln}}$ values were used instead of $\Delta S^{292.8}_{\text{fus}}$, as it was previously done with several non electrolyte drugs in EtOH + water cosolvent mixtures (10, 29 - 31). Table 4 summarizes the thermodynamic quantities of the sodium sulfonamide mixing processes. The Gibbs energy of mixing was positive in all the systems studied. It is interesting to note that the Gibbs energies of transfer from water to EtOH (calculated as the difference between Gibbs energies of mixing in EtOH and water) diminish in the following order: Na-SD (30.7 kJ mol^{-1}) > Na-SMR (22.9 kJ mol^{-1}) > Na-SMT (14.3 kJ mol^{-1}).

When the partial contributions of the ideal solution (related to the solute fusion process) and mixing processes to the enthalpy and entropy of solution were analyzed, it was found that $\Delta H^{0\text{id}}_{\text{soln}}$ and $\Delta S^{0\text{id}}_{\text{soln}}$ are positive (as it is shown in table 4), while the contribution of the thermodynamic quantities related to the mixing process toward the solution processes is variable. Thus, ΔH^0_{mix} is positive for all sodium sulfonamides except for Na-SD in the following cosolvent composition interval: 0.30 $\leq \mu_{\text{EtOH}} \leq$ 0.70. However, the entropy of mixing (ΔS^0_{mix}) is negative for Na-SD in all mixtures, whereas for Na-SMR and Na-SMT, this quantity is positive from neat water to 0.80 or 0.90 in mass fraction of EtOH, respectively. If the contribution of the mixing processes to the dissolution processes is considered, it can be concluded that, i) the mixing process is driven by enthalpy ($\Delta H^0_{\text{mix}} < 0$ and $\Delta S^0_{\text{mix}} < 0$) for Na-SD in the intervals 0.00 $\leq \mu_{\text{EtOH}} \leq$ 0.20 and 0.80 $\leq \mu_{\text{EtOH}} \leq$ 1.00, ii) the mixing process is driven by entropy ($\Delta H^0_{\text{mix}} > 0$ and $\Delta S^0_{\text{mix}} > 0$) for Na-SMR in the interval 0.00 $\leq \mu_{\text{EtOH}} \leq$ 0.80; and iii) the mixing process is also driven by entropy ($\Delta H^0_{\text{mix}} > 0$ and $\Delta S^0_{\text{mix}} > 0$) for Na-SMT in the interval 0.00 $\leq \mu_{\text{EtOH}} \leq$ 0.90. It was found that all other cases were not enthalpy or entropy-driven ($\Delta H^0_{\text{mix}} > 0$ and $\Delta S^0_{\text{mix}} < 0$).

The net variation in ΔH^0_{mix} values results from the contribution of several types of interaction. The enthalpy of the cavity formation (required for solute accommodation) is endothermic because the energy must be supplied against the cohesive forces of the solvent. This process decreases solubility. On the other hand, the enthalpy of solute-solvent interaction is exothermic and results mainly from ion-dipole, van der Waals and Lewis acid-base interactions. In the case of non-electrolyte drugs, the structuring of water molecules around the nonpolar groups of solutes (hydrophobic hydration) contributes to reduce the net heat of the mixing process to minor or even negative values in aqueous solutions, as it is the case of pure water (as it is shown in table 4 for Na-SD). Nevertheless, these drugs are electrolyte and, therefore, they interact with the solvent through ion-dipole interactions, which could lead to hydrophilic hydration around the anionic group.

Table 4. Apparent thermodynamic quantities related to the sodium sulfonamide mixing process in ethanol + water cosolvent mixtures at 292.8 K[a].

μ_{EtOH} [b]	ΔG^0_{mix} / kJ mol^{-1}	ΔH^0_{mix} / kJ mol^{-1}	ΔS^0_{mix} / J mol^{-1} K^{-1}	$T\Delta S^0_{mix}$ / kJ mol^{-1}	ζ_H [c]	ζ_{TS} [c]
Na-SD						
0.00	0.99 (0.14)	–6.13 (0.14)	–23.95 (0.13)	–7.12 (0.04)	0.463	0.537
0.10	2.00 (0.14)	–4.11 (0.14)	–20.51 (0.13)	–6.11 (0.04)	0.402	0.598
0.20	3.16 (0.15)	–0.50 (0.14)	–12.16 (0.13)	–3.67 (0.04)	0.121	0.879
0.30	4.73 (0.15)	1.28 (0.14)	–11.40 (0.13)	–3.44 (0.04)	0.271	0.729
0.40	6.23 (0.16)	2.77 (0.16)	–11.46 (0.13)	–3.46 (0.04)	0.444	0.556
0.50	8.27 (0.14)	3.41 (0.15)	–16.22 (0.13)	–4.86 (0.04)	0.413	0.587
0.60	10.54 (0.15)	3.42 (0.15)	–23.96 (0.14)	–7.12 (0.04)	0.324	0.676
0.70	13.12 (0.14)	1.98 (0.14)	–37.66 (0.15)	–11.13 (0.04)	0.151	0.849
0.80	17.35 (0.17)	–2.40 (0.15)	–67.1 (0.3)	–19.75 (0.09)	0.108	0.892
0.90	22.99 (0.23)	–4.49 (0.17)	–93.5 (0.8)	–27.48 (0.22)	0.141	0.859
1.00	31.71 (0.15)	–7.19 (0.14)	–132.5 (0.4)	–38.90 (0.11)	0.156	0.844
Na-SMR						
0.00	4.76 (0.18)	13.20 (0.19)	29.2 (0.3)	8.44 (0.09)	0.610	0.390
0.10	5.44 (0.17)	16.30 (0.16)	37.54 (0.27)	10.87 (0.08)	0.600	0.400
0.20	6.36 (0.17)	21.05 (0.26)	50.6 (0.4)	14.69 (0.13)	0.589	0.411
0.30	7.30 (0.16)	24.89 (0.15)	60.51 (0.24)	17.59 (0.07)	0.586	0.414
0.40	8.21 (0.18)	26.55 (0.15)	63.1 (0.4)	18.34 (0.11)	0.591	0.409
0.50	9.38 (0.18)	27.15 (0.16)	61.1 (0.4)	17.76 (0.11)	0.604	0.396
0.60	11.26 (0.23)	27.68 (0.22)	56.5 (0.5)	16.42 (0.15)	0.628	0.372
0.70	13.73 (0.24)	25.64 (0.19)	41.1 (0.4)	11.92 (0.12)	0.683	0.317
0.80	17.38 (0.28)	21.6 (0.3)	15.0 (0.3)	4.26 (0.09)	0.835	0.165
0.90	22.63 (0.28)	13.12 (0.19)	–32.05 (0.21)	–9.51 (0.06)	0.580	0.420
1.00	27.68 (0.23)	4.56 (0.19)	–78.5 (0.5)	–23.13 (0.14)	0.165	0.835
Na-SMT						
0.00	6.00 (0.17)	12.45 (0.20)	22.6 (0.4)	6.45 (0.12)	0.659	0.341
0.10	6.31 (0.17)	19.84 (0.20)	46.8 (0.5)	13.52 (0.16)	0.595	0.405
0.20	6.63 (0.14)	26.25 (0.15)	67.6 (0.5)	19.62 (0.13)	0.572	0.428
0.30	7.01 (0.14)	34.91 (0.16)	95.8 (0.5)	27.89 (0.16)	0.556	0.444
0.40	7.58 (0.15)	39.62 (0.16)	110.0 (0.7)	32.04 (0.20)	0.553	0.447
0.50	8.30 (0.17)	42.55 (0.18)	117.6 (0.9)	34.26 (0.25)	0.554	0.446
0.60	9.51 (0.16)	43.49 (0.22)	116.6 (0.8)	33.98 (0.25)	0.561	0.439
0.70	11.28 (0.17)	37.66 (0.17)	90.7 (0.6)	26.39 (0.19)	0.588	0.412
0.80	14.28 (0.17)	25.74 (0.17)	39.7 (0.4)	11.46 (0.11)	0.692	0.308
0.90	18.29 (0.14)	21.73 (0.16)	12.33 (0.23)	3.44 (0.07)	0.863	0.137
1.00	20.33 (0.15)	17.5 (0.4)	–9.09 (0.22)	–2.83 (0.07)	0.861	0.139

[a] Values in parentheses are standard deviations. [b] μ_{EtOH} is the mass fraction of ethanol in the solute-free cosolvent mixture. ζ_H and ζ_{TS} are the relative contributions by enthalpy and entropy toward Gibbs energy mixing.

As it was already stated, the energy of the cavity formation process should be lower as the proportion of EtOH increases because the polarity of the medium decreases, which is a fact that favors solute-solvent interactions, except for ion-dipole. This fact is shown in table 4, where ΔH^0_{mix} decreases as the proportion of cosolvent increases in EtOH-rich mixtures. According to Romero *et al.,* 1996 (32), for non-electrolytes in the initial portion of the solubility curve, the hydrogen bonding of the drug will increase according to the EtOH concentration. For large cosolvent proportions (from 0.60 in mass fraction of EtOH to neat EtOH), this interaction may be saturated, becoming a constant contribution. However, nonspecific and cavity effects are not saturated, and vary according to the EtOH concentration. Nevertheless, these considerations do not explain the behavior observed in water-rich mixtures, where the ion-dipole interactions predominate for all sodium sulfonamides.

Enthalpy-Entropy Compensation of sodium sulfonamide mixing process

According to the literature, weight graphs of $\Delta H^{0\text{-app}}_{soln}$ as a function of $\Delta G^{0\text{-app}}_{soln}$ at the mean harmonic temperature allows us to observe similar

mechanisms for the solution process according to the tendencies that can be seen in the results (33, 34).

In this context, figure 2 comprehensively shows that sodium sulfonamides, in the EtOH + water cosolvent system, present non-linear $\Delta H_{soln}^{0\text{-app}}$ vs. $\Delta G_{soln}^{0\text{-app}}$ compensation with negative slopes if the composition interval $0.60 \leq \mu_{EtOH} \leq 1.00$ is considered for all sodium sulfonamides; whereas from this EtOH proportion to neat water, positive slopes are obtained. According to this graph, it can be concluded that the driving function for the drug mixing processes is the entropy in the former case, while in the latter the driving function is the enthalpy. Nevertheless, the molecular and ionic events, which are involved in the dissolution of these drugs in this cosolvent, system are unclear.

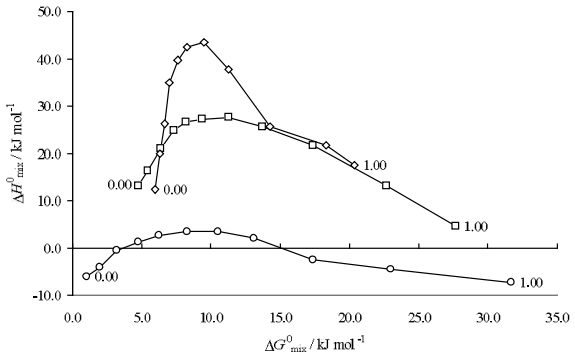

Figure 2. ΔH^0_{mix} vs. ΔG^0_{mix} enthalpy-entropy compensation plot for solubility of sodium sulfonamides in ethanol + water cosolvent mixtures at 292.8 K. (◊): Na-SMT, (□): Na-SMR: (○): Na-SD.

CONCLUSIONS

From all the previously discussed topics, it can be concluded that the mixing processes of these sodium sulfonamides in EtOH + water mixtures is variable according to the cosolvent composition. Also, non linear enthalpy-entropy compensation was found for the drugs evaluated in this cosolvent system. In this context, the solution processes in compositions from pure EtOH to the mixture having 0.60 in mass fraction of ethanol were found to be enthalpy-driven; whereas, the systems from this ethanol proportion to neat water were found to be enthalpy-driven. Finally, it can be stated that the thermodynamic quantities presented in this report broaden the physicochemical information about electrolyte drugs in aqueous and alcoholic solutions.

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