ALVARO JAVIER PATIÑO AGUDELO

EFEITO DE LÍQUIDOS IÔNICOS IMIDAZÓLICOS SOBRE PROCESSOS DE AGREGAÇÃO MOLECULAR ENVOLVENDO SURFACTANTES IÔNICOS

Tese apresentada à Universidade Federal de Viçosa, como parte das exigências do Programa de Pós-Graduação em Agroquímica, para obtenção do título de *Doctor Scientiae.*

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APROVADA: 15 de março de 2019.

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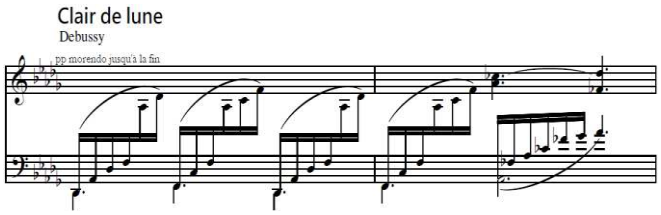

À minha esposa, Andrea Villalobos,

pela magnitude da sensibilidade

da íris dos seus olhos.

Que seja infinita nossa companhia

enquanto a vida passa.

Dedico

"But we must not forget that when radium was discovered no one knew that it would prove useful in hospitals. The work was one of pure science. And this is a proof that scientific work must not be considered from the point of view of the direct usefulness of it. It must be done for itself, for the beauty of science, and then there is always the chance that a scientific discovery may become like the radium a benefit for humanity"

Marie Curie

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SUMÁRIO

LISTA DE SÍMBOLOS E ABREVIAÇÕES

Lista de símbolos e abreviações dos capítulos 1 e 2:

$\boldsymbol{a_i}$**:** atividade química da espécie i;

$\boldsymbol{\alpha}$**:** grau de dissociação micelar;

$\boldsymbol{\alpha CD}$**:** α-ciclodextrina;

$\boldsymbol{\beta CD}$**:** β-ciclodextrina;

βCD-$C_{12}Py^+$:complexo de inclusão formado por β-ciclodextrina e o cátion dodecilpiridínio;

$\boldsymbol{CD}$**:** ciclodextrina;

$\boldsymbol{CD - S}$**:** complexo de inclusão formado por ciclodextrina e surfactante;

$\boldsymbol{cmc}$**:** concentração micelar crítica;

C_nPyBr: brometo de alquilpiridínio;

C_nTAB: brometo de N-alquiltrimetilamônio;

C_{12}PyX: haletos de N-dodecilpiridínio;

C_4mimBF_4: tetrafluoroborato de 1-butil-3-metilimidazólio;

C_4mimBr: brometo de 1-butil-3-metilimidazólio;

C_4mimPF_6: hexafluorofosfato de 1-butil-3-metilimidazólio;

$\boldsymbol{C^+}$**:** contra-íon do monômero de surfactante aniônico;

$\boldsymbol{D}$**:** coeficiente de difusão;

DLS: espalhamento dinâmico de luz;

DDAO: óxido de dodecil(dimetil)amino;

$\Delta\boldsymbol{G_{ic}^o}$**:** variação de energia livre Gibbs padrão de inclusão;

$\Delta\boldsymbol{G_{mic}^o}$**:** variação de energia livre de Gibbs padrão de micelização;

$\Delta\boldsymbol{H_{des}^o}$**:** variação de entalpia padrão da dessolvatação do surfactante;

$\Delta\boldsymbol{H_{des}^{o,CD}}$**:** variação de entalpia padrão da dessolvatação da região hidrofóbica da ciclodextrina;

$\Delta\boldsymbol{H_{elec}^o}$**:** variação de entalpia padrão das contribuições eletrostáticas;

ΔH_{ic}^{o}**:** variação de entalpia padrão de inclusão;

$\Delta H_{int}^{o,CD-S}$**:** variação de entalpia padrão da interação surfactante-ciclodextrina;

ΔH_{mic}^{o}**:** variação de entalpia padrão de micelização;

ΔH_{obs}**:** variação de entalpia observada;

ΔH_{T-T}^{o}**:** variação de entalpia padrão da interação surfactante-surfactante;

$\Delta \mathbf{P}$**:** variação de potência;

$\Delta S_{des}^{o,CD}$**:** variação de entropia padrão da liberação de moléculas de água da região hidrofóbica da ciclodextrina;

$\Delta S_{des}^{o,S}$**:** variação de entropia padrão da liberação de moléculas de água da região hidrofóbica do surfactante;

ΔS_{ic}^{o}**:** variação de entropia padrão de inclusão;

ΔS_{mic}^{o}**:** variação de entropia padrão de micelização;

$\Delta S_{tras-rota-conf}^{o,CD-S}$**:** variação de entropia padrão do número de graus de liberdade translacional, rotacional e conformacional das moléculas de surfactante e ciclodextrina;

$g_{(\tau)}^{(2)}$**:** função de correlação de segunda ordem;

γ**:** tensão interfacial;

γCD**:** γ-ciclodextrina;

H_i e H_f**:** entalpias do sistema nos estados termodinâmicos *i* e *f*, respectivamente;

ITC: calorimetria de titulação isotérmica;

I_1/I_3**:** coeficiente da intensidade das bandas vibracionais do espectro de emissão do pireno;

$\vec{j}$**:** fluxo do número efetivo de partículas por mol;

k**:** condutividade elétrica;

K_B**:** constante de Bolztmann;

K_{ic}**:** constante de equilibrio para o processo de inclusão;

K_{mic}**:** constante de equilíbrio para o processo de micelização;

K_W**:** constante de Wiseman;

LIs: líquidos iônicos;

$\boldsymbol{\lambda_{em}}$**:** comprimento de onda de emissão;

$\boldsymbol{\lambda_{ex}}$**:** comprimento de onda de excitação;

M: micela de surfactante;

$\boldsymbol{m}$**:** número de monômeros de surfactante;

$\boldsymbol{n}$**:** sítios de ligação;

$\boldsymbol{n_r}$**:** índice de refração da solução;

$\boldsymbol{\vec{\nabla}c}$**:** gradiente de concentração;

P: pressão;

$\boldsymbol{p^-}$**:** carga superficial efetiva da micela;

$\boldsymbol{q}$**:** energia trocada entre o sistema e vizinhança na forma de calor;

$\boldsymbol{q_{es}}$: vetor espalhamento;

$\boldsymbol{R}$**:** constante universal dos gases;

$\boldsymbol{R_H}$**:** raio hidrodinâmico;

S: surfactante;

S0, **S1** e **S2**: estados eletrônicos singletes representando o estado fundamental, o primeiro e o segundo estado excitado, respectivamente;

SDBS: dodecilbenzeno sulfonato de sódio;

SDS: dodecilsulfato de sódio;

SDSN: dodecilsulfonato de sódio;

$\boldsymbol{T}$**:** temperatura absoluta;

$\boldsymbol{T^-}$**:** monômero do surfactante aniônico ionizado;

$\boldsymbol{t}$**:** tempo;

$\boldsymbol{t_1}$ **e** $\boldsymbol{t_2}$**:** tempo no início e no final de cada injeção, respectivamente, em um experimento de
calorimetria de titulação isotérmica;

$\boldsymbol{U}$**:** energia interna;

$\boldsymbol{U_i}$ e $\boldsymbol{U_f}$**:** energias internas do sistema nos estados termodinâmicos *i* e *f*, respectivamente;

V_c**:** volume efetivo da cela calorimétrica;

V_i **e** V_f**:** volumes do sistema nos estados termodinâmicos *i* e *f*, respectivamente;

$\overline{v}$**:** parâmetro de ligação;

w**:** energia trocada entre o sistema e vizinhança na forma de trabalho.

Lista de símbolos e abreviações dos capítulos 3 e 4:

α**:** degree of counter-ion dissociation;

βCD: β-cyclodextrin;

βCD-$C_{12}Py^+$: inclusion complex formed by β-cyclodextrin and 1-dodecylpiridinium cation;

βCD-C_6mim^+: inclusion complex formed by β-cyclodextrin and 1-hexyl-3-methylimidazolium cation;

βCD-C_8mim^+: inclusion complex formed by β-cyclodextrin and 1-octyl-3-methylimidazolium cation;

cmc**:** critical micellar concentration;

C_nmim^+: 1-alkyl-3-methylimidazolium cation;

$C_n$mimX: 1-alkyl-3-methylimidazolium halides;

C_0mimCl: 1-methylimidazolium chloride;

C_{12}PyCl: 1-dodecylpyridinium chloride;

$C_{12}Py^+$: 1-dodecylpiridinium cation;

C_2mimCl: 1-ethyl-3-methylimidazolium chloride;

C_4mimBF_4: 1-butyl-3-methylimidazolium tetrafluoroborate;

C_4mimBr: 1-butyl-3-methylimidazolium bromide;

C_4mimCl: 1-butyl-3-methylimidazolium chloride;

C_4mimI: 1-butyl-3-methylimidazolium iodide;

C_4mimPF_6: 1-butyl-3-methylimidazolium hexafluorophosphate;

C_6mimCl: 1-hexyl-3-methylimidazolium chloride;

C_8mimCl: 1-octyl -3-methylimidazolium chloride;

DLS: dynamic light scattering;

$\boldsymbol{D_t}$**:** diffusion coefficient;

$\boldsymbol{\Delta G_{ic}^{o}}$**:** standard Gibbs free energy change of the inclusion complex formation;

$\boldsymbol{\Delta G_{mic}^{o}}$**:** standard Gibbs free energy change of micellization;

$\boldsymbol{\Delta H_{ag1}^{o}}$**:** standard enthalpy change when a mole of surfactant change from its monomeric state to its small aggregate state;

$\boldsymbol{\Delta H_{ag2}^{o}}$**:** standard enthalpy change when a mole of surfactant change from its small aggregate state to its micellar state;

$\boldsymbol{\Delta H_{ag3}^{o}}$**:** standard enthalpy change in post-micellar region;

$\boldsymbol{\Delta H_{ap-int}}$**:** apparent change in molar enthalpy of the interaction;

$\boldsymbol{\Delta H_{demic}}$**:** demicellization enthalpy change;

$\boldsymbol{\Delta H_{des}^{o,\beta CD}}$**:** standard enthalpy change of the β-cyclodextrin desolvation;

$\boldsymbol{\Delta H_{des}^{o,\beta CD-surf}}$**:** standard enthalpy change of the β-cyclodextrin-surfactant interaction;

$\boldsymbol{\Delta H_{des}^{o,surf}}$**:** standard enthalpy change of the surfactant desolvation;

$\boldsymbol{\Delta H_{elec}^{o}}$**:** standard enthalpy change of the electrostatic repulsion between the head groups of the surfactant;

$\boldsymbol{\Delta H_{ic}^{o}}$**:** standard enthalpy change of the inclusion complex formation;

$\boldsymbol{\Delta H_{int}}$**:** enthalpy change of the surfactant-cyclodextrin interaction;

$\boldsymbol{\Delta H_{mic}^{o}}$**:** standard enthalpy change of micellization;

$\boldsymbol{\Delta H_{mic}^{dil}}$**:** enthalpy change for the micelle dilution;

$\boldsymbol{\Delta H_{obs}}$**:** observed enthalpy change per mol of surfactant added into the sample calorimetric cell;

$\boldsymbol{\Delta H_{surf-surf}^{o}}$**:** standard enthalpy change of the surfactant-surfactant interactions;

$\boldsymbol{\Delta S_{ic}^{o}}$**:** standard entropy change of the inclusion complex formation;

$\boldsymbol{\Delta S_{mic}^{o}}$**:** standard entropy change of micellization;

$\boldsymbol{\Delta S_{rel}^{o,H_20-\beta CD}}$**:** standard entropy change of the water molecules releasing from the solvation shell of βCD cavity;

$\Delta S_{rel}^{o,H_20-surf}$**:** standard entropy change of the water molecules releasing from the solvation shell of surfactant;

$\Delta S_{tras-rota-conf}^{o,\beta CD-surf}$**:** standard entropy change of the number of translational-rotational-conformational degrees of freedom for surfactant and βCD molecules;

ILs: ionic liquids;

ITC: isothermal titration calorimetry;

I_1/I_3**:** quotient of pyrene vibrational band intensities;

K_{ic}**:** inclusion binding constant;

K_B**:** Boltzmann's constant;

λ_{max}**:** maximum emission wavelength;

n**:** stoichiometric ratio;

η**:** viscosity of suspension;

PEO: poly(ethylene oxide);

r**:** molar ratio between 1-dodecylpiridinium cation and β-cyclodextrin species;

R: gas constant;

R_H**:** hydrodynamic radius

SDBS: sodium dodecylbenzene sulfonate;

SDS: sodium dodecyl sulfate;

T: absolute temperature;

TAM III: thermal activity monitor;

T_d**:** denaturation protein;

X_{mic}**:** fraction of 1-dodecylpiridinium chloride monomers that formed micelles;

3D: three-dimensional structure of water.

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RESUMO

AGUDELO, Álvaro Javier Patiño, D.Sc., Universidade Federal de Viçosa, março de 2019. **Efeito de líquidos iônicos imidazólicos sobre processos de agregação molecular envolvendo surfactantes iônicos.** Orientador: Luis Henrique Mendes da Silva. Coorientadoras: Maria do Carmo Hespanhol e Ana Clarissa dos Santos Pires.

O estudo termodinâmico de processos de associação molecular em misturas aquosas de líquidos iônicos (LIs) é essencial para o entendimento e controle da solvatação de compostos anfifílicos que participam na ciência dos coloides e da química supramolecular. Entretanto, o efeito dos LIs sobre a termodinâmica dos processos de micelização de surfactantes aniônicos não modelo e de formação de complexos de inclusão entre surfactantes e compostos macromoleculares, permanece inexplorado. Neste trabalho foi estudada a termodinâmica de dois processos de agregação molecular: *i*) micelização do surfactante dodecilbenzeno sulfonato de sódio (SDBS) em soluções aquosas de cloreto de 1-alquil-3-metilimidazólio ($C_n$mimCl, n = 0, 2 ou 4) usando espectroscopia de fluorescência, calorimetria de titulação isotérmica (ITC), espalhamento dinâmico de luz e condutividade elétrica e *ii*) formação de complexos de inclusão entre o cation dodecilpiridinio ($C_{12}Py^+$) e β-ciclodextrina (βCD) em misturas formadas por água e haletos de 1-alquil-3-metilimidazólio ($C_n$mimX, sendo n = 0, 2, 4, 6 ou 8 e X = Cl^-, Br^- ou I^-) empregando ITC. O primeiro processo termodinâmico estudado mostrou que a estabilidade termodinâmica dos agregados de SDBS dependeu fortemente da estrutura molecular e concentração do LI, diminuindo na ordem $C_4mim^+ > C_0mim^+ > C_2mim^+$. Na máxima concentração estudada de LIs (1,00 mmol L^{-1}), o incremento da hidrofobicidade do cation imidazólio diminuiu o favorecimento entálpico, aumentando o ΔH^o_{mic} de −3,75 ± 0,07 kJ mol^{-1} (C_0mim^+) para −2,69 ± 0,01 kJ mol^{-1} (C_4mim^+). Por outro lado, o ganho entrópico foi afetado na tendência contrária, sugerindo que a alta hidrofobicidade do C_4mim^+ se sobrepõe ao efeito cosmotrópico dos LIs no *bulk* da solução. Foi sugerido que os cations imidazólios podem interagir com os monômeros de SDBS na superfície da micela, sendo esta interação, principalmente, via interações hidrofóbicas, π-π e eletrostáticas para o C_4mim^+ e C_2mim^+, e interações eletrostáticas e ligações de hidrogênio para o C_0mim^+. Para o segundo processo de agregação estudado, $C_{12}Py^+$ e βCD formaram complexos de inclusão com estequiometria 1:1 em água pura, sendo o processo entalpicamente ($\Delta H^o_{ic} = -9{,}2 \pm 0{,}1$ kJ mol^{-1}) e entropicamente ($T\Delta S^o_{ic} = 16{,}1 \pm 0{,}2$ kJ mol^{-1}) favorável. Entretanto, em soluções aquosas de LIs, todos os parâmetros termodinâmicos da complexação do βCD–$C_{12}Py^+$ variaram, dependendo tanto da

concentração como do tamanho da cadeia alquílica do cátion C_nmim^+. Para LIs de cadeia curta (C_nmim^+, $n \leq 4$), a entropia do sistema diminuiu, enquanto para LIs de cadeia longa (C_nmim^+, $n \geq 6$), foi observado uma redução dos valores de entalpia de ligação. Estes efeitos foram atribuídos à solvatação preferencial das cadeias dos surfactantes pelos LIs, à capacidade dos LIs em modificar a estrutura terciaria da água e à inclusão de moléculas de LI na cavidade da βCD.

ABSTRACT

AGUDELO, Álvaro Javier Patiño, D.Sc., Universidade Federal de Viçosa, March, 2019. **Effect of imidazole ionic liquids on processes of molecular aggregation involving ionic surfactants.** Adviser: Luis Henrique Mendes da Silva. Co-advisers: Maria do Carmo Hespanhol and Ana Clarissa dos Santos Pires.

Thermodynamic investigation of molecular association processes in aqueous mixtures containing ionic liquids (ILs) is essential to elucidate and control the solvation of amphiphilic compounds involved in colloid science and supramolecular chemistry. However, the effects of ILs on the thermodynamics of micellization processes of atypical anionic surfactants and on the formation of inclusion complexes between surfactants and macromolecular compounds remain unexplored. In this work, the thermodynamics of two molecular aggregation processes has been investigated: *i*) the micellization of the anionic surfactant sodium dodecylbenzene sulfonate (SDBS) in aqueous solutions of 1-alkyl-3-methylimidazolium chloride ($C_n$mimCl, n = 0, 2 or 4) using fluorescence spectroscopy, isothermal titration calorimetry (ITC), dynamic light scattering, and electrical conductimetry measurements; and *ii*) the formation of inclusion complexes between the cation dodecylpyridinium ($C_{12}Py^+$) and β-cyclodextrin (βCD) in mixtures formed with water and 1-alkyl-3-methylimidazolium halides ($C_n$mimX, n = 0, 2, 4, 6 or 8, and X = Cl^-, Br^- or I^-) analyzed by ITC. The first examined process showed that the thermodynamic stability of SDBS aggregates strongly depended on the molecular structure and concentration of IL, decreasing in the order $C_4mim^+ > C_0mim^+ > C_2mim^+$. For the maximum concentration of ILs (1.00 mmol L^{-1}), the increase in the hydrophobicity of the imidazolium cation decreased the extent of enthalpic favorableness, increasing ΔH^o_{mic} values from −3.75 ± 0.07 kJ mol^{-1} (C_0mim^+) to −2.69 ± 0.01 kJ mol^{-1} (C_4mim^+). On the other hand, the entropic contribution was affected in an opposite way, suggesting that the high hydrophobicity of C_4mim^+ can overcome the kosmotropic effect of ILs in the bulk of the solutions. It has also been proposed that the imidazolium cations can interact with SDBS monomers on the micelle surface, mainly by means of hydrophobic, π- π and electrostatic interactions for C_4mim^+ and C_2mim^+, and electrostatic interactions and hydrogen bonds for C_0mim^+. In the second investigation, $C_{12}Py^+$ and βCD formed complexes with 1:1 stoichiometry in pure water, and the process was both enthalpically ($\Delta H^o_{ic} = -9.2 \pm 0.1$ kJ mol^{-1}) and entropically ($T\Delta S^o_{ic} = 16.1 \pm 0.2$ kJ mol^{-1}) favorable. Nevertheless, in IL aqueous solutions, all thermodynamic parameters related to the βCD–$C_{12}Py^+$ complexation depended on both concentration and size of the alkyl

chain of the C_nmim^+ cation. For short-chain ILs (C_nmim^+, $n \leq 4$), the system entropy decreased, while for long-chain ILs (C_nmim^+, $n \geq 6$), the main effect was a decrease in the bond enthalpy values. These observations have been attributed to the preferential solvation of the surfactant chains by the ILs, to the ability of ILs to modify the tertiary structure of water molecules, and to the inclusion of IL molecules in the cavity of βCD.

Capítulo 1: Introdução, estado da arte e fundamentação experimental

1.1. *Introdução*

O estudo físico-químico da associação entre moléculas anfifílicas é essencial para o entendimento do desenvolvimento da vida em nível molecular. Entre os diversos processos de agregação existentes, a micelização de surfactantes e a formação de complexos supramoleculares entre compostos macrocíclicos e surfactantes têm sido amplamente estudados nas últimas décadas. As nanoestruturas formadas por estes compostos são fundamentais para estudos envolvendo reconhecimento molecular e para aplicações em liberação controlada de drogas, produção de cosméticos, remediação de solos e descontaminação de ambientes aquáticos.

Em água pura, os processos mencionados são dirigidos principalmente pelo aumento de entropia do sistema, dominado pela diferença entre a rede tridimensional de moléculas de água solvatando as regiões hidrofóbicas dos compostos anfifílicos livres (com baixa entropia e entalpia molar) e aquela concentrada no *bulk* da solução (com alta entropia e entalpia molar). Assim, quando a diferença de entropia entre as duas populações de água não é termodinamicamente compensada pelas interações entálpicas, ocorrendo no estado monomérico, o processo de associação molecular ocorre. Por outro lado, na presença de cossolutos, as duas populações de água têm suas redes tridimensionais afetadas, alterando as contribuições entálpicas e entrópicas que regem os processos de agregação molecular.

Os cossolutos são classificados em agentes cosmotrópicos (compostos que diminuem a entropia molar da água) e agentes caotrópicos (compostos que aumentam a entropia molar da água). Em geral, agentes cosmotrópicos são íons com alta densidade de carga (por exemplo Li^+, Na^+, Mg^+ ou Ca^{2+}) que estabilizam termodinamicamente os agregados moleculares, enquanto agentes caotrópicos são íons com baixa densidade de carga (por exemplo K^+, Cs^+, I^- ou Br^-) que desestabilizam termodinamicamente os processos de associação molecular. Portanto, a seleção apropriada de um cossoluto é crucial para monitorar e entender o papel das moléculas de água sobre a cinética e a termodinâmica dos processos de agregação molecular.

Nos últimos 15 anos, uma classe de cossolutos que tem chamado a atenção são os líquidos iônicos imidazólicos (LIs), definidos como sais líquidos em temperaturas menores que o ponto de ebulição da água, com propriedades estratégicas como alta condutividade, baixa pressão de vapor e baixa toxicidade. Sua estrutura química é baseada num cátion imidazólio que possui uma carga deslocalizada, ideal para avaliar simultaneamente interações não covalentes como interações hidrofóbicas, eletrostáticas, π-π e ligações de hidrogênio, sendo uma limitação dos cossolutos clássicos.

Os LIs vêm ocupando um lugar cada vez mais importante em aplicações práticas e na pesquisa básica envolvendo sistemas coloidais onde ocorre a formação de agregados. Existe na literatura científica uma extensa informação relacionada à formação de micelas de surfactantes em misturas de água e LIs, mas nenhum relato de seu efeito sobre a formação de complexos de inclusão. Neste contexto, muitas questões fundamentais ainda precisam ser respondidas, dentre as quais encontram-se: *i*) qual é o efeito de LIs imidazólicos sobre a termodinâmica de micelização de surfactantes aniônicos contendo anéis aromáticos ligados em seus grupos cabeça? *ii*) o balanço hidrofóbico/hidrofílico dos LIs afeta igualmente os parâmetros termodinâmicos associados à formação de micelas e complexos de inclusão? *iii*) existe uma competição entre moléculas de água e cátions imidazólicos pelas superfícies hidrofóbicas das ciclodextrinas e surfactantes simultaneamente? *iv*) qual é a relação entre íons cosmotrópicos e caotrópicos clássicos com o efeito promovido pelos cátions imidazólicos?

Para abordar o entendimento destas questões, nesta tese foi proposto avaliar o efeito de LIs imidazólicos sobre a termodinâmica de micelização do surfactante dodecilbenzeno sulfonato de sódio (SDBS) e a termodinâmica de inclusão do cation dodecilpiridinio ($C_{12}Py^+$) em β-ciclodextrina em solução aquosa.

1.2. *Organização da tese*

O corpo estrutural desta tese está baseado em 5 capítulos que foram construídos e apresentados de forma estratégica, para direcionar ao leitor na compreensão dos desafios científicos presentes ao longo destes 4 anos de formação doutoral.

Este capitulo (escrito em português) refere-se à compilação da informação teórico-experimental que proporciona as bases físico-químicas dos dois processos de associação molecular abordados nesta tese de doutorado, a micelização de surfactantes iônicos e a formação de complexos de inclusão entre compostos macrocíclicos e

detergentes em solução aquosa. Além do anterior, uma descrição histórica e molecular dos líquidos iônicos imidazólicos é proporcionada.

O *Capitulo 2 (escrito em português)* relata os objetivos gerais e específicos que foram os precursores para o desenvolvimento das duas propostas apresentadas nos capítulos 3 e 4.

O *Capitulo 3 (escrito em inglês)* trata do efeito de líquidos iônicos imidazólicos de cadeia curta sobre a termodinâmica de micelização do surfactante aniônico SDBS.

O *Capitulo 4 (escrito em inglês)* aborda a termodinâmica de formação de complexos de inclusão entre βCD e $C_{12}Py^+$ em solução aquosa de haletos de 1-alquil-3-metilimidazólio.

Ambos os capítulos 3 e 4, estão escritos em formato de artigos científicos e na presente data estão em fase de submissão em revistas cientificas de relevância internacional, na área de físico-química de coloides e química supramolecular.

O *Capitulo 5 (escrito em português)* apresenta as conclusões gerais desta tese de doutorado.

1.3. *Estado da arte*

Esta seção apresenta um resumo dos conceitos básicos, além das informações mais relevantes presentes na literatura científica sobre a termodinâmica de micelização de surfactantes iônicos, formação de complexos de inclusão surfactante-ciclodextrina e efeito de líquidos iônicos sobre estes processos de agregação molecular em solução aquosa.

1.3.1. Micelização de surfactantes iônicos

1.3.1.1. Surfactantes: propriedades físicas e químicas

Os surfactantes são uma classe de compostos químicos com a capacidade de alterar a tensão superficial de soluções aquosas. Essa característica deve-se à natureza anfifílica de sua estrutura molecular, uma região hidrofóbica (chamada de cauda) e uma região hidrofílica (chamada de cabeça).[1] Geralmente a cauda do surfactante é constituída por ligações covalentes simples entre átomos de carbono,[2] em que a baixa polarizabilidade entre eles confere pouca afinidade pelas moléculas de água. A cabeça do surfactante é formada por grupos de átomos onde a somatória do momento dipolar é

diferente de zero, por exemplo, óxido de etileno,[3] grupos funcionais como o sulfato[4] e o fosfato,[5] compostos heterocíclicos como o piridinio[6] e o imidazólio,[7] e sais quaternárias de nitrogênio,[8] que permitem uma afinidade com moléculas de água. A Figura 1.1 apresenta a estrutura do surfactante dodecilbenzeno sulfonato de sódio (SDBS).

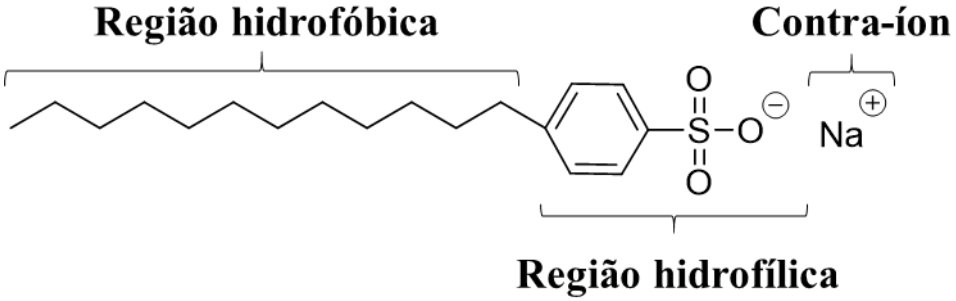

Figura 1.1. Estrutura molecular do surfactante aniônico SDBS. São destacadas nos colchetes as regiões hidrofóbica e hidrofílica, assim como o contra-íon.

Os surfactantes são classificados em não iônicos, iônicos e zwitteriônicos. Surfactantes não iônicos são moléculas anfifílicas com ausência de carga, enquanto os surfactantes iônicos possuem carga no grupo cabeça, podendo ser negativa (surfactante aniônico) ou positiva (surfactante catiônico). Os surfactantes zwitteriônicos são os mais complexos, visto que a carga global do mesmo é zero, mas na sua estrutura química estão presentes dois grupos polares com cargas opostas.

Em solução aquosa, o comportamento dos compostos anfifílicos é bem definido e sua representação é mostrada na Figura 1.2.

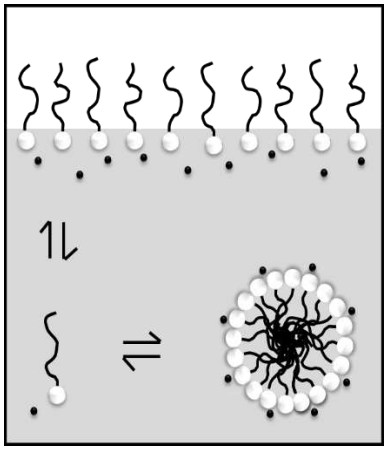

Figura 1.2. Micelização de surfactantes iônicos em solução aquosa.

Quando pequenas quantidades de surfactante são colocadas em um sistema formado por água e ar, é observada uma diminuição da tensão interfacial (γ) da água, proporcional à concentração do tensoativo em solução ($[S]_{sol}$). Isso se deve ao equilíbrio de moléculas livres de surfactante (monômeros) presentes no *bulk* da solução e adsorvidas na interface, onde o grupo cabeça interage com a água e a cauda com o ar.[9] O aumento da concentração do surfactante até a concentração micelar crítica (cmc), com saturação da interface, favorece a agregação dos monômeros em micelas. A cmc pode ser determinada por diferentes técnicas instrumentais a partir do ponto de inflexão (ou descontinuidade) observado quando uma propriedade física é medida em função da concentração do detergente (Figura 1.3).

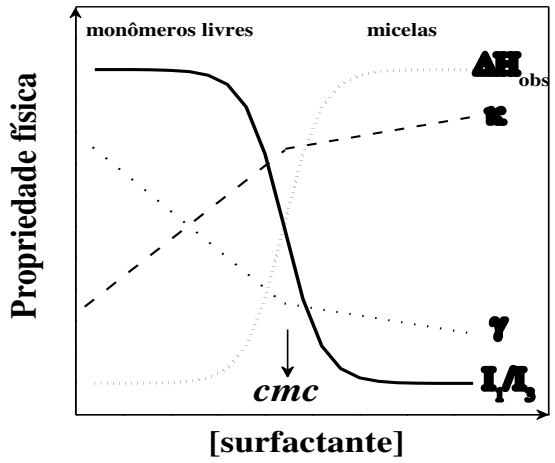

Figura 1.3. Comportamento de algumas propriedades físicas associado à formação de micelas de surfactante. Propriedades físicas: (ΔH_{obs}) variação de entalpia observada (se o processo for endotérmico),[10] (k) condutividade elétrica,[11] (γ) tensão interfacial[12] e (I_1/I_3) razão das intensidades das bandas vibrônicas do pireno.[13]

A formação de micelas por surfactantes em água é um processo entropicamente dirigido e, para o seu entendimento, é necessário fazer uma avaliação das moléculas de água solvatando as regiões hidrofóbicas do surfactante (grupo 1) e presentes no *bulk* da solução (grupo 2).[10]

Apesar das moléculas de água dos grupos 1 e 2 fazerem parte do mesmo sistema, elas possuem características distintas. As moléculas de água do grupo 1 tem baixa entropia e entalpia molar comparada com as moléculas do grupo 2. Para solvatar a região hidrofóbica do surfactante, as moléculas de água devem maximizar as ligações de

hidrogênio entre elas para compensar a fraca interação com a cauda do surfactante, promovendo a perda de suas entropias configuracional e rotacional.

No limite, quando a concentração do surfactante em solução ($[S]_{sol}$) tende ao valor da cmc, a diferença da entropia molar entre as populações de água dos grupos 1 e 2 é tão elevada que quando $[S]_{sol} = cmc$, a natureza a fim de diminuir a energia livre de Gibbs do sistema, libera as moléculas de água que estavam solvatando as caudas dos monômeros do surfactante para o *bulk* da solução, levando à formação de micelas.

Para um maior entendimento deste processo, a seguir será apresentada uma abordagem termodinâmica sobre agregação de surfactantes.

1.3.1.2. Termodinâmica de micelização de surfactantes iônicos

A termodinâmica química é uma das áreas da físico-química que permite a caracterização completa das forças motrizes que regem o processo de micelização de surfactantes, tais como variação da energia livre de Gibbs padrão de micelização (ΔG^o_{mic}), variação da entalpia padrão de micelização (ΔH^o_{mic}) e variação da entropia padrão de micelização ($T\Delta S^o_{mic}$).

Em solução aquosa, o processo de micelização para um tensoativo aniônico (T^-C^+) pode ser representado na seguinte equação química 1.1:

$$mT^-{}_{(aq)} + (m-p)C^+{}_{(aq)} \rightleftharpoons M^{p-}{}_{(aq)} \qquad 1.1$$

em que T^- é o monômero do surfactante ionizado, C^+ é o contra-íon, e M é a micela formada com m monômeros e com carga superficial efetiva p^-. A constante de equilíbrio (K_{mic}) para esse processo de micelização é dada pela equação 1.2:

$$K_{mic} = \frac{a_{M^{p-}}}{(a_{C^+})^{m-p}(a_{T^-})^m} \qquad 1.2$$

onde a_i representa a atividade química da espécie i em solução. Para trabalhos experimentais onde a concentração dos analitos está no regime de diluição infinita, as atividades das espécies são equivalentes às suas concentrações, assim a equação 1.2 fica;

$$K_{mic} = \frac{[M^{p-}]}{[C^+]^{m-p}[T^-]^m} \quad 1.3$$

aplicando o logaritmo de base e na equação 1.3, temos a equação 1.4:

$$lnK_{mic} = ln[M^{p-}] - (m-p)ln[C^+] - mln[T^-] \quad 1.4$$

Usando a K_{mic} é possível calcular o ΔG^o_{mic} fazendo uso da relação fundamental de Gibbs[14] (equação 1.5).

$$\Delta G^o_{mic} = -RTlnK_{mic} \quad 1.5$$

em que R é a constante universal dos gases (R = 8,31 J K^{-1} mol^{-1}) e T é a temperatura (298,15 K) na qual realiza-se o experimento. Substituindo a equação 1.4 na equação 1.5, é obtida a equação 1.6:

$$\Delta G^o_{mic} = -RT[ln[M^{p-}] - (m-p)ln[C^+] - mln[T^-]] \quad 1.6$$

A equação 1.6 pode ser expressada por mol de surfactante, dividindo cada membro por m, resultando na equação 1.7:

$$\Delta G^o_{mic} = RT\left[-\frac{ln[M^{p-}]}{m} + \left(1 - \frac{p}{m}\right)ln[C^+] + ln[T^-]\right] \quad 1.7$$

Já que o número de monômeros (m) envolvido na formação das micelas de surfactante é elevado, na equação 1.7 podem ser realizadas algumas aproximações. O termo $-ln[M^{p-}]/m$ pode ser desconsiderado e, em água pura, a $[T^-]$ pode ser aproximada ao valor da cmc, enquanto $[C^+]$ é igual à cmc. Desta forma a equação 1.7 pode ser reescrita como (equação 1.8):

$$\Delta G^o_{mic} = RT(2-\alpha)ln(cmc) \quad 1.8$$

em que $\alpha = p/m$ é denominado como o grau de dissociação micelar. α expressa o número de C^+ dissociados no *bulk* da solução que fazem parte dos monômeros de T^- que formam o agregado molecular. O ΔG^o_{mic} é a variação da energia livre de Gibbs padrão de

micelização quando 1 mol de surfactante no estado monomérico forma micelas com carga superficial p^- em solução.

Para entender o significado do potencial termodinâmico do processo de micelização é necessário avaliar a energética associada às interações intermoleculares (ΔH^o_{mic}) e os graus de liberdade ($T\Delta S^o_{mic}$) das espécies químicas envolvidas no processo de agregação molecular. A variação da entalpia padrão de micelização (ΔH^o_{mic}) pode ser considerada como a somatória de 3 eventos ocorrendo simultaneamente (equação 1.9):

$$\Delta H^o_{mic} = \Delta H^o_{des} + \Delta H^o_{elec} + \Delta H^o_{T-T} \qquad 1.9$$

onde ΔH^o_{des}, ΔH^o_{elec} e ΔH^o_{T-T} são as variações de entalpia padrão da dessolvatação do surfactante, das contribuições eletrostáticas e das interações surfactante-surfactante, respectivamente. O ΔH^o_{des} é positivo, uma vez que quebrar interações água-água sobre uma superfície hidrofóbica demanda mais energia do que a liberada pelas interações água-água no *bulk* da solução, quando ocorre o processo de micelização.[15] O ΔH^o_{elec} é positivo devido principalmente à repulsão eletrostática gerada pelos grupo cabeça do surfactante no estado micelar.[16] O ΔH^o_{T-T} é uma contribuição entalpicamente favorável, já que representa as interações fracas exercidas pelos tensoativos no estado micelar, sendo estas dependentes da natureza molecular do detergente. Determinando o ΔG^o_{mic} e ΔH^o_{mic}, a variação da entropia padrão de micelização ($T\Delta S^o_{mic}$) pode ser calculada, utilizando a equação de Gibbs (equação 1.10):[17]

$$\Delta G^o_{mic} = \Delta H^o_{mic} - T\Delta S^o_{mic} \qquad 1.10$$

A variação da entropia padrão de micelização ($T\Delta S^o_{mic}$) possui as mesmas componentes apresentadas para o ΔH^o_{mic} (equação 1.9), entretanto, a responsável principal pela magnitude e o sinal (positivo) deste parâmetro termodinâmico é a contribuição vinda do ganho entrópico promovido pela liberação de moléculas de água das regiões hidrofóbicas do surfactante para o *bulk* da solução quando as micelas de tensoativo são formadas.

Depois de analisar a termodinâmica de micelização de surfactantes iônicos em água pura, será apresentado como a natureza molecular dos responsáveis pela micelização (surfactantes e solvente) e a variação da temperatura, afetam este processo termodinâmico.

1.3.1.3. Consequências da temperatura e das características moleculares do surfactante e solvente sobre o processo de micelização

A micelização de surfactantes iônicos depende drasticamente da *i*) estrutura molecular do surfactante, *ii*) natureza do solvente e *iii*) temperatura.

i) Estrutura molecular do surfactante: quando a agregação de surfactantes iônicos é monitorada em água pura e temperatura de 298,15 K, parâmetros estruturais do detergente como tamanho da região hidrofóbica, natureza do contra-íon e estrutura molecular do grupo cabeça, influenciam os parâmetros termodinâmicos de micelização. Aqui será apresentado um exemplo que abrange a generalidade para cada efeito.

Bezan e colaboradores,[18] estudaram a agregação de brometos de alquilpiridínio (C_nPyBr, $11 \leq n \leq 15$). O aumento do n incrementou a estabilidade termodinâmica das micelas de C_nPyBr, evidenciado pela diminuição dos valores de ΔG^o_{mic}, passando de −33,5 (n = 11) para −48,6 kJ mol^{-1} (n = 15). Este efeito foi promovido principalmente pelo ganho entrópico, vindo da liberação de um maior número de moléculas de água das regiões hidrofóbicas dos monômeros para o *bulk* da solução.

Rosen et al,[19] monitoraram a micelização de haletos de N-dodecilpiridínio ($C_{12}PyX$, X = Cl^- ou Br^-). O valor da cmc foi 17,8 e 11,3 mmol L^{-1} para o $C_{12}PyCl$ e $C_{12}PyBr$, respectivamente. Esta diferença está relacionada à hidratação do ânion em solução (Cl^- > Br^-), que promove diferenças mais abruptas entre a entropia molar da população de moléculas de água solvatando a cauda dos monômeros e a presente no *bulk* da solução. A estabilidade termodinâmica das micelas de $C_{12}Py^+$ foi maior quando o contra-íon foi Br^- ($\Delta G^o_{mic,Br^-} = -20{,}1$ kJ mol^{-1}; $\Delta G^o_{mic,Cl^-} = -18{,}9$ kJ mol^{-1}). Entretanto, os valores de $T\Delta S^o_{mic}$ e ΔH^o_{mic} foram maiores para o surfactante com o contra-íon Cl^-, sugerindo que um maior número de moléculas de água são liberadas das camadas de solvatação do ânion com maior densidade de carga quando a micelização acontece.

Para examinar o efeito do grupo cabeça, podem ser comparados os surfactantes dodecilbenzeno sulfonato de sódio (SDBS)[20] e dodecilsulfato de sódio (SDS).[21] Ambos os tensoativos, possuem o mesmo comprimento da cadeia alquílica (12 carbonos), sendo que a principal diferença molecular entre estes é a troca de um átomo de oxigênio (SDS) por um anel benzênico (SDBS). O valor da cmc foi menor para o surfactante mais hidrofóbico ($cmc_{SDBS} = 2{,}8$ mmol L^{-1}; $cmc_{SDS} = 8{,}2$ mmol L^{-1}). Consequentemente, um

maior ganho entrópico ($T\Delta S^o_{mic,SDBS} = 26{,}0$ kJ mol^{-1}; $T\Delta S^o_{mic,SDS} = 14{,}9$ kJ mol^{-1}) e menos energia entálpica liberada ($\Delta H^o_{mic,SDBS} = -2{,}5$ kJ mol^{-1}; $\Delta H^o_{mic,SDS} = -6{,}2$ kJ mol^{-1}) foram determinados para o SDBS quando comparado com o SDS.

ii) Efeito da natureza do solvente: misturas aquosas contento aditivos inorgânicos ou orgânicos (classificados como polares e não polares) permitem controlar as propriedades termodinâmicas de micelização de um surfactante sobre uma temperatura definida. Neste item será destacado o efeito de cossolutos inorgânicos sobre a micelização do SDBS à temperatura de 298,15 K.

Geralmente, a presença de cossolutos inorgânicos faz diminuir a cmc de surfactantes iônicos, e alta quantidade de sal promove o aumento do tamanho das micelas. Sood et al,[22] avaliaram o efeito de sais da série Hofmeister sobre a cmc e tamanhos dos agregados de SDBS. A cmc diminui exponencialmente com o aumento da concentração do eletrólito em solução na ordem $MgCl_2$ > KCl > NH_4Cl > NaCl. Além disso, o aumento da concentração de sal até 2 mmol L^{-1}, aumentou o número de agregação das micelas, consequentemente aumentando seu diâmetro hidrodinâmico. Estes fenômenos acontecem porque o aumento da concentração de eletrólito dentro do sistema promove uma diminuição do α, devido a uma maior concentração de íons que diminuem a repulsão eletrostática entre os grupos cabeça na superfície da micela. Este fenômeno depende da cosmotropicidade dos ions em análise, uma vez que este fator está ligado a uma maior blindagem eletrostática.[23–27]

iii) Efeito da temperatura: a modulação desta propriedade termodinâmica produz efeito sobre a hidratação de superfícies hidrofóbicas e hidrofílicas, portanto processos de agregação envolvendo moléculas anfifílicas são afetados. Chauhan et al.,[28] analisaram o efeito da temperatura sobre a cmc do surfactante SDBS em água pura. O aumento da temperatura de 293,15 para 313,15 K aumentou linearmente a cmc do SDBS de 1,2 para 1,51 mmol L^{-1}.

No estado monomérico, as frações hidrofóbica e hidrofílica do detergente estão hidratadas, sendo permitida apenas a hidratação da superfície hidrofílica quando os agregados são formados. Com o aumento da temperatura, o grau de hidratação da cauda e da região hidrofílica diminuem, estes fatores promovem efeitos diferentes, sendo que o primeiro favorece e o segundo desfavorece a agregação. No caso do SDBS, o aumento linear da cmc com a temperatura, deve-se à predominância do segundo fator.[29] Este

comportamento não é padrão para todos os surfactantes, uma vez que depende da natureza molecular do surfactante.

Depois de analisar os princípios básicos da micelização de surfactantes iônicos, na seção seguinte será analisado o processo de formação de complexos de inclusão.

1.3.2. Formação de complexos de inclusão ciclodextrina-surfactante

1.3.2.1. Ciclodextrinas: histórico e estrutura molecular

As ciclodextrinas (CD_s) foram descobertas pelo cientista A. Villiers em 1891 a partir da degradação do amido pelo *Bacillus amylobacter*. Na primeira metade do século XX, foi realizada a caracterização química rigorosa por Franz Schardinger.[30] Até os anos 1970, as CD_s possuíam baixo grau de pureza e a quantidade produzida era limitada. Entretanto, devido ao desenvolvimento da biotecnologia foi possível a clonagem das enzimas glicosiltransferases (CGTases) responsáveis pela degradação do amido, permitindo a produção industrial de CD_s naturais.[31]

As CD_s são compostos macrocíclicos constituídos por unidades de D-glicopiranose ligadas via α(1-4).[32] No processo de síntese, as CD_s com maior rendimento são as αCD, βCD e γCD com, 6, 7 e 8 unidades de glicose, respectivamente (Figura 1.4).

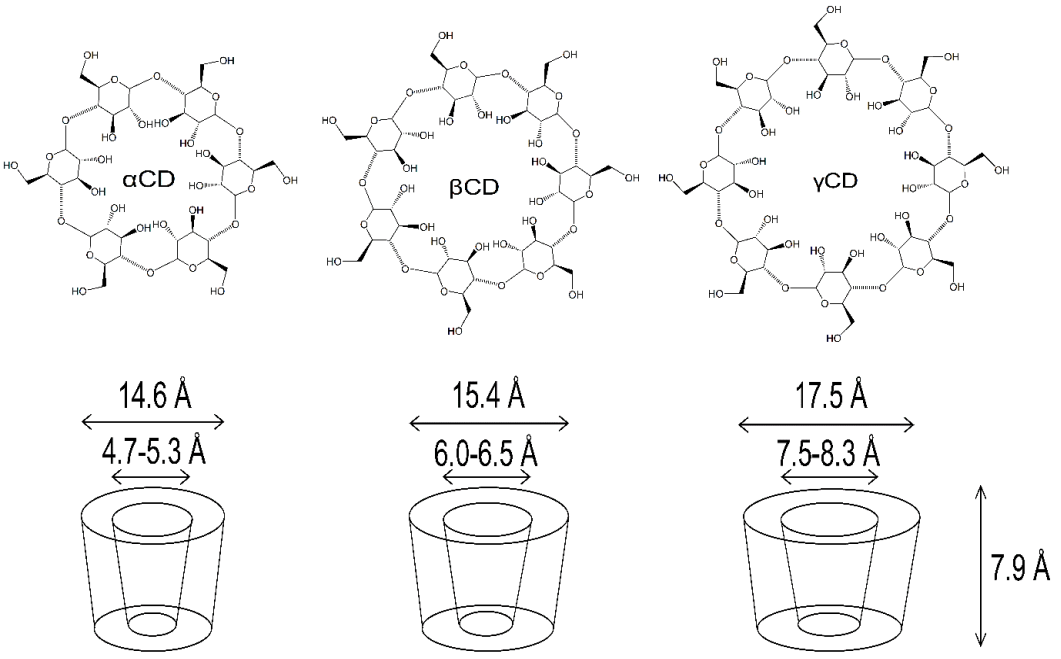

Figura 1.4. Estrutura da α, β e γCD. Adaptado da Figura 1 do trabalho de Venturini et al.[33]

A disposição espacial das unidades de glicose limita a rotação das ligações α(1-4), forçando a CD a assumir uma configuração tronco-cônica, com uma cavidade côncava e oca de 7,9 Å de profundidade. Os diâmetros superior e inferior da cavidade são, 4,7 e 5,3 Å para αCD, 6,0 e 6,5 Å para βCD, 7,5 e 8,3 Å para γCD. A posição dos grupos funcionais hidroxila (-OH), na superfície externa da CD, confere característica hidrofílica, enquanto a ausência de dipolos permanentes entre os átomos que se encontram na superfície da cavidade da CD confere característica hidrofóbica, tornando assim as CD_s compostos anfifílicos. Em seguida será apresentada uma sinopse sobre a termodinâmica da interação surfactante-ciclodextrina em solução aquosa.

1.3.2.2. Termodinâmica da formação de complexos de inclusão ciclodextrina−surfactante ($CD - S$)

Denomina-se complexo de inclusão o agregado molecular formado por duas ou mais moléculas, onde uma possui uma cavidade (hospedeiro) agindo como receptor de uma segunda denominada hóspede.[34] Entre os diversos hospedeiros, a CD é estratégica porque forma complexos com metais,[35] corantes,[36] antibióticos[37] e surfactantes.[38]

Um dos complexos de inclusão que vem chamando a atenção da comunidade científica nos últimos anos é o formado por CD e surfactantes,[39–41] já que a presença do composto macrocíclico em solução altera a atividade superficial dos surfactantes, aumentando a concentração micelar crítica (cmc), propriedade físico-química essencial para sua aplicação.[42– 44]

Para a formação do complexo $CD - S$, é necessário promover a dessolvatação das espécies livres e, dependendo da natureza química do hóspede e hospedeiro, diversas interações não covalentes, como eletrostáticas, de van der Waals, interações hidrofóbicas e ligações de hidrogênio, acontecem.[45]

A formação de um complexo $CD - S$ com estequiometria 1:1 pode ser representado como a seguinte equação química:

$$CD_{(aq)} + S_{(aq)} \rightleftharpoons CD - S_{(aq)} \qquad 1.11$$

onde CD e S são as espécies químicas livres e $CD - S$ é o complexo formado no equilíbrio. A constante de equilíbrio (K_{ic}) do processo de inclusão é representada como:

$$K_{ic} = \frac{a_{CD-S}}{a_{CD}a_S} \quad 1.12$$

a_{CD}, a_S e a_{CD-S} são as atividades correspondentes. Considerando que em baixas concentrações os valores das atividades das espécies em solução são aproximadamente iguais as suas concentrações, a equação 1.12 torna-se:

$$K_{ic} = \frac{[CD-S]}{[CD][S]} \quad 1.13$$

As concentrações das espécies químicas para um experimento de inclusão podem ser obtidas usando diferentes técnicas experimentais, como por exemplo condutividade,[46] espectroscopia de fluorescência,[47] tensão interfacial[48] e calorimetria de titulação isotérmica.[49] O K_{ic} do processo de inclusão $CD-S$ é uma função termodinâmica importante pois permite determinar a variação de energia livre de Gibbs padrão (ΔG_{ic}^{o}) para o processo de associação $CD-S$ usando a equação 1.5. O ΔG_{ic}^{o} é definido como a variação da energia livre Gibbs padrão quando 1 mol de S e CD passam do estado livre em solução para o estado complexado. Esse parâmetro termodinâmico, nas condições padrão, permite caracterizar a predominância do complexo hóspede-hospedeiro ou de suas espécies não complexadas no equilíbrio. Para entender a magnitude do ΔG_{ic}^{o}, é necessário avaliar as contribuições entálpicas e entrópicas dentro do sistema onde o processo de inclusão ocorre.

A variação da entalpia padrão de inclusão (ΔH_{ic}^{o}) pode ser obtida diretamente fazendo uso da calorimetria de titulação isotérmica[50] ou indiretamente utilizando a aproximação de van't Hoff.[51] Para um sistema onde o complexo $CD-S$ é formado, o ΔH_{ic}^{o} pode ser definido como a somatória de 3 processos:

$$\Delta H_{ic}^{o} = \Delta H_{des}^{o,S} + \Delta H_{des}^{o,CD} + \Delta H_{int}^{o,CD-S} \quad 1.14$$

$\Delta H_{des}^{o,S}$ e $\Delta H_{des}^{o,CD}$ são as variações de entalpia padrão da dessolvatação das regiões hidrofóbicas do surfactante e ciclodextrina. $\Delta H_{int}^{o,CD-S}$ é a variação de entalpia padrão da interação entre o surfactante e a ciclodextrina. Apesar de $\Delta H_{des}^{o,S}$ e $\Delta H_{des}^{o,CD}$ serem variações de entalpia associadas à dessolvatação de regiões hidrofóbicas, essas contribuições são

distintas. Na cauda do surfactante as moléculas de água maximizam as ligações de hidrogênio entre elas, sendo necessário fornecer energia para romper as interações água-água da camada de solvatação, tornando o $\Delta H_{des}^{o,S}$ endotérmico.[52] Por outro lado, na cavidade da CD, as moléculas de água possuem ligações água-água menos fortes do que no *bulk* da solução, fazendo com que $\Delta H_{des}^{o,CD}$ seja exotérmico.[53] O $\Delta H_{int}^{o,CD-S}$ é exotérmico devido principalmente à interações de van der Waals e forças dispersivas. O módulo do ΔH_{ic}^{o} será definido pelo balanço entre $\left|\Delta H_{des}^{o,CD} + \Delta H_{int}^{o,CD-S}\right|$ e $\left|\Delta H_{des}^{o,S}\right|$.

Determinando o ΔG_{ic}^{o} e ΔH_{ic}^{o}, a variação da entropia padrão de inclusão ($T\Delta S_{ic}^{o}$) pode ser medida indiretamente, utilizando a equação de Gibbs (equação 1.10). Uma análise paralela ao ΔH_{ic}^{o} pode ser aplicada para a $T\Delta S_{ic}^{o}$:

$$T\Delta S_{ic}^{o} = T\Delta S_{des}^{o,S} + T\Delta S_{des}^{o,CD} + T\Delta S_{tras-rota-conf}^{o,CD-S} \qquad 1.15$$

onde $T\Delta S_{des}^{o,S}$ e $T\Delta S_{des}^{o,CD}$ são as variações de entropia padrão da dessolvatação das regiões hidrofóbicas do surfactante e da ciclodextrina. $T\Delta S_{tras-rota-conf}^{o,CD-S}$ é a variação de entropia padrão da interação entre a ciclodextrina e o surfactante. As moléculas de água solvatando a cauda do surfactante estão altamente estruturadas e, quando o complexo $CD-S$ é formado, atingem a estruturação das moléculas de água presentes no *bulk*, tornando o $T\Delta S_{des}^{o,S}$ positivo.[54] O $T\Delta S_{des}^{o,CD}$ é negativo, uma vez que as moléculas de água dentro da cavidade da CD possuem maior entropia molar do que as moléculas de água no *bulk* da solução (a cavidade da βCD tem espaço para aproximadamente 11 moléculas de água, sendo que apenas 7 moléculas de água estão presentes).[45] O $T\Delta S_{tras-rota-conf}^{o,CD-S}$ é negativo devido ao fato das espécies livres de CD e S possuírem entropia translacional, rotacional e conformacional maior do que em seu estado complexado. Assim o valor do $T\Delta S_{ic}^{o}$ vai depender essencialmente das variações de entropia associadas à dessolvatação das regiões hidrofóbicas.

1.3.2.3. Efeitos sobre a formação de complexos de inclusão surfactante-ciclodextrina como consequência de alterações moleculares dos responsáveis pela associação molecular e da variação da temperatura.

Além dos parâmetros já estudados na seção de micelização para surfactantes iônicos (estrutura molecular do surfactante, natureza do solvente e temperatura), os quais modulam as interações não covalentes do sistema, a estrutura molecular da ciclodextrina (αCD, βCD e γCD) possui um papel fundamental na formação de complexos de inclusão. Nesta breve revisão, são apresentados e discutidos os efeitos gerais de cada um dos parâmetros avaliados.

i) Estrutura molecular da ciclodextrina: Brocos et al.,[55] estudaram a formação de complexos de inclusão entre SDS e ciclodextrinas na temperatura de 298,15 K, utilizando como solvente água pura. O diâmetro interno da ciclodextrina permite formar complexos com diferentes estequiometrias ($n = [SDS]/[CD]$), por exemplo αCD forma complexos com $n = 1$ (22%) e $n = 0{,}5$ (78%), βCD com $n = 1$ (89%) e $n = 0{,}5$ (11%) e γCD com $n = 1$ (~85%) e $n = 2$ (~15%). Estes resultados estão relacionados a dois fatores: *a)* no estado monomérico, a porção hidrofóbica do surfactante tende a estar enovelada para minimizar as interações com as moléculas do solvente e *b)* quanto maior o diâmetro interno da ciclodextrina, maior pode ser a entropia conformacional do hóspede. A formação destes complexos possui valores de ΔH_{ic}^{o} e $T\Delta S_{ic}^{o}$ diferentes. O $\Delta H_{ic,n\,=\,1}^{o}$ e $\Delta H_{ic,n\,=\,0{,}5}^{o}$ são mais negativos para αCD, devido a uma maior eficiência na exclusão das moléculas de água da cavidade do composto macrocíclico e uma maior interação hospede-hospedeiro. O $\Delta H_{ic,n\,=\,2}^{o}$ foi positivo, devido à repulsão eletrostática entre os grupos cabeça do SDS. O $T\Delta S_{ic,n\,=\,1}^{o}$ e $T\Delta S_{ic,n\,=\,0{,}5}^{o}$, foram mais negativos para αCD, uma vez que a interação efetiva entre αCD e SDS diminui a entropia conformacional da cauda do surfactante, forçando assumir a conformação trans. Quando as ciclodextrinas foram βCD e γCD, o $T\Delta S_{ic}^{o}$, foi positivo devido ao ganho entrópico vindo da liberação de uma maior quantidade de moléculas de água com baixa entropia molar, quando a formação dos complexos de inclusão ocorre.

ii) Estrutura molecular do surfactante: entre as diferentes características moleculares que podem ser avaliadas dos surfactantes iônicos para a formação de complexos de inclusão, o tamanho da cauda do detergente é a que tem apresentado efeitos significativos. Benko et al.,[56] avaliaram a inclusão de brometos de N-alquiltrimetilamônio

(C_nTAB, $10 \leq n \leq 16$) em βCD em temperatura de 298,15 K e em água pura. A estequiometria para estes complexos foi 1. O aumento do número de carbonos da cadeia alquílica de 10 para 16 promoveu o aumento da K_{ic} de 5000 para 64000 L mol^{-1}, respectivamente. Isto indicou que ocorre uma maior estabilidade termodinâmica da formação dos complexos com o aumento do número de carbonos. Esta estabilidade foi promovida pelo aumento de n que tornou os valores de ΔH_{ic}^{o} mais exotérmicos de −6,22 para −14,75 kJ mol^{-1}, uma vez que os valores de $T\Delta S_{ic}^{o}$ permaneceram quase constantes. Isto implica que ocorre um maior conteúdo de energia liberada vindo das interações hidrofóbicas e dispersivas entre a cauda do surfactante e cavidade da ciclodextrina.

iii) Efeito da temperatura: Benko et al.,[57] avaliaram o efeito da temperatura (7 temperaturas entre 288 e 348 K) sobre a termodinâmica da formação de complexos de inclusão entre βCD e surfactante (SDS, dodecilsulfonato de sódio (SDSN), brometo de dodeciltrimetilamônio ($C_{12}TAB$) e óxido de dodecil(dimetil)amino (DDAO)) em água pura). O valor da K_{ic} foi similar para todos os complexos, sugerindo uma mínima contribuição do grupo cabeça e do contra-íon dos monômeros do surfactante. O aumento da temperatura até 348 K não alterou a estequiometria do complexo formado ($n \approx 1$) e gradualmente diminuiu o valor do K_{ic} até atingir o 30% do valor obtido em 288 K para todos os surfactantes. Esta variação do K_{ic} promoveu mínimas alterações no valor de ΔG_{ic}^{o} (em torno de 1,5 kJ mol^{-1}). Entretanto os valores de ΔH_{ic}^{o} e $T\Delta S_{ic}^{o}$ diminuíram linearmente com o aumento da temperatura. Como já foi mencionado, o processo de formação de complexos de inclusão está associado a 3 processos que ocorrem simultaneamente; dessolvatação do surfactante ($\Delta H_{des}^{o,S} > 0$; $T\Delta S_{des}^{o,S} > 0$); dessolvatação da cavidade da ciclodextrina ($\Delta H_{des}^{o,CD} < 0$; $T\Delta S_{des}^{o,CD} < 0$) e interação surfactante-ciclodextrina ($\Delta H_{int}^{o,CD-S} < 0$; $T\Delta S_{tras-rota-conf}^{o,CD-S} < 0$). Assim, o trabalho de Benko indica que para baixas temperaturas (288 K) o processo é entropicamente dirigido, principalmente pela liberação das moléculas de água da cauda do surfactante para o *bulk* da solução. Com o aumento da temperatura (289-327 K) o papel dos outros dois processos aumenta, tornando o processo de inclusão entropicamente e entalpicamente dirigido. Já para temperaturas elevadas (329-348 K) o processo é somente entalpicamente dirigido, devido à predominância da contribuição energética vinda da interação entre a cauda do surfactante e a cavidade da ciclodextrina.

iv) Efeito da polaridade do solvente: as características moleculares do solvente afetam a termodinâmica de formação de complexos de inclusão. Entretanto, são poucos

os trabalhos que estudam o efeito de cossolutos e cossolventes sobre a formação de complexos supramoleculares envolvendo surfactantes e compostos macrocíclicos. Cossolventes como isopropanol,[58] butanol[59] e N-acetamida[60] tem sido estudados em diferentes proporções sobre a termodinâmica de inclusão de alquiltrimetilamônio em βCD. O aumento da concentração de cada um destes cossolventes diminui a K_{ic}, alterando a estabilidade termodinâmica dos complexos em solução. As conclusões sobre este comportamento estão relacionadas com o aumento da estabilidade dos monômeros livres de surfactante em presença dos compostos orgânicos, promovendo um enfraquecimento nas interações surfactante-ciclodextrina. Em contraste ao efeito dos cossolventes, Czapkiewicz et al.,[61] reportaram o efeito do cossoluto NaCl sobre a inclusão de C_{12}PyCl na βCD. A presença de 100 mmol L^{-1} de NaCl promoveu uma maior estabilidade do complexo formado, como evidenciado pelo aumento da constante de inclusão, devido ao efeito cosmotrópico do sal.

Depois de analisar os processos de micelização de surfactantes e formação de complexos de inclusão $CD - S$ em água pura, na seção seguinte será apresentado o que são líquidos iônicos e seu efeito sobre a micelização de surfactantes modelo.

1.3.3. Líquidos iônicos

1.3.3.1. Propriedades moleculares

Os líquidos iônicos (LIs) são sais orgânicos ou semi-orgânicos em estado físico líquido à temperaturas abaixo de 373,15 K.[62] Embora o primeiro LI sintetizado foi o nitrato de etilamônio ($C_2NH_8NO_3$) na segunda década do século XX, somente a partir de 1980 foi possível sintetizar um LI estável à umidade do ar, sendo estes baseados no cátion imidazólio (Figura 1.5).[63]

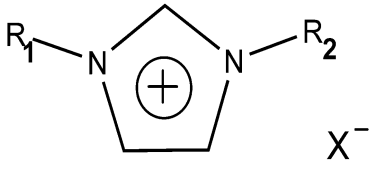

Figura 1.5. Estrutura molecular de um clássico líquido iônico com cátion imidazólio. R_1 e R_2 são cadeias alquílicas com comprimento de 0 até n carbonos. X^- é o ânion do LI. Uma grande variedade pode ser encontrada no trabalho de Kohno et al.[64]

O cátion imidazólio é formado por um anel aromático com carga deslocalizada e cadeias alquílicas (R_1 e R_2) ligadas a átomos de nitrogênio que modulam seu caráter hidrofóbico. Os LIs com cátion imidazólio possuem propriedades como alta estabilidade térmica e química, alta solubilidade em água, baixa pressão de vapor e baixa toxicidade, tornando-os compostos químicos com alto potencial para substituir os tradicionais solventes orgânicos, os quais são tóxicos e utilizados em grande extensão em processos industriais.[65,66]

1.3.3.2. Efeito de líquidos iônicos sobre processos de agregação molecular

Apesar dos LIs imidazólicos já serem reportados na literatura há quase 40 anos, só no início da primeira década do século XXI foi iniciado seu estudo em sistemas coloidais, onde ocorre agregação molecular.[67] O efeito de LIs sobre processos de associação tem sido analisado usando o LI como solvente[68–70] ou misturas de água e LI,[71–73] sendo essa última estratégica para monitorar o papel das moléculas de água nas camadas de solvatação das regiões hidrofóbicas dos compostos anfifílicos e presentes no *bulk* da solução.

O efeito de misturas aquosas formadas por eletrólitos comuns e água,[74–77] ou compostos orgânicos e água[78–81] sobre processos de agregação molecular, são bem diferentes dos efeitos promovidos por soluções aquosas formadas por LI. Isto ocorre devido à estrutura molecular estratégica dos líquidos iônicos que permite monitorar interações iônicas (não possível em cosolutos orgânicos clássicos) e interações hidrofóbicas (não possível em eletrólitos simples) simultaneamente.

O efeito de soluções aquosas de LI já tem sido amplamente estudado na formação de micelas de surfactantes iônicos, enquanto que ainda não existem estudos do efeito destes solventes sobre a formação de complexos de inclusão entre surfactantes e ciclodextrinas.

Serão focados trabalhos e análises do efeito de líquidos iônicos imidazólicos em solução aquosa sobre a micelização do dodecilsulfato de sódio (SDS) em temperatura de 298,15 K, surfactante modelo com mais reportes na literatura científica, e que apresenta alta similaridade com o surfactante SDBS, o qual é o alvo de estudo no capitulo 3.

Behera et al.,[82,83] estudaram o efeito de C_4mimBF_4 e C_4mimPF_6 na micelização do SDS em solução aquosa. Em concentrações baixas de LI (≤ 2% m/m), foi observada uma diminuição da cmc do SDS, entretanto quando a concentração do LI aumentou até

30% m/m, os valores da cmc permaneceram constantes e aumentaram para o C_4mimPF_6 e C_4mimBF_4, respectivamente. Para concentrações baixas de LI, estes se comportam como eletrólitos simples, e em altas concentrações os cátions imidazólio migram do *bulk* da solução para fazer parte das micelas de SDS, como evidenciado pelo crescimento do agregado micelar a partir de medidas do número de agregação.

Ferreira et al,[84] realizaram medidas de calorimetria de titulação isotérmica e condutividade elétrica para determinar o efeito de haletos de C_4mimX (X = Cl^- ou Br^-) sobre a micelização do SDS. O aumento da concentração de LI até 10 mmol L^{-1} diminuiu exponencialmente o valor da cmc, sendo o efeito mais pronunciado na presença do LI com ânion Cl^-. Seus resultados apontaram que a cadeia alquílica do cátion C_4mim^+ interage com o núcleo hidrofóbico da micela e o anel imidazólio com os grupos cabeça do SDS.

Smirnova et al.,[85] compararam o efeito do eletrólito simples NaCl e os LIs cloreto de 1-alquil-3-metilimidazólio no processo de agregação do SDS, usando as mesmas técnicas do trabalho de Ferreira et al. Eles determinaram que o aumento da concentração e hidrofobicidade de LI no sistema promovem uma diminuição da cmc e torna entalpicamente mais favorável o processo de micelização, como observado pelas medidas de ΔH^o_{mic}. Os resultados obtidos destes parâmetros foram comparados com os obtidos pelo NaCl, e foi observada uma maior diminuição destes parâmetros na presença dos LIs, devido a interações eletrostáticas e hidrofóbicas do cation imidazólio agindo simultaneamente em solução, as quais promovem a solvatação preferencial da cauda do surfactante.

Depois da compilação da informação sobre a complexação entre surfactantes e ciclodextrinas, micelização de surfactantes e o efeito de LIs na agregação de compostos anfifílicos, serão apresentados os métodos experimentais usados nesta tese.

1.4. *Métodos e ferramentas instrumentais para estudar processos de associação molecular*

Para os químicos experimentais, são essenciais quatro etapas ao longo da produção de um resultado de qualidade: *i*) revisão da literatura; *ii*) rigor no preparo das soluções que contém os sistemas de estudo; *iii*) domínio sobre as técnicas instrumentais usadas para monitorar as propriedades físico-químicas do fenômeno avaliado e *iv*) tratamento dos dados obtidos a partir das medidas experimentais.

Esta seção está destinada ao detalhamento das ferramentas instrumentais usadas para o monitoramento da formação de micelas e de complexos de inclusão, assim como os métodos utilizados para o tratamento dos dados.

1.4.1. Calorimetria de titulação isotérmica (ITC)

A calorimetria de titulação isotérmica (ITC - sigla proveniente do inglês *Isothermal Titration Calorimetry*), é a única técnica instrumental capaz de monitorar diretamente a variação de entalpia de processos de associação molecular, tais como micelização de surfactantes,[86] formação de complexos de inclusão[87] e interação polímero-surfactante.[88]

A Figura 1.6 apresenta o esquema básico de um calorímetro de titulação isotérmica.

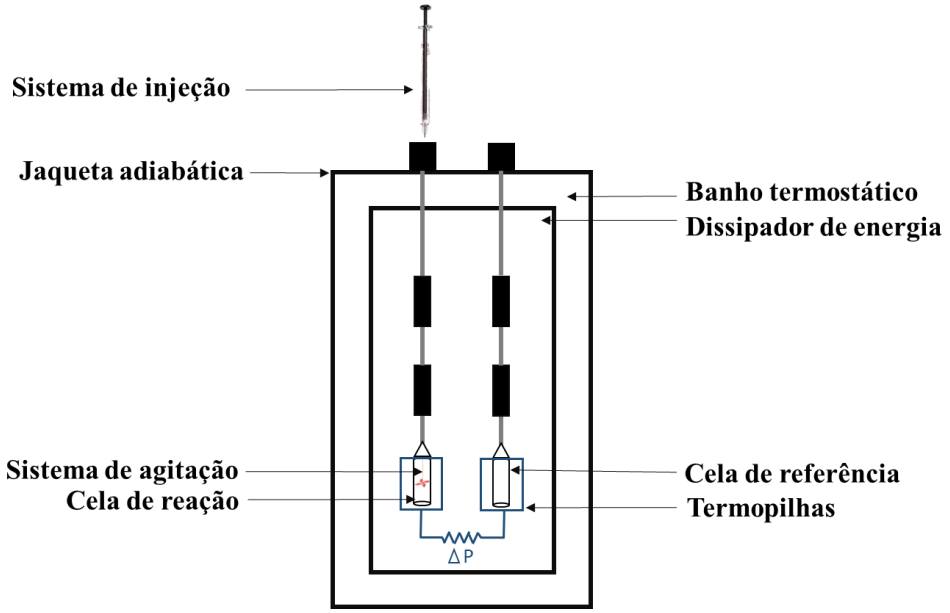

Figura 1.6. Diagrama esquemático dos principais componentes de um calorímetro de titulação isotérmica.

Para um experimento de micelização de surfactante, volumes (5-10 μL) de uma solução aquosa de surfactante (com concentração de 10 a 20 vezes o valor da cmc), controlados pelo sistema de injeção, são titulados sobre um volume específico do solvente (2,7 mL), presente na cela de reação à temperatura e pressão constante. Nesse processo de injeção, uma nova mistura é formada na cela de reação, onde interações

intermoleculares são rompidas e formadas, promovendo uma variação de temperatura (ΔT). Além da cela de reação, o calorímetro possui uma segunda cela (com 2,7 mL do solvente) que atua como referência do estado inicial das interações intermoleculares presentes no solvente. Para atingir o equilíbrio térmico (se o processo for exotérmico), uma quantidade de energia na forma de calor (q) flui entre a cela de reação e a vizinhança, detectada por dispositivos eletrônicos chamados termopilhas. Quando q flui através das termopilhas, um diferencial de potencial elétrico surge proporcional ao ΔT. O sinal reportado pelo equipamento é denominado termograma, um gráfico que mostra a variação de potência (ΔP) em função do tempo (t) (Figura 1.7).

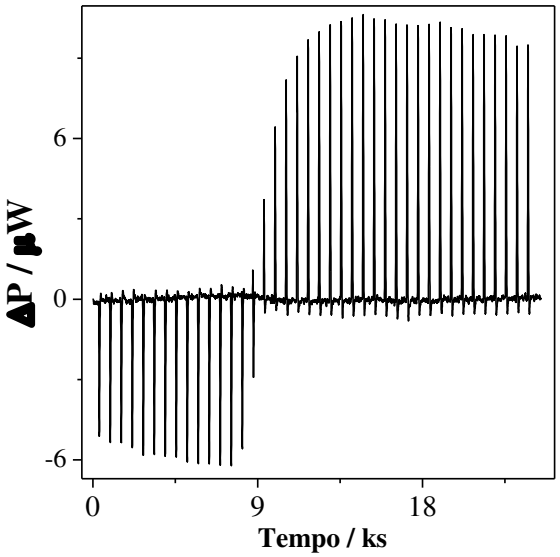

Figura 1.7. Termograma obtido em um monitor de atividade térmica (TAM III-ITC) a partir da titulação de uma solução de dodecilsulfato de sódio (SDS) 50 mmol L^{-1} preparada em solução aquosa de 2 mmol L^{-1} de brometo de 1-butil-3-metilimidazólio (C_4mimBr), sobre uma solução de 2 mmol L^{-1} de C_4mimBr. As injeções do titulante foram de 5 μL e o volume inicial do titulado na cela de reação foi de 2,7 mL. Adaptado da Figura 3 do trabalho de Ferreira et al.[84]

Cada uma das injeções do titulante gera um pico onde a área pode ser integrada como apresentado pela equação 1.16:

$$q = \int_{t_1}^{t_2} P dt \qquad 1.16$$

onde q é a energia na forma de calor de cada injeção no intervalo de tempo t_1 e t_2. Como o experimento dentro do calorímetro é realizado a pressão constante, q corresponde a

variação de entalpia observada (ΔH_{obs}) do processo termodinâmico decorrente da nova mistura formada entre o titulante e o titulado, a seguir será apresentada a dedução.

A primeira lei da termodinâmica para o processo de titulação define:

$$dU = dq + dw \tag{1.17}$$

em que U é a energia interna do sistema, q e w são as quantidades energia, na forma de calor e trabalho, trocadas entre o sistema e a vizinhança. Para um processo termodinâmico com estado inicial e final definido, à pressão constante (P), temos que:

$$\int_i^f dU = \int_i^f dq + P\int_i^f dV \tag{1.18}$$

Resolvendo as integrais e agrupando os termos, a equação 1.19 é obtida:

$$q = \left(U_f + PV_f\right) - (U_i + PV_i) \tag{1.19}$$

Em termodinâmica a soma $U + PdV$ denomina-se entalpia (H), assim a equação 1.19 torna-se:

$$\Delta H = H_f - H_i = q \tag{1.20}$$

Deste modo, temos que a variação de energia na forma de calor, transferida do sistema para a vizinhança, é numericamente igual à variação da entalpia do processo ocorrendo na nova mistura formada. A seguir será apresentado o uso da ITC em sistemas onde ocorre a formação de micelas de surfactantes iônicos.

1.4.1.1. Uso da calorimetria de titulação isotérmica no estudo da formação de micelas de surfactantes iônicos

A caracterização termodinâmica completa (ΔG^o_{mic}, ΔH^o_{mic} e $T\Delta S^o_{mic}$) por uma única técnica instrumental, para o processo de micelização de surfactante iônicos, só é possível utilizando condutividade elétrica. A grande desvantagem de utilizar o método condutimétrico, é o uso da aproximação de van't Hoff para calcular o ΔH^o_{mic}, o qual torna

as mudanças translacionais, conformacionais e rotacionais das espécies químicas em solução, além dos efeitos do solvente, independentes da temperatura do experimento.[89]

Apesar da ITC não permitir a determinação de todos os parâmetros termodinâmicos de micelização, por falta do α, parâmetros importantes como cmc e ΔH^o_{mic} podem ser calculados.[90]

Para entrarmos no uso da ITC no processo de micelização, cabe destacar que a parte experimental foi explicada na seção 1.4.1., permitindo que nesta seção entremos apenas nas análises dos resultados. A Figura 1.8 apresenta a variação de entalpia observada (ΔH_{obs}) em função da concentração de dodecilsulfato de sódio (SDS), à temperatura de 298,15 K.

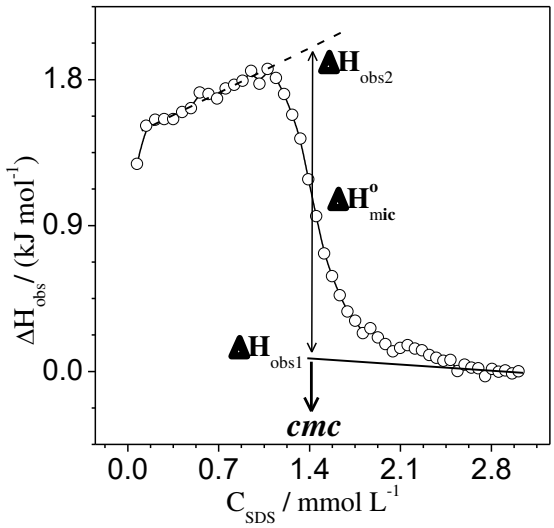

Figura 1.8. ΔH_{obs} *versus* [SDS] para a micelização de SDS em tampão pH 7,4 à temperatura de 298,15 K. (−,--) Ajustes lineares para obter o parâmetro ΔH^o_{mic}. Adaptado da Figura 6b do trabalho de Rengifo et al.[91]

O formato das curvas ΔH_{obs}x[surfactante] depende de fatores como solvatação-dessolvatação dos monômeros de surfactante, número de agregação, temperatura e formato do agregado.[92] Em geral, as curvas de micelização obtidas por ITC podem ser divididas em 3 regiões de concentração: *i*) região pré-micelar (≈ 0-1,1 mmol L^{-1} de SDS), onde os valores de ΔH_{obs} são relacionados com a diluição e quebra das micelas, solvatação e diluição dos monômeros de surfactante; *ii*) região de transição (≈ 1,1-2,1 mmol L^{-1} de SDS) na qual a cmc pode ser obtida e *iii*) região pós-micelar (≈ 2,1-2,9 mmol L^{-1} de SDS), onde os valores de ΔH_{obs} referem-se à diluição das micelas.

A cmc é determinada no ponto de inflexão da curva utilizando o método da primeira derivada ($\partial(\Delta H_{obs})/\partial([SDS])$),[93] sendo o valor reportado por Rengifo et al.[91]

para o SDS em tampão pH 7,4, 1,33 mmol L^{-1}. Em ordem de obter o ΔH^o_{mic}, é necessário extrapolar a região pré-micelar (--) e pós-micelar (−) até o valor obtido de cmc. Com esse procedimento dois valores de variação de entalpia observada são obtidos, ΔH_{obs2} e ΔH_{obs1}, e a diferença entre eles permite calcular o valor ΔH^o_{mic}:

$$\Delta H^o_{mic} = \Delta H_{obs1} - \Delta H_{obs2} \quad (1.21)$$

O processo de micelização do SDS em tampão pH 7,4, é entalpicamente dirigido, uma vez que as interações monômero-monômero apresentam valores de ΔH_{obs} mais positivos que os obtidos da interação micela-micela.

1.4.1.2. Uso da calorimetria de titulação isotérmica no estudo da formação de complexos ciclodextrina−surfactante ($CD - S$)

A calorimetria de titulação isotérmica permite realizar uma caracterização termodinâmica completa (ΔG^o_{ic}, ΔH^o_{ic} e $T\Delta S^o_{ic}$) da formação do complexo $CD - S$ em um único experimento. Um conceito básico da formação de complexos de inclusão para o tratamento de dados de ITC é o parâmetro de ligação ($\bar{v}$). O $\bar{v}$ representa o número de mols ligados de S por cada mol de CD, no sistema. O valor numérico de $\bar{v}$ pode variar entre 0 e o número de sítios de ligação. O $\bar{v}$ pode ser definido matematicamente como:

$$\bar{v} = \frac{[S]_L}{[CD]_T} = \frac{[CD - S]}{[CD] + [CD - S]} = \frac{K_{ic}[CD][S]}{[CD] + K_{ic}[CD][S]} = \frac{K_{ic}[S]}{1 + K_{ic}[S]} \quad (1.22)$$

sendo para n sítios de ligação:

$$\bar{v} = n\frac{K_{ic}[S]}{1 + K_{ic}[S]} \quad (1.23)$$

Para cada injeção no experimento de calorimetria, a energia na forma de calor absorvida ou liberada (q_i) é relacionada a um processo de ligação como:

$$q_i = \Delta H_i \, \Delta(mols\ S_i) \quad (1.24)$$

onde ΔH é a variação de entalpia por mol de S ligado e $\Delta(mols\ S_i)$ é a quantidade ligada de S na injeção i. O número de mols de S pode ser representado em função da concentração, podendo reescrever a equação 1.24 como:

$$q_i = \Delta H\ V_c\left([S]_{,i} - [S]_{i,i-1}\right) = \Delta H\ V_c(\bar{v}_i[CD]_{T,i} - \bar{v}_{i-1}[CD]_{T,i-1}) \quad 1.25$$

em que V_c representa o volume efetivo da cela calorimétrica e $[CD]_T$ é a concentração total de CD na cela de reação para a injeção i. Desta forma, a energia na forma de calor total (Q), liberada ou absorvida depois de N injeções pode ser descrita como:

$$Q = \sum_{i=1}^{N} q_i = \Delta H\ V_c\ [CD]_T\bar{v} = \Delta H\ V_c\ [CD]_T n\frac{K_{ic}[S]}{1 + K_{ic}[S]} \quad 1.26$$

Ao longo de um experimento de ITC, a $[S]$ não ligado é uma variável desconhecida, sendo necessário operacionalmente introduzir na equação 1.26 uma variável conhecida, como a quantidade de surfactante total.

Levando em consideração que:

$$[S]_T = [S] + [CD - S]\ \rightarrow [S] = [S]_T - [CD - S] \quad 1.27$$

e

$$n[CD]_T = [CD] + [CD - S]\ \rightarrow [CD] = n[CD]_T - [CD - S] \quad 1.28$$

Assim podemos utilizar a expressão matemática de K_{ic} (equação 1.13) e colocar os termos em função das concentrações totais de cada uma das espécies e da concentração do complexo formado (equação 1.29).

$$K_{ic}([S]_T - [CD - S])(n[CD]_T - [CD - S]) = [CD - S] \quad 1.29$$

multiplicando e igualando a zero, temos que:

$$K_{ic}[S]_T n[CD]_T - K_{ic}n[CD]_T[CD - S] - [CD - S] - K_{ic}[S]_T[CD - S] + K_{ic}[CD - S]^2 = 0 \quad 1.30$$

reordenando e agrupando os termos, a equação 1.30 fica:

$$[CD - S]^2 + \left(-n[CD]_T - [S]_T - {K_{ic}}^{-1}\right)[CD - S] + [S]_T n[CD]_T = 0 \qquad 1.31$$

A equação 1.31 é uma equação quadrática, que pode ser simplificada em:

$$[CD - S]^2 + b[CD - S] + c = 0 \qquad 1.32$$

onde $b = -n[CD]_T - [S]_T - {K_{ic}}^{-1}$ e $c = [S]_T n[CD]_T$. Solucionando a equação 1.32 é obtido:

$$[CD - S] = \frac{-b - \sqrt{b^2 - 4c}}{2} \qquad 1.33$$

substituindo a equação 1.33 na equação 1.26 obtém-se:[94]

$$Q = \frac{\Delta H_{ic}^o \, V_c}{2K_{ic}}\left[1 + K_{ic}[S]_T + nK_{ic}[CD]_T - \sqrt{(1 + K_{ic}[S]_T + nK_{ic}[CD]_T)^2 - 4n{K_{ic}}^2[CD]_T[S]_T}\right] \qquad 1.34$$

A equação 1.34 é uma função sigmóide que permite calcular os parâmetros de inclusão (K_{ic}, n e ΔH_{ic}^o) para a ligação de um ligante a uma macromolécula com n sítios idênticos e independentes. É de destaque que K_{ic} é igual à constante de Wiseman (K_W) somente quando n = 1, relatando a não cooperatividade deste processo.[95]

Para exemplificar a aplicação do modelo da equação 1.34, na Figura 1.9 é apresentada a energia na forma de calor associada à ligação de um ligante a uma macromolécula em função da razão molar.

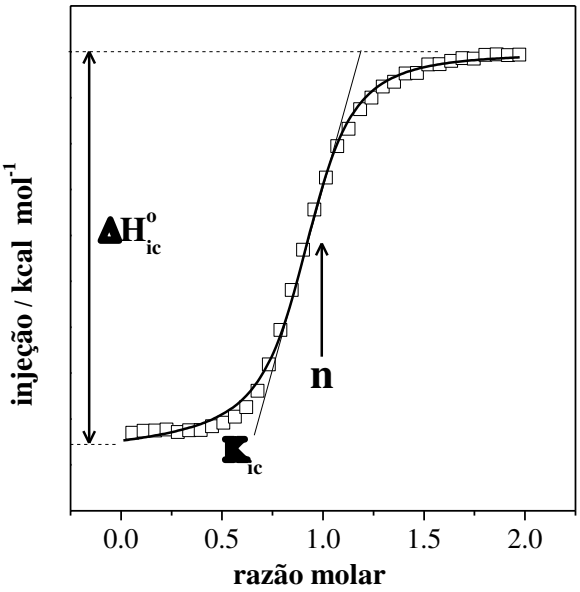

Figura 1.9. Energia na forma de calor da interação hospede-hospedeiro *versus* razão molar. (−) Ajuste para obter parâmetros como n, ΔH_{ic}^{o} e K_{ic}. Adaptado da Figura 1 do trabalho de Callies et al.[96]

A variação da entalpia padrão de inclusão (ΔH_{ic}^{o}) é obtida da diferença entre os dois patamares do perfil sigmoide. A razão molar na mudança de concavidade da função (ponto de inflexão) representa a estequiometria do complexo formado (n). A tangente que passa pelo centro do ajuste sigmoide permite determinar a constante do processo de inclusão (K_{ic}), utilizada para o cálculo da variação da energia livre de Gibbs padrão de inclusão. Calculando a ΔH_{ic}^{o} e ΔG_{ic}^{o} é possível calcular a variação de entropia padrão de inclusão ($T\Delta S_{ic}^{o}$).

1.4.2. Condutividade elétrica

Para obter a concentração micelar crítica (cmc) de surfactantes iônicos e o grau de dissociação micelar (α), é necessário utilizar uma técnica altamente sensível à variação da mobilidade de íons em solução, a condutividade elétrica (k).[28]

A condutividade elétrica (k) é uma grandeza relacionada à habilidade de uma solução para conduzir uma corrente elétrica. A condutividade elétrica depende do número e mobilidade dos íons e partículas presentes na solução. Em sistemas coloidais de surfactantes iônicos,[84] a k aumenta linearmente com o aumento da concentração do surfactante $[S]$. Para concentrações menores à cmc, o valor da inclinação, $\partial(k)/\partial([S])$, é diferente do que para concentrações acima à cmc, possibilitando diferenciar as regiões

relacionadas à interação monômero-monômero e micela-micela. A mudança de inclinação $\partial(k)/\partial([S])$ para as duas regiões vem da mobilidade de um menor número de íons em solução, devido ao confinamento de uma quantidade de monômeros de surfactante que saem do *bulk* da solução para formar o agregado micelar. A cmc é determinada no intercepto entre o comportamento linear da interação monômero-monômero e micela-micela (equação 1.35).

$$a_1[cmc] + b_1 = a_2[cmc] + b_2 \qquad 1.35$$

a_i é o coeficiente angular e b_i o intercepto em $[S] = 0$. α é obtido da razão entre a_2 e a_1. Este parâmetro expressa a quantidade de contra íons do surfactante que ficam livres em solução. Usando a cmc e o α é possível determinar o potencial termodinâmico do processo de micelização usando a equação 1.8.

1.4.3. Espectroscopia de fluorescência

Uma ferramenta instrumental clássica para a determinar a concentração da formação agregados de surfactantes em solução aquosa é a espectroscopia de fluorescência.[97,98]

A fluorescência é definida pela união internacional de química pura e aplicada (IUPAC-por suas siglas em inglês), como a luminescência que ocorre essencialmente somente durante a irradiação de uma substancia por radiação eletromagnética. A Figura 1.10 apresenta o diagrama de Jablonski, que representa os principais processos físicos que ocorrem na excitação-emissão de uma espécie química com propriedades fluorescentes.

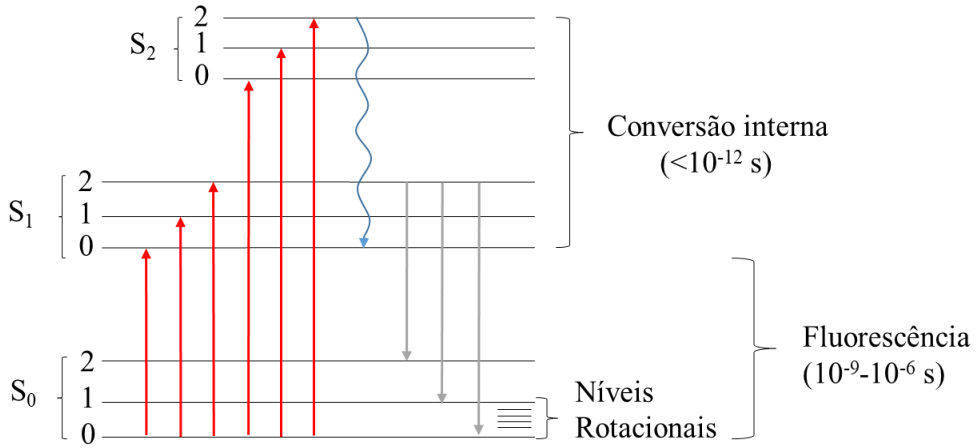

Figura 1.10. Diagrama de Jablonski.[99]

Na Figura 1.10, as notações utilizadas são S para os estados eletrônicos singletes do estado fundamental (S0), primeiro (S1) e segundo (S2) excitado. Cada um dos S, possui subníveis vibracionais representados com números inteiros 1, 2 e 3, onde cada subnível tem subníveis rotacionais.

Num experimento típico de fluorescência, à temperatura ambiente (298,15 K), a maioria das moléculas presentes dentro do sistema com propriedades fluorescentes estão ocupando o mais baixo nível vibracional do estado eletrônico fundamental (S0), no entanto com a absorção de radiação eletromagnética elas são levadas para níveis roto-vibrônicos excitados em S1 ou S2. Depois do processo de excitação, em um tempo < 10^{-12} segundos, as moléculas relaxam até atingir o menor subnível vibracional em S1 (processo de conversão interna). Em S1, ocorre uma transição da molécula para um dos subníveis vibracionais de S0, emitindo um fóton num tempo entre 10^{-9} e 10^{-6} segundos. A emissão do fóton ocorre em comprimento de onda (λ_{em}) maior que o comprimento de onda da radiação eletromagnética que entrou em contato com as moléculas (λ_{ex}), devido à perda de energia gerada pelas colisões entre as moléculas no estado excitado.

O fenômeno de fluorescência é monitorado em um equipamento chamado de espectrofluorímetro. Quando uma solução com propriedades fluorescentes é colocada na cela de amostra, vários parâmetros do equipamento precisam ser ajustados, tais como o λ_{ex} e a faixa de λ_{em} da espécie química em análise, voltagem da fotomultiplicadora, aberturas das fendas de emissão e excitação, tempo de leitura e temperatura da medida, sendo cada uma das anteriores dependente da natureza e concentração do fluoróforo estudado.

O resultado obtido na medida é um espectro de emissão de fluorescência, que pode apresentar uma ou mais bandas associadas às diferentes transições dos fótons fluorescentes. As mudanças de intensidade das bandas obtidas num espectro são dependentes de fatores como temperatura,[100] concentração do fluoróforo[101] e polaridade do solvente.[102]

1.4.3.1. Uso da espectroscopia de fluorescência para a determinação da *cmc* de surfactantes iônicos

Os experimentos de fluorescência são baseados na medida da fluorescência intrínseca de espécies químicas, como por exemplo proteínas, ao estudar a interação proteína-analito por meio do monitoramento da intensidade de fluorescência de resíduos de triptofano,[103] ou a fluorescência extrínseca de moléculas em solução, como por exemplo sondas moleculares para determinação da *cmc* de surfactantes.[104]

A sonda fluorescente mais usada para a determinação do parâmetro de agregação molecular de surfactantes é o pireno, um hidrocarboneto policíclico aromático de baixa solubilidade em água. A Figura 1.11 apresenta o espectro de emissão do pireno a 298,15 K.[105]

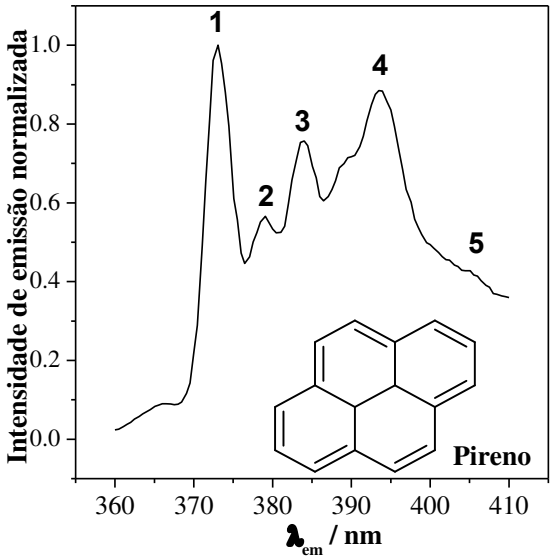

Figura 1.11. Espectro de emissão do pireno quando excitado em 335 nm. Internamente é apresentada a estrutura molecular do pireno.

Quando o pireno é excitado em 335 nm, é obtido um espectro de emissão de fluorescência com 5 bandas vibrônicas entre 360 nm e 410 nm. Sendo que para a determinação da cmc são importantes a primeira ($I_1 = 373\,nm$) e a terceira ($I_3 = 384\,nm$) banda, já que a razão entre estas (I_1/I_3) é um parâmetro sensível para avaliar mudanças na polaridade do microambiente onde está localizado o pireno em solução. Para um solvente com alta polaridade como água ou misturas água e eletrólitos simples, a razão das bandas é próxima de 1,8, enquanto que para solventes com baixa polaridade como hexano, essa razão é aproximadamente 0,6.[106]

Depois de discutir as propriedades de fluorescência do pireno, será apresentada uma performance de um experimento para determinar a cmc do surfactante SDS usando como solvente misturas de água + C_4mimBr. O experimento consiste na titulação de ≈ 15 injeções de uma solução concentrada de SDS sobre um volume definido de uma solução aquosa de pireno 1,2 μmol L^{-1} (concentração para evitar a formação de excímeros) presente na cela do fluorímetro, ambas as soluções foram preparadas nas misturas água + C_4mimBr. A cada injeção é coletado o espectro do pireno e realizado o cálculo da razão das bandas. A Figura 1.12 apresenta a razão I_1/I_3 em função do log([SDS]) a 298,15 K.

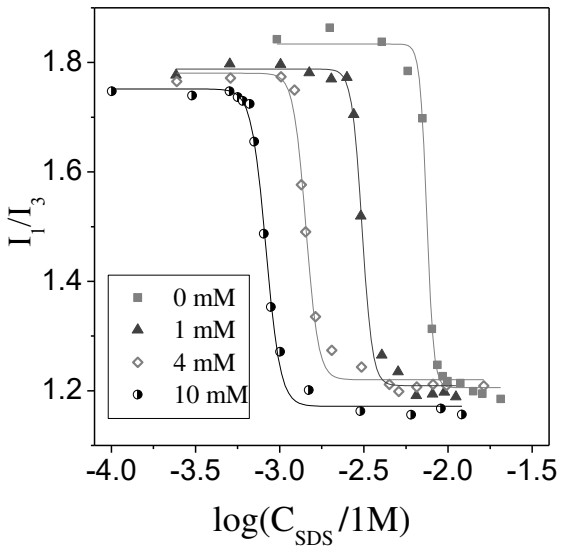

Figura 1.12. Coeficiente da intensidade das bandas vibracionais (I_1/I_3) do espectro de emissão do pireno *versus* log[SDS] em diferentes concentrações de C_4mimBr a 298,15 K. Adaptado da Figura 2 do trabalho de Ferreira et al.[84]

Para o experimento realizado em água pura, em baixas concentrações de SDS ($\approx -2{,}9 \leq \log[\text{SDS}] \leq -2{,}1$), a razão I_1/I_3 permanece constante com valor próximo a 1,8, o qual caracteriza as moléculas de pireno em um ambiente de alta polaridade. Acima

de ≈ log[SDS] = −2,1, a razão I_1/I_3 inicia um abrupto decrescimento, indicando a redução da polaridade no microambiente ao redor das moléculas de pireno. Essa diminuição da polaridade foi caracterizada pela partição gradual do policíclico aromático do *bulk* da solução para o núcleo das primeiras micelas formadas de SDS. Quando o valor de log[SDS] atinge ≈ −2,0, a razão I_1/I_3 permanece constante, indicando a partição quase total das moléculas de pireno para o núcleo das micelas.

A determinação da cmc é realizada utilizando o ajuste sigmoide de Boltzmann (–);[107]

$$\frac{I_1}{I_3} = \frac{A_1 - A_2}{1 + e^{\frac{(x-x_0)}{\Delta x}}} \qquad 1.36$$

A_1 e A_2 representam os limites superior (interação monômero-monômero) e inferior (interação micela-micela) do ajuste sigmoide, respectivamente. O x_0 (cmc = 7,47 ± 0.01 mmol L^{-1}) é o ponto de inflexão. Na cmc, a probabilidade de um monômero de surfactante adicionado permanecer no *bulk* da solução ou dentro de uma micela é de 50%.

1.4.4. Espalhamento dinâmico de luz (DLS)

O alto potencial tecnológico dos surfactantes depende da dinâmica das micelas em solução, da interação agregado-agregado e da transferência de substâncias químicas entre a fase dispersa e contínua. Estas características dependem do raio hidrodinâmico (R_H) dos agregados poli-moleculares e a técnica por excelência para determiná-lo é o espalhamento dinâmico de luz (DLS - sigla proveniente do inglês *Dynamic Light Scattering*).[108]

O DLS baseia-se em dois fenômenos que ocorrem simultaneamente quando um feixe de um laser incide sobre um sistema micelar, sendo eles o movimento Browniano e o efeito Tyndall. O movimento Browniano refere-se à mobilidade aleatória dos nanoagregados dentro de um sistema coloidal, enquanto que o efeito Tyndall descreve a re-emissão da radiação incidente pela existência de entes espalhadores dentro do sistema.[108]

O procedimento teórico-experimental para determinar o R_H das micelas de surfactante usando DLS consiste em 4 passos: *i*) avaliar a função de correlação de segunda

ordem ($g^{(2)}_{(\tau)}$) com o tempo (t), *ii*) calcular o vetor espalhamento (q_{es}), *iii*) determinar o coeficiente de difusão (D) e *iv*) usar a relação de Stokes-Einstein.[109]

i) *Função de correlação de segunda ordem* ($g^{(2)}_{(\tau)}$): a flutuação da intensidade de luz espalhada está relacionada com o tamanho e dinâmica das micelas espalhadoras. A ferramenta instrumental encarregada de medir estas flutuações é a espectroscopia de correlação de fótons. Esta técnica permite monitorar de forma quantitativa a luz espalhada em função do tempo, e os dados experimentais são processados fazendo uso da função de correlação da intensidade espalhada ($g^{(2)}_{(\tau)}$);

$$g^{(2)}_{(\tau)} = \frac{\langle I_{(t)} I_{(t+\tau)} \rangle}{\langle I \rangle^2} \qquad 1.37$$

onde τ e $I_{(t)}$, representam o tempo de correlação e intensidade no tempo *t*, respectivamente. As funções de correlação da intensidade espalhada e do campo elétrico espalhado se relacionam pela relação de Siegert;[109]

$$g^{(2)}_{(t)} = 1 + \beta \left| g^{(1)}_{(\tau)} \right|^2 \qquad 1.38$$

Para dispersões coloidais onde os espalhadores têm o mesmo tamanho e têm sua difusão livre, condição normalmente alcançada para dispersões diluídas, tem-se que a função de correlação da intensidade decai exponencialmente no tempo, da seguinte forma:

$$g^{(2)}_{(t)} = 1 + \beta e^{(-2Dq^2_{es}t)} \qquad 1.39$$

Onde β é um parâmetro fator geométrico relacionado com o número de áreas de coerência do detector, D é o coeficiente de difusão e q_{es} é o módulo do vetor espalhamento. Esta função, assim como funções de correlação de forma geral, informam o grau de correlação entre os valores de intensidade espalhada separadas por um intervalo de tempo τ. Esse grau de correlação, por sua vez, está intimamente relacionado com a velocidade das flutuações provocadas pelo movimento browniano.

A Figura 1.13 apresenta as curvas da função de correlação $(g^{(2)}_{(t)} - 1)/\beta$ em função do tempo (t) obtidas de uma emulsão água-tolueno a 298,2 K.

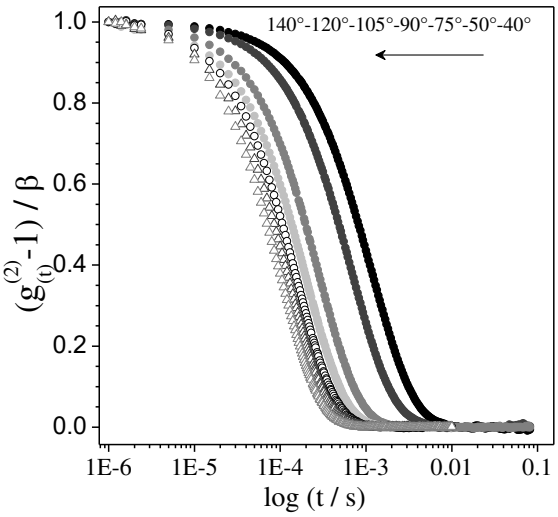

Figura 1.13. Função de correlação de segunda ordem *versus* log (t/s) para ângulos de espalhamento entre 40° e 140° a 298,15 K. Sistema coloidal formado por uma fase continua de tolueno e uma fase dispersa de gotículas de água. Resultado obtido na disciplina métodos experimentais da física II da pós-graduação em física da Universidade Federal de Viçosa.

Na equação 1.39 o grau de correlação é determinado pelo tempo característico de decaimento da exponencial dado por $(Dq_{es}^2)^{-1}$. O grau de correlação também depende do ângulo de observação da luz espalhada. Para ângulos baixos (digamos, 40°) a correlação é mais forte, ou seja, o tempo característico do decaimento exponencial é relativamente grande, porque variações de posição relativas das partículas precisam ser grandes para promoverem flutuações consideráveis na intensidade da luz espalhada, o que acontece com períodos maiores de τ. Em contraste, para ângulos maiores (140°) um mesmo movimento relativo das partículas provoca flutuações mais rápidas na intensidade espalhada. Em outras palavras, a função de correlação decai com um tempo característico menor. Isso justifica a dependência do tempo de correlação com o módulo do vetor de espalhamento que, por sua vez, depende do ângulo de observação, como será visto mais adiante.

Rearranjando a equação 1.39 temos:

$$g_{(t)}^{(2)} = 1 + \beta e^{(-2\Gamma t)} \qquad 1.40$$

Aplicando um ajuste exponencial adequado para cada curva da Figura 1.13 é possível calcular o valor de Γ, o qual possui contribuição tanto do q_{es} como do valor de D.

ii) *Vetor espalhamento (q_{es})*: este é definido como:

$$q_{es} = \frac{4\pi n_r}{\lambda}\left(\text{sen}\,\frac{\theta}{2}\right) \qquad 1.41$$

em que n_r, λ e θ são o índice de refração da solução, o comprimento de onda do laser e ângulo de posição do detector.

iii) *Coeficiente de difusão (D)*: da equação 1.40 define-se a taxa de relaxação como:

$$\Gamma = Dq_{es}^2 \qquad 1.42$$

Conhecendo Γ e q_{es}^2 é possível determinar o coeficiente de difusão. Por definição da primeira lei de Fick, o D é a constante de proporcionalidade entre o fluxo do número efetivo de partículas por mol ($\vec{j}$) e o gradiente de concentração ($\vec{\nabla}c$) numa dada posição e tempo ($\vec{r}, t$) (equação 1.43).

$$\vec{j}(\vec{r}, t) = -D\vec{\nabla}c(\vec{r}, t) \qquad 1.43$$

Em outras palavras, esta constante de proporcionalidade fornece informação sobre a dinâmica do sistema coloidal em estudo, mostrando como as micelas de surfactante se movimentam na fase continua.

iv) *Relação de Stokes-Einstein*: com a obtenção do coeficiente de difusão e utilizando a relação de Stokes-Einstein (equação 1.44), para difusão de partículas esféricas através de um determinado líquido, é possível calcular finalmente o raio hidrodinâmico (R_H).

$$D = \frac{K_B T}{6\pi\eta R_H} \qquad 1.44$$

Em que K_B é a constante de Boltzmann (1,3806 x10^{-23} J K^{-1}), T é a temperatura na qual foi realizado o experimento, η é a viscosidade do sistema coloidal e D é o coeficiente de difusão.

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Capítulo 2: Objetivos

2.1. *Objetivos gerais*

O objetivo principal desta tese de doutorado é contribuir desde uma visão termodinâmica ao entendimento do efeito de líquidos iônicos imidazólicos nos processos de agregação molecular. Para atingir este desafio, duas propostas foram elaboradas:

Proposta 1: avaliar a termodinâmica de micelização do surfactante aniônico dodecilbenzeno sulfonato de sódio (SDBS) em soluções aquosas de líquidos iônicos imidazólicos;

Proposta 2: estudar o efeito de líquidos iônicos sobre a termodinâmica da formação de complexos de inclusão entre β-ciclodextrina e o surfactante catiônico cloreto de dodecilpiridinio.

2.2. *Objetivos específicos*

2.2.1. Objetivos específicos da proposta 1:

- Determinar os parâmetros termodinâmicos de micelização (ΔG^o_{mic}, ΔH^o_{mic} e $T\Delta S^o_{mic}$) do SDBS em água pura e em soluções aquosas de LI, usando as técnicas de condutividade elétrica, espectroscopia de fluorescência, calorimétrica de titulação isotérmica e;
- Utilizar o espalhamento dinâmico de luz para calcular o R_H dos agregados micelares;
- Monitorar o efeito do balanço hidrofóbico/hidrofílico do cation imidazólio sobre a termodinâmica de micelização.

2.2.2. Objetivos específicos da proposta 2:

- Empregar a calorimetria de titulação isotérmica para determinar os parâmetros termodinâmicos da formação do complexo de inclusão (ΔG^o_{ic}, ΔH^o_{ic} e $T\Delta S^o_{ic}$) entre surfactante e ciclodextrina em água pura e em misturas de água e LI;
- Estudar o efeito da hidrofobicidade do cation imidazólio e cosmotropicidade do anion do LI sobre a formação de complexos de inclusão;
- Monitorar as consequências termodinâmicas da alteração das redes tridimensionais das moléculas de água presentes no *bulk* e solvatando as regiões hidrofóbicas da ciclodextrina e do surfactante devido à presença dos LIs.

« Qu'y a-t-il derrière les discussions face à face pour soutenir les prépositions ? Des textes. Et qu'y a-t-il derrière les textes pour les assurer ? D'autres textes qui deviennent de plus techniques parce qu'ils font appel à de plus en plus d'articles. Et qu'y a-t-il derrière ces articles ? Des graphiques, des inscriptions, des étiquetés, des tableaux, des cartes, disposés en étages. Et derrière ces inscriptions ? Des instruments, de toute taille, de tous âges, de tous prix dont la fonction est le marquage et l'enregistrement de différences traces. Et derrière les instruments qui y a-t-il donc ? Divers porte-parole qui commentent les graphiques et exprimant « simplement » ce qu'ils « veulent dire ». Et derrière ces porte-parole afin d'assurer qu'ils disent vrai ? D'autres rangées d'instruments. Et derrière celles-ci ? Des épreuves de force pour évaluer la résistance des liens qui unissent les représentants à leurs représentés »

« Les chercheurs ne disent rien de plus que ce qui est inscrit, mais sans leurs commentaires les inscriptions disent considérablement moins ! Il y a un mot pour décrire cette étrange situation, un mot très important pour ce qui suit, c'est le mot de délégué ou de porte-parole. L'auteur se comporte comme s'il était le porte-parole de ce qui s'inscrit sur l'écran de l'appareil »

« Porte-parole se dit de quelqu'un qui parle pour d'autres qui ne parlent pas »

Bruno Latour
La science en action- Editions La Découverte, Paris, 2005

Capítulo 3: Aggregation of sodium dodecylbenzene sulfonate: weak molecular interactions modulated by the alkyl chain length of imidazolium cation species

Abstract

Ionic liquids (ILs) can modify cooperative processes in aqueous solutions to a large extent, including anionic surfactant aggregation. Here, the micellization of sodium dodecylbenzene sulfonate (SDBS) was evaluated in low concentrations of 1-alkyl-3-methylimidazolium chlorides ($C_n$mimCl, n = 0, 2, and 4) aqueous solutions using fluorescence spectroscopy, isothermal titration calorimetry, dynamic light scattering, and conductimetry measurements. The thermodynamic stability of SDBS aggregates strongly depended on the IL structure and concentration, decreasing in the order C_4mim^+ > C_0mim^+ > C_2mim^+. At 1.00 mmol L^{-1} of the ILs, the increase of the hydrophobicity of the imidazolium cation decreased the enthalpic favorableness, changing ΔH^o_{mic} from −3.75 ± 0.07 kJ mol^{-1}, for C_0mim^+, to −2.69 ± 0.01 kJ mol^{-1}, for C_4mim^+. On the other hand, the entropic feasibility changed in the opposite tendency, suggesting that the higher hydrophobicity of C_4mim^+ overcame the kosmotropic effect of ILs cations in the bulk. We have suggested that the imidazolium cations can interact with the SDBS monomers on the micelle surface, being this interaction mainly via hydrophobic, π-π and electrostatic interactions for C_4mim^+ and C_2mim^+, and electrostatic interactions and hydrogen bonds for C_0mim^+.

Keywords: surfactant micellization, ionic liquids, sodium dodecylbenzene sulfonate, self-assembly, aggregation.

3.1. *Introduction*

The self-assembly of chemical compounds in aqueous solution is a fundamental physical-chemical process in both theorical and practical aspects. Among several molecules with the ability to form aggregates through self-assembly, surfactants have been of great interest due to their wide potential of application in pharmaceutical formulations,[1] environmental remediation,[2] foam generation,[3] personal care products,[4] nanomaterial synthesis,[5] and drug delivery.[6] Surfactants can form aggregates above a narrow concentration range (critical micellar concentration - cmc) at fixed temperature and pressure,[7,8] being the cmc dependent on several properties, specially the molecular properties of the solvent,[9,10] that can be modulated by the addition of cosolutes or cosolvents.

The effect of cosolutes and cosolvents, such as: methanol, propanol, acetamide, urea, NaBr and ethylene glycol[11–13] on the surfactant micellization have been widely studied. In the last fifteen years, ionic liquids (ILs), which are electrolytes with bulky organic cations and inorganic (or organic) anions,[14] have attracted attention as cosolutes,[15–17] because they are able to affect, simultaneously, the ionic and hydrophobic interactions in aqueous systems, modulating different cooperative processes.[18–20] The more commonly studied ionic liquids are those formed by the imidazolium cation due to their negligible vapor pressure, high ionic conductivity, and because their polarity can be modulated by the number of carbons in the lateral chain on the imidazolium cation ring.[21]

The imidazolium ILs have been extensively used to investigate the thermodynamics of micellization of model ionic surfactants, such as sodium dodecyl sulfate (SDS). In 2009, Behera et al.[22,23] studied the effect of hydrophilic (1-butyl-3-methylimidazolium tetrafluoroborate – C_4mimBF_4) and hydrophobic (1-butyl-3-methylimidazolium hexafluorophosphate – C_4mimPF_6) ILs on the SDS micellization, using fluorescence spectroscopy and conductivity. For both ILs, two distinct effects were observed: for C_4mimBF_4, the cmc decreased when the IL concentration increased up to 2% wt, but increased when the IL concentration rose from 2 to 30% wt; for C_4mimPF_6, the cmc of SDS decreased when the IL concentration increased up to 0.1% wt, then remained constant with increasing IL concentration up to 30% wt. From these results, Behera et al. concluded that ILs acted as simple electrolytes at lower concentrations, and they were partitioned from the bulk to the SDS micelles core at higher concentrations, promoting micelle growth. In 2009, Smirnova et al.[24] evaluated the SDS micellization in

1-alkyl-3-methylimidazolium salts + water mixtures by conductivity and calorimetry. The increase of concentration and carbon number in the lateral chain of the imidazolium cation promoted a more intense decrease in the cmc and micellization enthalpy change values (ΔH_{mic}^{o}) when compared with the simple electrolyte NaCl, which was attributed to the changes in electrostatic and hydrophobic interactions in the system promoted by IL cations. In 2013, Pal et al.[25] examined the behavior of SDS solutions in the presence of 3-methyl-5-methylimidazolium hexafluorophosphate making use of conductivity and fluorescence spectroscopy. The IL caused an increase in the cmc values of SDS due to the presence of solvophobic interactions between IL and surfactant chains, enhancing the formation of mixed micelles. In 2015, Ferreira et al.,[19] evaluated the effect of 1-butyl-3-methylimidazolium halides on the SDS micellization using calorimetry, fluorescence, and conductivity, showing that the imidazolium cation introduces its small carbonic chain into the hydrophobic part of the micelle.

Micellar systems based on more complex anionic surfactants such as sodium dodecylbenzene sulfonate (SDBS) have also been studied in the presence of imidazolium ionic liquids. In 2010, Rai et al.[26] studied the size of SDBS micelles in the presence of the C_4mimPF_6 IL using dynamic light scattering (DLS). They showed that for the SDBS concentration of 300 mmol L^{-1} and a concentration of approximately 100 mmol L^{-1} of IL (3:1 stoichiometric ratio) the diameter of the micelles increased drastically by approximately 20 nm due to the incorporation of the imidazolium cations in the micellar aggregate. In 2015, Chabba et al.[27] studied the formation of SDBS micelles in aqueous solutions in the presence of C_nmimCl (n = 8, 10 or 12) using tensiometry, fluorescence spectroscopy and dynamic light scattering techniques. Their results suggested that the formation of mixed micelles and vesicles occurred due to the balance between cation-π, hydrophobic, strong electrostatic and H-bonding interactions.

Despite the extensive information about the effect of imidazole-based ILs on the micellization of model ionic surfactants, systematic thermodynamic studies involving the effect of the hydrophobic/hydrophilic balance of these ILs on the micellization of atypical anionic surfactants, especially in low concentrations of IL, are scarce.

Here, we evaluated the effect of low concentrations of 1-alkyl-3-methylimidazolium chlorides (Figure A.S1 in the supporting information) on the micellization thermodynamic of sodium dodecylbenzene sulfonate (SDBS) in aqueous solution using isothermal titration calorimetry, fluorescence spectroscopy, electrical conductivity, and dynamic light scattering. Thermodynamic parameters of micellization

(ΔG^o_{mic}, ΔH^o_{mic} and $T\Delta S^o_{mic}$) were obtained for IL with distinct concentrations and hydrophobicity.

3.2. *Materials and methods*

3.2.1. Materials:

Sodium dodecylbenzene sulfonate (99.0%), 1-methylimidazolium chloride (95.0%), 1-ethyl-3-methylimidazolium chloride (≥ 95.0%), and 1-butyl-3-methylimidazolium chloride (≥ 98.0%) were obtained from Sigma-Aldrich (USA). Sodium chloride (99.0%) was purchased from Isofar (Brazil). All reagents were used without further purification. Deionized water (18MΩ cm at 298.2 K) from a Milli-QII (Millipore, USA) was used to prepare all solutions.

3.2.2. Fluorescence measurements

Fluorescence emission spectra of pyrene in SDBS aqueous solutions in the presence/absence of ILs were obtained in a Cary Eclipse fluorescence spectrophotometer (Agilent technologies, USA) equipped with a xenon flash lamp and a thermostatic bath (298.2 ± 0.1 K). The excitation was performed at 335 nm and the selected slits were 2.5 nm for excitation and 2.0 nm for emission. The pyrene concentration in solution was 1.2 μmol L^{-1} in order to prevent formation of excimers. The ratio of intensity of the first (I_1 at 373 nm) and third (I_3 at 384 nm) bands of pyrene emission spectra were calculated and analyzed as a function of the log([SDBS] / mol L^{-1}). Experiments using fluorescence emission to monitor the intrinsic fluorescence of benzene ring of SDBS surfactant in the presence and absence of C_4mimCl were also conducted, in which the excitation was performed at 224 nm and the fluorescence emission intensity maximum was collected at 290 nm.

3.2.3. Absorbance measurements

Absorbance spectra of SDBS aqueous solutions at 298.2 ± 0.1 K were obtained in a Cary 50 Probe spectrophotometer from Varian (Mulgrave, Victoria, Australia).

3.2.4. Conductivity measurements

The conductivity measurements were carried out at a digital conductivity meter (DM-32) equipped with a conductivity cell (DMC-010M) from Digimed (Brazil). The conductivity cell was calibrated with aqueous KCl solution (0.01 mmol L^{-1}) supplied by Digimed (Brazil). The temperature of the systems was kept at 298.2 ± 0.1 K during all experiments by circulating thermostated liquid within a jacketed cell where the measurements were performed. The experimental procedure consisted of consecutive additions of 25 or 35 μL of a concentrated surfactant aqueous solution (50 times the SDBS cmc: this value had been previously obtained by the pyrene method) into 10 mL of the solvent present in the jacketed cell. Surfactant addition were performed using a gastight Hamilton Syringe (250 μL).

3.2.5. Isothermal titration calorimetry (ITC)

The calorimetry measurements were performed in a thermal-activity monitor (TAM III) system (TA Instruments-USA) controlled by the TAM assistant^{TM} dedicated software. The calorimeter is equipped with two cells (sample and reference) of 4.0 mL reaction volume. The titrant was added using a Hamilton syringe (500 µL) controlled by a 3810 syringe pump and the calorimetric vessel content was stirred at 180 rpm (3.0 s^{-1}) using a gold helix stirrer. To determine the aggregation parameters of the SDBS in solution, 50 (or 65) consecutive injections of 5 (or 3) µL of concentrated SDBS solution (20 or 200 times above cmc: this value had also been previously determined by fluorescence experiments) were added into the sample cell containing 2.7 mL of solvent at 298.15000 ± 0.00001 K. The solutions inside of syringe, sample and reference cells were prepared in the same solvent. The time interval between two consecutive injections was 900 s, time required for the stabilization of heat flow and regeneration of the baseline.

3.2.6. Dynamic light-scattering measurements

The hydrodynamic radius (R_H) of SDBS micelles in water and water + IL mixtures was determined by DLS on a Zetasizer Nano-ZS (Malvern Instruments, England) at 298.2 ± 0.1 K. The Zetasizer is equipped with a 4-mW He/Ne laser with wavelength of 632.8 nm. The scattered intensity was monitored under a detection angle of 173° with respect

to the source. The micelle hydrodynamic radius (R_H) was calculated using the Stokes-Einstein equation:

$$R_H = \frac{K_B T}{3\pi\eta D_t} \quad (3.1)$$

where K_B is the Boltzmann's constant, T is the absolute temperature, η is the viscosity of the suspension, and D_t is the diffusion coefficient.

3.3. *Results and discussion*

3.3.1. Fluorescence spectroscopy

To evaluate the effect of ILs on the self-assembly of SDBS in aqueous solution, it is necessary to determine its aggregation properties in pure water. The SDBS cmc values reported in the literature vary in a large range, depending on the technique used to access this property and the intrinsic nature of the sample, since this surfactant can be obtained as a mixture of isomers.[28] A method commonly used to measure the cmc of surfactants in aqueous solutions is the fluorometric approach in the presence of pyrene, a fluorescent probe of polarity change.[29] The ratio between the intensities of the first (I_1, at 373 nm) and third (I_3, at 384 nm) bands of pyrene emission spectrum is a sensitive parameter to evaluate changes in the polarity of the environment around it, and it was used to investigate the aggregation behavior of the SDBS. Figure 3.1 shows the I_1/I_3 ratios of pyrene *versus* log([SDBS] / mol L^{-1}) obtained from the titration of concentrated SDBS solution into pyrene aqueous solution.

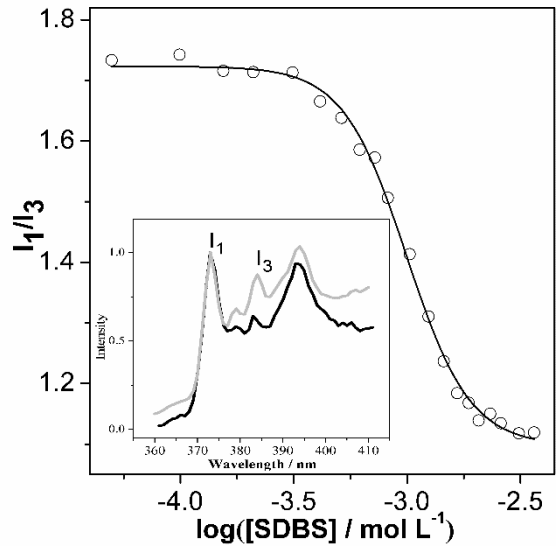

Figure 3.1. Quotient of pyrene vibrational band intensities (I_1/I_3) *versus* log([SDBS] / mol L^{-1}), at 298.2 K. Pyrene concentration was 1.2 μmol L^{-1}. (−) Fitted model to determine the cmc. Inset: normalized pyrene fluorescence spectrum in the pre-micellar (-) and post-micellar (-) region of the SDBS aqueous solution.

At lower SDBS concentrations, corresponding to log([SDBS] / mol L^{-1}) < −3.5, the I_1/I_3 ratio remained almost constant at around 1.74, which confirms that pyrene molecules are in a microenvironment of high polarity as effected by solvation water molecules. For log([SDBS] / mol L^{-1}) values higher than −3.5, the I_1/I_3 ratio abruptly decreases, indicating the reduction of the polarity in the microenvironment around the pyrene molecules due to its gradual partition from the aqueous medium to the hydrophobic core of the SDBS aggregates. Then, when log([SDBS] / mol L^{-1}) reaches −2.5, the I_1/I_3 ratio becomes constant again, indicating that pyrene was almost completely partitioned to the core of the micelles. The cmc of the SDBS was defined as the surfactant concentration at the inflection point of the sigmoidal fit to the I_1/I_3 *versus* log([SDBS] / mol L^{-1}) plot, being equal to 1.01 ± 0.02 mmol L^{-1} in pure water. Different authors have reported cmc values for SDBS in the range of 1.30 – 3.36 mmol L^{-1} at 298.15 K.[28,30–32] The values have been obtained by different techniques, such as conductivity, interfacial tension, isothermal titration calorimetry and fluorescence spectroscopy, being the smallest value obtained by the last technique.

To evaluate the effect of different imidazolium ILs on the SDBS micellization, the cmc of the surfactant was obtained in aqueous systems containing C_0mimCl, C_2mimCl, and C_4mimCl at different concentrations. Experiments in the presence of NaCl

were also performed for comparison. The curves obtained from the fluorometric experiments are shown in Figures A.S2 – A.S5 in the supporting information.

Figure 3.2 shows the *cmc* values of SDBS obtained by the pyrene method as a function of the cosolute concentration.

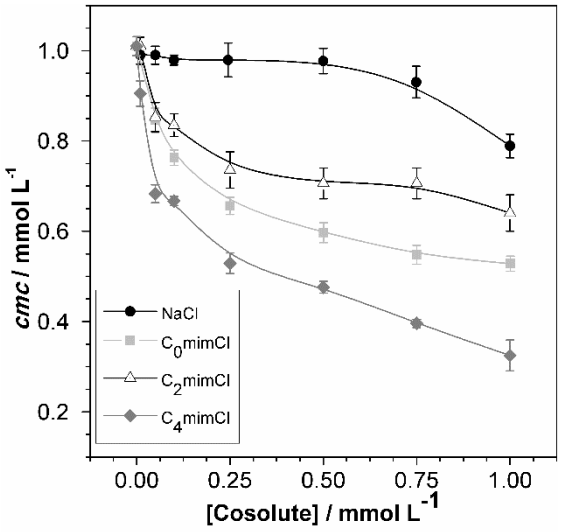

Figure 3.2. Values of *cmc* for SDBS as a function of cosolute concentration. All *cmc* values were obtained by the pyrene method, at 298.2 K.

The *cmc* values decreased with increasing cosolute concentrations and depended on the cosolute nature, showing that the evaluated electrolytes had different abilities to stabilize the SDBS micelles. Since all cosolutes investigated here are formed by the chloride anion, the different abilities of the electrolytes to modify the SDBS aggregation were due to the cation structure. For example, at a concentration of cosolute of 1.00 mmol L^{-1}, the replacement of NaCl by C_4mimCl reduced the *cmc* by three times. In addition, for all investigated cosolute concentrations, the ILs presented higher ability than NaCl to promote micellar stabilization, with *cmc* decreasing in the order C_2mimCl < C_0mimCl < C_4mimCl.

Sodium is a small cation with a high charge density and a hydrated radii equal to 3.58 Å,[33] and its effect on the anionic surfactants micellization is well known.[34] Sodium ions act on the electric double layer of the SDBS micelles, shielding the electrostatic repulsion among the negatively-charged headgroups, and decreasing the electrostatic repulsion among them. The imidazolium cations of the ILs are positively charged, and also promote a similar effect of electrostatic shielding on the SDBS micelle surface.[35] However, because imidazolium cations are large ions that present a delocalized charge

due to the π-electrons resonant effect in the imidazolium ring,[36] we should expect a less intense effect of the ILs compared with the NaCl to shield the electrostatic repulsion among the monomers headgroups on the micelle, which would lead to higher cmc values in presence of the ILs, suggesting that ILs can also act by modulating hydrophobic interactions in the system. Despite this, since the influence of the ILs on the cmc values was not linearly dependent on the number of carbons in the lateral chain of the imidazolium ring, the hydrophobic/hydrophilic balance of the ILs did not determine the stabilization of the SDBS micelles, probably indicating that this stabilization was also a result of a balance between electrostatic and hydrophobic forces controlling the formation of IL-SDBS interactions.

Smirnova et al.[24] determined that the increase of imidazolium IL ($C_n$mimBr, n = 4 or 10) concentration decreased exponentially the cmc of SDS, being this effect more pronounced with the increase in the number of carbons atoms in the lateral chain of the IL cation, probably due to the formation of mixed micelles between surfactant and IL cations. Ferreira at al.[19] reported the effect of the C_4mimCl on the cmc of the SDS, showing that the IL concentration increase promoted an exponential decrease in the cmc values of SDS. This behavior was attributed to the C_4mim^+ shield on electrostatic repulsion among the headgroups on the micelle surface, in addition to the preferential solvation of the surfactant tail by the lateral chain of the imidazolium cations. Jiao et al.[37] evaluated the effect of 0.10 mmol L^{-1} tetraalkylammonium bromides ILs on the aggregation of the 1-butyl-3-methylimidazolium dodecylsulfate (C_4mimDS) and verified that the increase in the number of carbons atoms in the IL cation promoted a greater stabilization of micelles of C_4mimDS than that observed in pure water or in the presence of simple electrolytes (the same behavior observed by Smirnova´s paper, but different from our results). Although the effect reported by Jiao's work is not of an imidazolium IL, and the cation did not possess an aromatic ring, the contribution obtained by these authors brings relevant information about the effect promoted by the increase in the hydrophobicity of IL cation. To our knowledge there are no reports in the scientific literature about the effect of ILs on the micellization thermodynamics of SDBS in aqueous solutions, but the results obtained in this work are consistent with those already reported for SDS, which differs from SDBS by the absence of a benzene ring between the sulfonate group and the linear carbonic chain.

3.3.2. Electrical conductivity

To evaluate the thermodynamic stability of the SDBS micelles in the studied systems, the standard Gibbs free energy change of micellization (ΔG^o_{mic}) for the SDBS micellization process was calculate using equation 3.2.

$$\Delta G^o_{mic} = (2 - \alpha)RTln(cmc) \quad 3.2$$

where R, T, and α are the gas constant, absolute temperature, and the degree of counter-ion dissociation, respectively, in which α value was obtained by equation 3.3:

$$\alpha = \frac{s_2}{s_1} \quad 3.3$$

where s_1 and s_2 are the slopes of the linear fits in the pre- and post-micellar regions in the conductometric curves The curves obtained from the conductivity experiments in pure water and IL + water mixtures (or NaCl + water mixtures) are shown in Figures A.S6 – A.S10 in the supporting information.

Table 3.1 shows the values of α and ΔG^o_{mic} obtained for all evaluated systems.

Table 3.1. Degree of counter-ion dissociation (α) and standard Gibbs free energy change (ΔG^o_{mic}) values obtained for the SDBS micellization in different IL + water or NaCl + water mixtures, at 298.2 K.

Concentration	NaCl		C_0mimCl		C_2mimCl		C_4mimCl	
mmol L^{-1}	α	ΔG^o_{mic}	α	ΔG^o_{mic}	α	ΔG^o_{mic}	α	ΔG^o_{mic}
0.00	0.84	−19.8	0.84	−19.8	0.84	−19.8	0.84	−19.8
0.01	0.82	−20.3	0.83	−20.0	0.81	−20.4	0.86	−19.9
0.05	0.82	−20.3	0.87	−19.8	0.81	−20.7	0.87	−20.4
0.10	0.82	−20.6	0.88	−19.9	0.81	−20.7	0.87	−20.5
0.25	0.81	−20.5	0.88	−20.4	0.82	−21.1	0.88	−21.2
0.50	0.81	−20.4	0.88	−20.6	0.83	−21.1	0.88	−21.2
0.75	0.82	−20.4	0.88	−20.9	0.85	−20.7	0.89	−21.7
1.00	0.81	−20.7	0.88	−20.9	0.86	−20.8	0.90	−22.0

The linear regression analysis of the conductivity data in the pre- and post-micellar regions had a correlation factor higher than 0.9992; the error associated with the α and ΔG^o_{mic} values was lower than 1.5%; the cmc values used to calculated the ΔG^o_{mic} values were obtained by the pyrene method. The unit of ΔG^o_{mic} parameter is kJ mol^{-1}.

The α value obtained in pure water was 0.84 ± 0.01, indicating that only 16% of sodium ions present in the system are adsorbed on the SDBS micelle surface. Similar

result (0.82) was reported by Chauan et al.[38] Moreover, the α values depended slightly on the concentration and cation nature of the cosolute. While NaCl promoted a gradual decrease in the α values from 0.84 (pure water) to 0.81 (1.00 mmol L^{-1} NaCl), the C_nmim^+ cation (n = 0 or 4) promoted an increase in the α values as the IL concentration increased, indicating that C_0mim^+ and C_4mim^+ cations acted directly on the micelle surface. It is therefore likely that C_0mim^+ and C_4mim^+ imidazolium cations adsorbed at the interface of the SDBS micelles, decreasing their negative charge density and promoting a decrease in the number of Na^+ ions on the micellar surface. The effects of C_4mim^+ and C_0mim^+ were similar. Interestingly, while the positive charge density of the imidazolium cations increases in the order $C_4mim^+ < C_0mim^+$, the hydrophobicity increases in the opposite order, showing that the adsorption of imidazolium cations on the micelle surface is a complex process.

For C_2mim^+, the α values showed two behaviors: for C_2mim^+ concentrations lower than 0.25 mmol L^{-1}, the α values decreased, and then they increased for higher concentrations, suggesting that higher amounts of this IL are needed to promote its adsorption on the SDBS micelle surface.

Independent of the investigated system, the ΔG^o_{mic} value was negative, indicating that SDBS micelles predominated on SDBS monomers in the equilibrium state. When the cosolute concentration increased up to 1.00 mmol L^{-1}, the stability of the SDBS micelles as compared to the monomers in the solution followed the electrolyte sequence NaCl < C_2mimCl < C_0mimCl < C_4mimCl. In order to comprehend this relative thermodynamic stability of the SDBS aggregates in the presence of cosolutes, it is necessary to evaluate the energetic contributions associated with intermolecular interactions and chemical species distribution in the system when the SDBS self-assembly process occurs. These contributions can be better elucidated by determining the enthalpy and entropy changes associated with the aggregation process, as discussed in the following sections.

3.3.3. Isothermal titration calorimetry

The balance of intermolecular interactions responsible for the formation of SDBS aggregates in water + ILs mixtures can be studied by isothermal titration calorimetry (ITC) measurements. ITC is a technique that allows to monitor the enthalpic change in cooperative molecular processes.[39] From ITC thermograms, the observed enthalpy change (ΔH_{obs}) due to the addition of SDBS into the calorimetric vessel can be obtained

as a function of surfactant concentration, and the critical micellar concentration (cmc) and the enthalpy change associated with the micelle formation in solution (ΔH^o_{mic}) can be determined. Furthermore, information about the intermolecular interactions between solvent and SDBS monomers, among SDBS monomers in the bulk, and among SDBS monomers in the micelle can be accessed. Figure 3.3 presents the ITC curve obtained by titrating a 25 mmol L^{-1} SDBS solution into pure water, at 298.15 K, obtained from the ITC thermogram showed in Figure A.S11 in the supporting information.

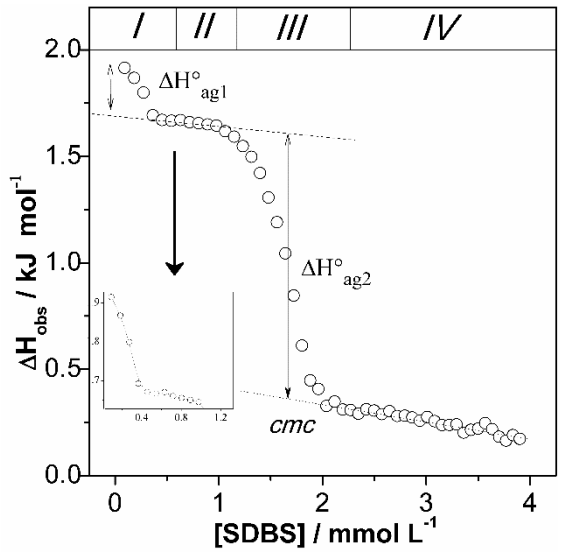

Figure 3.3. ΔH_{obs} values *versus* [SDBS] in pure water, at 298.15 K. Inset: magnification of the curve up to 1.2 mmol L^{-1} (region *I*).

The calorimetric curve for titration of surfactant solution, containing a high micelle concentration, in pure water presented a sigmoidal profile generally found for other surfactants when the SDBS concentration was higher than 0.59 mmol L^{-1}. However, different from traditional profiles found in the literature, for SDBS concentrations lower than the cmc ([SDBS] / mmol $L^{-1} \leq 0.59$), the curve exhibited a distinct behavior, as indicated in the inset of Figure 3.3 (region *I*).

The common sigmoidal profile for [SDBS] / mmol $L^{-1} > 0.59$ can be divided in three parts, i.e., pre-micellar, micelle formation, and post-micellar regions,[40] indicated by regions *II*, *III*, and *IV*, respectively, in Figure 3.3. In region *II* ($0.59 <$ [SDBS] / mmol $L^{-1} \leq 1.20$), the concentration of SDBS in the sample calorimetric cell is lower than the cmc, and the ΔH_{obs} values are associated with a set of processes that occur after the dilution and breakup of SDBS micelles. Then, for $1.20 <$ [SDBS] / mmol $L^{-1} \leq 2.27$

(region *III*), the addition of the concentrated surfactant solution promotes only a partial breakup of the added micelles, and the cmc value is attained in the calorimetric cell. The cmc value can be obtained using the first derivative method.[41] Finally, in region *IV*, 2.27 < [SDBS] / mmol $L^{-1} \leq 3.92$, the beginning of the post-micellar region is attained, where all micelles added into the sample calorimetric cell are diluted.

In region *I,* the ΔH_{obs} values decreased intensely until becoming almost constant in region *II*, suggesting an extra cooperative process of the surfactant monomers, probably involving the formation of small aggregates, such as dimers, trimmers, etc. The SDBS molecules have an aromatic ring and can interact with one another by π-π stacking interactions, promoting the formation of small SDBS aggregates. Since the SDBS benzene ring has an intrinsic fluorescence that allows to determine the excimers formation,[42,43] the SDBS solutions were analyzed using fluorescence and UV-vis spectroscopies in order to verify the formation of SDBS small aggregates in pure water (Figure A.S12 in the supporting information shows the curves of fluorescence intensity and absorbance against the SDBS concentration).

When the SDBS concentration increased up to 0.12 mmol L^{-1} (Figure A.S12a), the fluorescence emission intensity increased linearly from 38.9 ± 0.5 to 126.5 ± 0.4 due to the increase in the content of SDBS monomers in the system. Above 0.20 mmol L^{-1} SDBS, fluorescence quenching started, and the intensity of the emission band of SDBS decreased exponentially. This progressive quenching agrees with the model that describes the equilibrium between chemical species in monomeric (fluorescent) and aggregate (non-fluorescent) state in solution.[44,45] According to fluorescence data, the increase in the SDBS concentration up to 0.20 mmol L^{-1} also promoted a linear response in the absorbance (Figure A.S12b). Then, for [SDBS] = 0.24 mmol L^{-1}, the linearity of the absorbance *versus* [SDBS] curve was lost, probably due to the attractive non-covalent interaction (π-π interactions principally) between SDBS monomers that changed the energy of the electronic states of SDBS and, consequently, the molar absorptivity of the surfactant.

The critical surfactant concentrations, 0.20 and 0.24 mmol L^{-1}, obtained by fluorescence and UV-vis spectroscopies, respectively, are very close to 0.26 mmol L^{-1}, which corresponds to the inflection point of the first abrupt decrease in the ΔH_{obs} values in the calorimetric curve (region *I,* Figure 3.3), confirming our hypotheses. In view of this, to access the enthalpy change associated with the micelle formation, we should consider the following equilibria involved in the SDBS micellization process:

$$nDBS^-_{(aq)} + (n-p)S^+_{(aq)} \rightleftharpoons S_{n-p}DBS_n{}^{p-}_{(aq)} \quad 3.4$$
$$mS_{n-p}DBS_n{}^{p-}_{(aq)} + (mp-q)S^+_{(aq)} \rightleftharpoons S_{mn-q}DBS_{nm}{}^{q-}_{(aq)}$$

where DBS^- and S^+ are the ionized surfactant monomers and the sodium ion, respectively, $S_{n-p}DBS_n{}^{p-}$ are the small aggregates of surfactant (n = 1, 2, 3, ...) with effective charge p^-, and $S_{mn-q}DBS_{nm}{}^{q-}$ are the formed micelle with nm monomers and effective surface charge q^-. The standard micellization enthalpy change of SDBS (ΔH^o_{mic}), i.e., the enthalpy change when one mole of surfactant monomer are transformed into micelles, can be calculated using equation 3.5, considering that the micellization occurred in the two stages showed in equation 3.4.

$$\Delta H^o_{mic} = \Delta H^o_{ag1} + \Delta H^o_{ag2} \quad 3.5$$

where ΔH^o_{ag1} and ΔH^o_{ag2} are the aggregation enthalpy changes when one mole of SDBS change from its monomeric state to its small aggregate state (region *I* in Figure 3.3), and from its small aggregate state to its micellar state (region *III* in Figure 3.3), respectively. The ΔH^o_{ag2} value was determined by subtracting $\Delta H_{obs,2}$ from $\Delta H_{obs,1}$, the observed enthalpy changes obtained in the extrapolation point of the linear behavior in regions *IV* and *II* (dotted lines shown in Figure 3.3) at the cmc.[46] However, because there was not a first plateau in the beginning of the calorimetric curve, a Boltzmann's sigmoidal equation[47] was fitted to the experimental data of regions *I* and *II*, and used to calculate the ΔH^o_{ag1} value from the difference of ΔH_{obs} values in the lower and upper limits of the sigmoidal function (fits and details of ΔH^o_{ag1} and ΔH^o_{ag2} determination are shown in Figure A.S13 in the supporting information).

The value of ΔH^o_{mic} for SDBS in pure water was −1.52 ± 0.04 kJ mol^{-1} (ΔH^o_{ag1}= −0.26 ± 0.02 kJ mol^{-1} and ΔH^o_{ag2}= −1.26 ± 0.03 kJ mol^{-1}), showing that the micellization process was enthalpically favorable. Hait et al.,[28] using ITC, obtained a ΔH^o_{mic} value for SDBS in pure water equal to −1.0 kJ mol^{-1}, 33% smaller than that obtained here. This difference could be attributed to the energetic contribution of the formation of SDBS small aggregates to ΔH^o_{mic}, which was not considered in Hait's work because the first point of his calorimetric curve was higher than 0.50 mmol L^{-1} SDBS, i.e., above the critical concentration for the formation of the small aggregates.

The change in the system enthalpy due the micellization process of SDBS can be considered as the sum of three independent terms:

$$\Delta H_{mic}^{o} = \Delta H_{des}^{o} + \Delta H_{elec}^{o} + \Delta H_{surf-surf}^{o} \quad 3.6$$

where ΔH_{des}^{o} is the standard enthalpy change associated with the desolvation process of the SDBS monomers, including the disruption of interactions between water and SDBS and among water molecules surrounding the hydrophobic chain of SDBS, besides the formation of water-water interactions in the bulk, being probably positive;[48] ΔH_{elec}^{o} is the standard enthalpy change associated with the electrostatic repulsion between the headgroups of the SDBS in the aggregate state, being positive;[49] and $\Delta H_{surf-surf}^{o}$ is the standard enthalpy change due to the formation of interactions among the SDBS tails in the micellar state, including dispersive interactions between the carbon chains and π-π stacking interactions between the aromatic rings, being negative. In this way, ΔH_{mic}^{o} for SDBS in pure water was exothermic because $|\Delta H_{des}^{o}| + |\Delta H_{elec}^{o}| < |\Delta H_{surf-surf}^{o}|$, indicating an important role of the weak interactions to the standard enthalpy change of micellization.

From the ΔG_{mic}^{o} and ΔH_{mic}^{o} values, it was possible to determine the entropy change of micellization ($T\Delta S_{mic}^{o}$) using the fundamental relationship given by:

$$\Delta G_{mic}^{o} = \Delta H_{mic}^{o} - T\Delta S_{mic}^{o} \quad 3.7$$

where T is the absolute temperature. In pure water, $T\Delta S_{mic}^{o}$ was 18.3 ± 0.2 kJ mol^{-1}, showing that the SDBS micellization process was entropically favorable, mainly due to the entropic gain from the release of water molecules solvating the SDBS tail when the micellization process occurred.

3.3.3.1. ITC curves for SDBS micellization in the presence of ILs

To understand how ILs affected the enthalpic and entropic contributions for ΔG_{mic}^{o}, the ΔH_{mic}^{o} and $T\Delta S_{mic}^{o}$ parameters should be determined in the presence of ILs. Then, calorimetric curves of SDBS aggregation were obtained in the presence of different

ILs concentrations. Figure 3.4 shows the effects of C4mimCl on ΔH_{obs} *versus* [SDBS] curves.

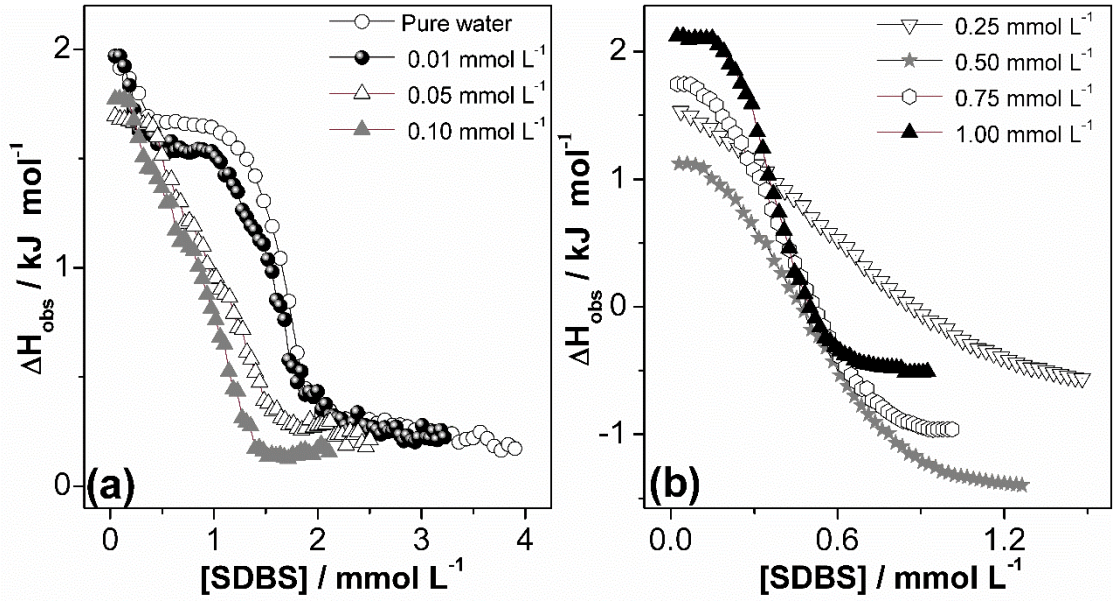

Figure 3.4. ΔH_{obs} *versus* [SDBS] curves obtained by dilution of SDBS solution in C4mimCl + water mixtures, at 298.15 K: **(a)** [C4mimCl] ≤ 0.10 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} ≤ [C4mimCl] ≤ 1.00 mmol L^{-1}.

At 0.01 mmol L^{-1} C4mimCl, the calorimetric curve profile was very similar to that observed in pure water. However, as the IL concentration increased, that profile was replaced with a single sigmoid, suggesting that the transition of SDBS monomers to micelles occurred without the formation of small aggregates. This hypothesis was confirmed by experiments of intrinsic fluorescence of SDBS in the presence of C4mimCl (Figure A.S14 in the supporting information), that showed only one abrupt transition in the SDBS intrinsic fluorescence *versus* [SDBS] curve.

These results indicated that the C4mimCl increased the stability of the SDBS monomers and micelles relative to the stability of the small aggregates. The hydrophobic chain of C_4mim^+ cation probably facilitated its intimate adsorption on the SDBS micelle interface through hydrophobic interactions, providing an efficient shielding between the SDBS headgroups due to the following phenomena: *i*) establishment of electrostatic interactions between the negatively-charged headgroup of the surfactant and the positively-charged imidazolium ring; *ii*) π-π stacking interactions between imidazolium and benzene rings in IL and surfactant, respectively, *iii*) cation-π interactions between imidazolium cation and the benzene ring of the surfactant molecule, and *iv*) H-bonding

interactions between the more acidic proton of the imidazolium ring and the SDBS headgroup.[27,50]

To verify that the stability of the SDBS small aggregates was not exclusively affected by electrostatic contributions, calorimetric curves were obtained in the presence of NaCl (Figure A.S15 in the supporting information). Interestingly, for all NaCl concentrations studied, the ΔH_{obs} *versus* [SDBS] curve profiles were very similar to those in pure water, showing that only electrostatic shielding cannot promote the destabilization of the small aggregates, thereby requiring specific interactions between the cosolute and surfactant monomers, which can be promoted by synergistic effects of hydrophobic, dispersive and electrostatic interactions occurring simultaneously.

For the series of ILs evaluated in this work, the balance among those interactions can be affected by the alkyl tail length of the imidazolium cation, thereby highlighting the importance of analyzing the calorimetric curves in the presence of C_2mimCl and C_0mimCl. Figures 3.5 and 3.6 show the ΔH_{obs} *versus* [SDBS] curves obtained in the presence of C_2mimCl and C_0mimCl, respectively.

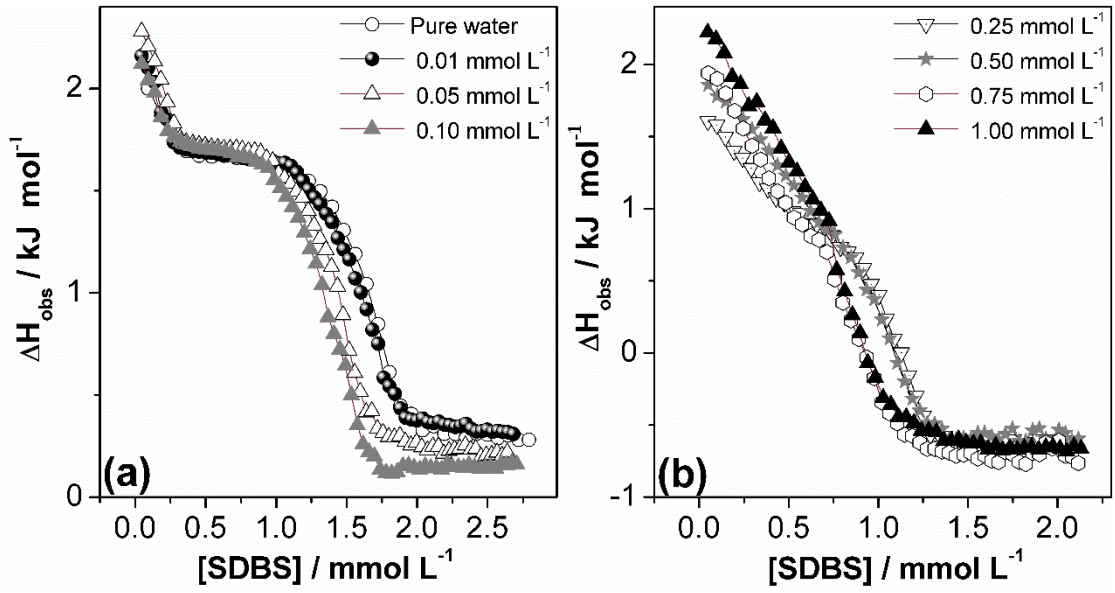

Figure 3.5. ΔH_{obs} *versus* [SDBS] curves obtained by dilution of SDBS solution in C_2mimCl + water mixtures. **(a)** [C_2mimCl] ≤ 0.10 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} ≤ [C_2mimCl] ≤ 1.00 mmol L^{-1}, at 298.15 K.

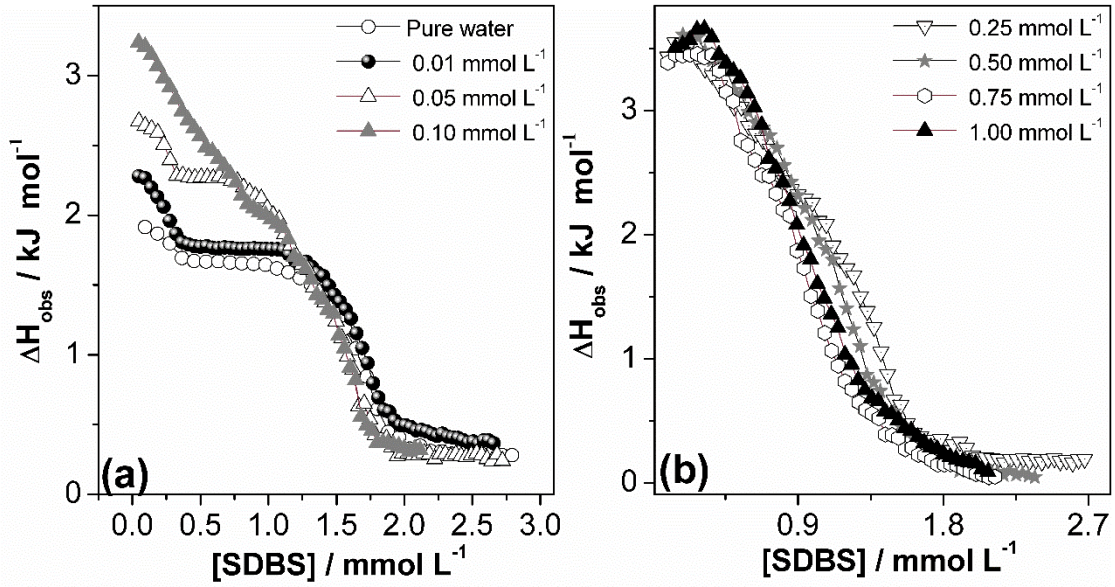

Figure 3.6. ΔH_{obs} *versus* [SDBS] curves obtained by dilution of SDBS solution in C_0mimCl + water mixtures. **(a)** [C_0mimCl] ≤ 0.10 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} ≤ [C_0mimCl] ≤ 1.00 mmol L^{-1}, at 298.15 K

For C_2mimCl concentrations up to 0.10 mmol L^{-1} (Figure 3.5a), the calorimetric curves maintained the same profile observed in pure water, indicating that the formation of small aggregates also occurred for this IL concentration range. However, for C_2mimCl concentrations from 0.25 to 1.00 mmol L^{-1} (Figure 3.5b), the calorimetric curve profiles were very distinct from those observed in the presence of C_4mimCl, and the ΔH_{obs} values decreased almost linearly up to the SDBS concentration equal to 0.75 mmol L^{-1}. Then, although the ΔH_{obs} values were abruptly decrease above that SDBS concentration, the slope of the curve changed. These results suggested that, at higher C_2mim^+ concentrations, the SDBS small aggregates were formed and progressively became micelles with increasing SDBS concentration. The C_2mimCl cation contains methyl and ethyl moieities in the imidazolium ring, while C_4mimCl has methyl and butyl groups, suggesting that the direct transition from monomers to micelle did not occurr for C_2mimCl within the investigated concentration range due to the lower hydrophobicity of C_2mimCl, which did not favor intense hydrophobic interactions between the hydrophobic tails of both SDBS and IL in the micelle.

For C_0mimCl concentration up to 0.05 mmol L^{-1} (Figure 3.6a), the calorimetric curves were also characterized by the presence of four regions, indicating the occurrence of the same processes discussed previously from Figure 3.3. However, within the concentration range associated with the formation of small aggregates, ΔH_{obs} increased

as the C_0mimCl concentration increased. At 0.10 mmol L^{-1} C_0mimCl, the calorimetric curve presented the same behavior as that observed for higher C_2mimCl concentrations (Figure 3.5b), indicating a higher ability of this IL to interact with the SDBS monomers in the micelle. For the higher concentrations of C_0mimCl (Figure 3.6b) the calorimetric curves were very similar to those obtained for higher concentrations of C_4mimCl, indicating the direct transition of SDBS monomer to micelles.

3.3.3.2. Thermodynamic parameters of aggregation

Comparing the values of ΔG_{mic}^{o} for C_4mimCl with those for C_0mimCl and C_2mimCl (Table 3.1), we can observe that the SDBS micelles were more thermodynamically stable in the presence of C_4mimCl, probably associated with the higher hydrophobicity of the C_4mim^+ cation. Nonetheless, the comparison between C_0mimCl and C_2mimCl showed a more complex effect of the imidazolium structure on the stabilization of those aggregates, indicating that the hydrophilic/hydrophobic balance of the ILs did not determine the capacity of the IL to promote the stability of the SDBS micelles. The extra acidic hydrogen present in the imidazolium cation of the C_0mimCl probably plays an important role on the micelle stability.

Therefore, we propose that the hydrophobic interaction between the SDBS and the lateral chain with four carbon atoms in C_4mimCl is determinant of the higher thermodynamic stability of the aggregates of SDBS in this IL in comparison with C_2mimCl and C_0mimCl, whereas the hydrogen bond between the more acidic proton of the imidazolium ring of C_0mimCl and the sulfonate group of the SDBS is determinant of the higher thermodynamic stability of the SDBS aggregates in C_0mimCl than in C_2mimCl.

To corroborate our hypothesis, the contributions of enthalpy change, ΔH_{mic}^{o}, and entropy change, $T\Delta S_{mic}^{o}$, of SDBS micellization were obtained, allowing to completely elucidate the thermodynamic information of the investigated process. All thermodynamic parameters associated with the SDBS micellization process are compiled in Tables A.S1 and A.S2 in the supporting information.

The ΔH_{mic}^{o} values were negative in all investigated conditions, being that the increase in the IL concentration rendered the process more enthalpically favorable. On the other hand, although the ΔH_{mic}^{o} values were more exothermic in the presence of lower

NaCl concentrations (0.01 mmol $L^{-1} \leq$ [NaCl] $\leq$ 0.75 mmol L^{-1}) than in pure water, the ΔH_{mic}^{o} values gradually increased as the NaCl concentration increased up to 1.0 mmol L^{-1}. Additionally, for the same concentration of the ILs, the ΔH_{mic}^{o} values were more negative in the order C_0mimCl > C_2mimCl > C_4mimCl. Regardless of the nature and quantity in solution of the electrolyte used, the thermodynamic process of micellization was entropically driven ($T\Delta S_{mic}^{o} > 0$); however, different trends were observed depending on the nature of the cation. The increase in the NaCl and C_4mimCl concentrations made the $T\Delta S_{mic}^{o}$ values more positive when compared with the value obtained in pure water, while the opposite effect was observed for C_0mimCl and C_2mimCl.

In the presence of NaCl concentrations up to 0.75 mmol L^{-1}, sodium ions promoted a higher electrostatic shielding and increased the fraction of ions in the double electric shell of the micelle, contributing to the more negative ΔH_{mic}^{o} values than those obtained in pure water. However, as Na^+ and Cl^- ions concentrations increased, the α values remained almost constant, suggesting that the increase in ΔH_{mic}^{o} values was mainly because of other effects promoted by Na^+ and Cl^- ions in the bulk of solution.

In the Hofmeister series, the Na^+ and Cl^- ions can behave as kosmotropic or chaotropic agents depending on the system.[51–54] Our results suggest that the sodium and chloride ions act as a structure-breaker of the three-dimensional network of water molecules, decreasing the molar enthalpy of the water molecules in the bulk. This effect is more pronounced at higher salt concentrations. Therefore, ΔH_{mic}^{o} became less negative as NaCl concentration increased, due to the decrease in the ΔH_{des}^{o} contribution in equation 3.6. This is probably associated with the lower energy released from the formation of water-water interactions in the bulk when the micellization process occurs in the presence of NaCl. Simultaneously, the rotational and translational freedom degrees of the water molecules increased in the bulk, and the $T\Delta S_{mic}^{o}$ values increased when the water molecules migrated from the solvation layer of the surfactant tail to the bulk, where the water molecules have higher molar entropy than those in the bulk composed of pure water, due to the chaotropic effect of the salt.

To understand the effect of ILs on the values of ΔH_{mic}^{o} and $T\Delta S_{mic}^{o}$, it is necessary to consider three aspects of these electrolytes in solution: *i*) the capacity of $C_n$mim$^+$ to interact with DBS^- monomers through specific interactions; *ii*) the chaotropic behavior

of Cl^- anions; and *iii*) the kosmotropic behavior of imidazolium cations, being this effect enhanced with the number of carbon atoms of the side chain.[55]

The chaotropic effect of the chloride anion acts equally for all ILs, and it would lead to less negative ΔH_{mic}^{o} values with increasing IL concentration, as observed in the presence of NaCl. Therefore, this effect was not predominant to determine the IL effect on the SDBS micellization. However, the kosmotropic effect promoted by C_nmim^+ would act in an opposite way, making ΔH_{mic}^{o} values more negative at higher IL concentrations, as observed here. At the same time, specific interactions (hydrogen bonds and/or hydrophobic interactions) between the imidazolium cations and the SDBS monomers in the micelle could promote the substitution of sodium ions by the imidazolium cation in the double electric layer of the SDBS aggregates. This hypothesis is corroborated by the α values, which slightly increased in the presence of C_4mimCl, C_0mimCl, and higher concentrations of C_2mimCl. The replacement of Na^+ with C_nmim^+ contributed to a synergistic effect that decreased the electrostatic repulsion among SDBS headgroups in the micellar state to a larger extension than sodium ions. Thus, the higher the IL concentration, the higher the imidazolium cation concentration in the micelle interface, returning more negative ΔH_{mic}^{o} values.

Concerning the dependence of the thermodynamic parameters on the hydrophobic/hydrophilic balance, the ΔH_{mic}^{o} values became less negative and the $T\Delta S_{mic}^{o}$ values became more positive in the same order in which the IL hydrophobicity increased, i.e, $C_0mim^+ < C_2mim^+ < C_4mim^+$. Both specific interactions between C_nmim^+ and DBS^- and the kosmotropic behavior of imidazolium cation are highly dependent on the hydrophobic/hydrophilic balance of the imidazolium cation, determining those different effects of the cations on the thermodynamic parameters of aggregation. It is well known that the removal of water molecules from a hydrophobic region is an endothermic process,[48] being necessary to spend energy to desolvate the hydrophobic parts of the SDBS monomers and IL cations that could be introduced in the inner of the micelle. The carbon chain of the C_4mim^+ was probably introduced in the inner of the micelle, releasing a higher number of water molecules with high partial molar enthalpy and low partial molar entropy to the bulk compared with C_0mim^+ and C_2mim^+, explaining the more positive $T\Delta S_{mic}^{o}$ values and less negative ΔH_{mic}^{o} values for that IL. Nonetheless, for the $C_n$mimCl ionic liquids with $n < 4$, a decrease in the $T\Delta S_{mic}^{o}$ values occurred in the system

with the increase of the IL concentration, indicating that the kosmotropic effect of these cations in the bulk of the solution predominated.

3.3.3.3. SDBS aggregates at higher surfactant concentration

If some of the IL cations entered the SDBS micelle, the increase of micellar concentration into the calorimetric cell would generate micelles with different IL concentrations in them, since the IL concentration is fixed. In order to verify this hypothesis, calorimetric curves of SDBS titration into electrolyte solutions were also obtained for a broader range of surfactant concentrations. Figure 3.7 presents the calorimetric curves obtained at that condition, where the final SDBS concentration in the calorimetric cell was at least 6 times higher than the cmc, in the presence of the studied electrolytes.

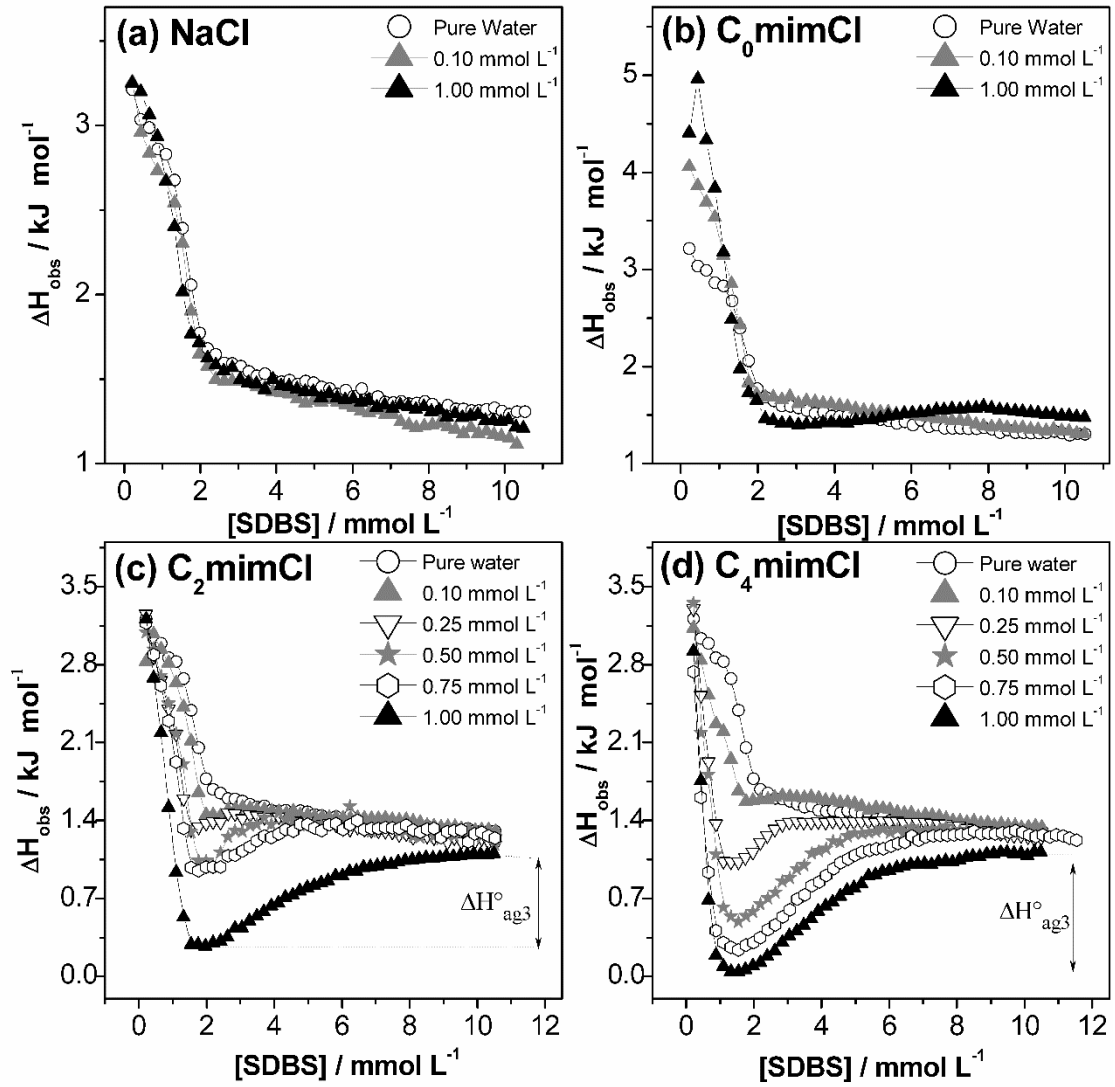

Figure 3.7. ΔH_{obs} *versus* [SDBS] curves obtained by titration of SDBS 200 mmol L^{-1} in 2.7 mL of aqueous solutions containing different concentrations of **(a)** NaCl, **(b)** C_0mimCl, **(c)** C_2mimCl, and **(d)** C_4mimCl, at 298.15 K.

For all investigated conditions, in the range of SDBS concentration up to approximately 2 mmol L^{-1}, the ΔH_{obs} values changed due to the thermodynamic processes already described above (Figures 3.3 – 3.6 and A.S15). However, for SDBS concentrations higher than 2 mmol L^{-1}, especially in the presence of C_2mimCl and C_4mimCl, the ΔH_{obs} values increased with increasing SDBS concentration until attaining a plateau, indicating the occurrence of an endothermic thermodynamic event. This event was more intense with the increase in the concentration and hydrophobicity of the IL cation. Interestingly, this effect was less intense for high IL concentration when the IL was C_0mimCl and did not appear in the presence of NaCl.

In pure water, for SDBS concentrations higher than the cmc, the ΔH_{obs} values are related to the micelle dilution when the SDBS concentrated solution is added into the calorimetric cell, reflecting micelle-micelle interactions that are basically driven by electrostatic repulsions. In the presence of NaCl, which acts by shielding the electrostatic repulsions, no effects on the ΔH_{obs} values were observed, compared to the values in pure water (Figure 3.7a), probably due to the slight decrease in the ionization degree of the micelles promoted by this electrolyte (Table 3.1) which overcame the shielding effect.

For the calorimetric curves obtained in the presence of C_2mimCl and C_4mimCl (Figures 3.7c and 3.7d), the new endothermic event above SDBS 2 mmol L^{-1} corroborated our hypothesis that the imidazolium cations of the IL interacted with the micelles in a different way, compared to NaCl. Recently, our research group,[19] investigating the effect of C_4mimBr and C_4mimCl on the SDS micellization, also observed a similar profile in the calorimetric curves for high SDS concentrations in the presence of IL. The endothermic event in the post-micellization region was attributed to the restructuring of the mixed SDS/IL micelles in the system forming smaller aggregates, which was confirmed by dynamic light scattering.

Dynamic light scattering (DLS) allows to obtain physical properties of dispersed phases in colloidal systems,[56] such as hydrodynamic radius (R_H) and the dynamics of aggregates in solution. Here, DLS was used to evaluate the effect of the electrolytes on the R_H values of SDBS aggregates, hence enabling to examine the effect of the IL cation structure on the aggregates size for higher SDBS concentrations. Figure 3.8 shows the R_H values *versus* SDBS concentrations curves obtained in pure water and in the presence of 1.00 mmol L^{-1} of the studied electrolytes.

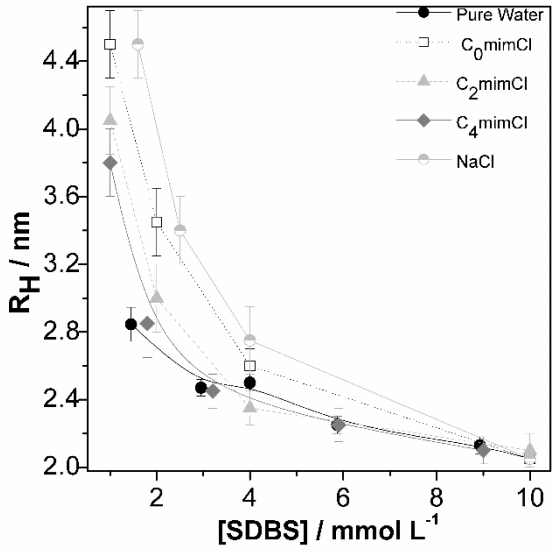

Figure 3.8. R_H *versus* [SDBS] curves obtained in pure water and in presence of 1.00 mmol L^{-1} electrolyte, at 298.15 K.

In pure water, the increase of SDBS concentration exponentially decreased the R_H of the aggregates, as already reported in the literature.[57] The same behavior was observed in the presence of all electrolytes, but the R_H values were higher than those in pure water, following the order NaCl > C_0mim^+ > C_2mim^+ > C_4mim^+, except for the highest SDBS concentration applied (10 mmol L^{-1}), for which the R_H values were the same for all systems (around 2.1 nm).

In water, the decrease in R_H with increasing SDBS concentration was attributed to the increase in the diffusion coefficient of the micelles in the system promoted by the higher electrostatic repulsion among the aggregates as the concentration of micelles increases. The same hypothesis could be considered for the effect observed in the presence of the other electrolytes.

The higher R_H values of the aggregates in the presence of NaCl at the lower SDBS concentration was probably associated with the increase in the aggregation number of the micelles, which was promoted by the decrease in the electrostatic repulsion among the SDBS head groups on the micelle surface.[58] This effect was also promoted by the ILs, but, as we have suggested, the ILs could stay closer to the micelle surface, especially with regard to C_4mim^+, which introduced its hydrophobic tail into the hydrophobic core of the micelle, preventing that a higher number of SDBS monomers entered the micelle. Interestingly, C_2mim^+, being less hydrophobic than C_4mim^+, and C_0mim^+, which interacted with the sulfonate groups in the micelle surface through hydrogen bonds,

promoted a less intense effect due to the low concentration of these cations on the micelle surface.

The clearer evidence that the effects of the electrolytes on the change of the micelle size were different is the endothermic event for the higher SDBS concentration (Figure 3.7). We believe that when the SDBS concentration in the system increased, attaining 10 mmol L^{-1} for instance, the n(SDBS):n(C_4mimCl) molar ratio increased, promoting the dilution of IL on the micelle surface, being this result in agreement with those obtained by Rai et al.[26] It is important to highlight that the location of the C_0mim^+ cation on the micelle surface is not like that of Na^+, C_2mim^+, and C_4mim^+ cations.

The enthalpy change associated with the endothermic process (ΔH^o_{ag3}) in Figure 3.7 was determined in the same way as we determined ΔH^o_{ag2}, using a Boltzmann sigmoid fit. All obtained values are reported in Table 3.2.

Table 3.2. Standard enthalpy change (ΔH^o_{ag3}) values calculated in the SDBS post-micellar region for different water + IL mixtures, at 298.15 K.

[IL]	C_0mimCl	C_2mimCl	C_4mimCl
mmol L^{-1}	ΔH^o_{ag3} / kJ mol^{-1}		
0.10	−	0.06 ± 0.01	0.04 ± 0.01
0.25	−	0.12 ± 0.01	0.37 ± 0.00
0.50	−	0.39 ± 0.01	0.94 ± 0.03
0.75	−	0.40 ± 0.02	1.32 ± 0.04
1.00	0.18 ± 0.01	1.06 ± 0.04	1.30 ± 0.04

−Enthalpy change not observed

The ΔH^o_{ag3} values were positive and increased with the concentration and hydrophobicity of the IL cation. According to our hypothesis, when micelles in the syringe, at very higher concentration, are introduced in the calorimetric cell just above the cmc, n(SDBS):n(C_4mimCl) is low, and IL cations can be transferred from the bulk to the micellar surface, promoting the rupture and the formation of several intermolecular interactions that promoted a favorable (or less unfavorable) enthalpic balance. This interaction included electrostatic shielding between the headgroups and specific interactions, such as π-π interactions between the aromatic rings of the surfactant and IL cation, π-cation interaction, and H-bonds. The more hydrophobic the cation, the more favorable the interactions associated with the IL cation being transferred. As the SDBS micelle concentration increased in the calorimetric cell, the free IL cation concentration

decreased, and less stable micelles were formed, without the presence of IL, promoting an increase in ΔH^o_{ag3}.

3.4. *Conclusions*

In this work, the thermodynamic of SDBS micellization was investigated in aqueous solutions with low concentrations of 1-alkyl-3-methylimidazolium chlorides ($C_n$mimCl, n = 0, 2, or 4). Fluorometric and calorimetric measurements showed that the aggregation process of the surfactant occurred in two stages: first, a transition of monomers to dimers (or small aggregates), followed by the formation of micelles.

The feasibility of the micellization process strongly depended on the cation structure of the IL, indicating that C_4mim^+ prevented the formation of small aggregates, inducing an one-step aggregation process. The aggregation was enthalpically and entropically driven for all evaluated conditions, being the thermodynamic parameters of micellization (ΔG^o_{mic}, ΔH^o_{mic}, and $T\Delta S^o_{mic}$) highly dependent on the nature and concentration of the IL in the system. The thermodynamic stability of the micelles increased in the order $C_4mim^+ > C_0mim^+ > C_2mim^+$, while the enthalpic and entropic favorableness change in the order $C_0mim^+ > C_2mim^+ > C_4mim^+$ and $C_4mim^+ > C_2mim^+ > C_0mim^+$, respectively.

We have suggested that C_nmim^+ interacted differently with the SDBS aggregates. For ILs with n ≥ 2, the interaction occurred on the surface of the aggregate via the SDBS tails within the micelles, increasing the R_H of the aggregate. However, when n = 0, although the R_H of the micelles increased, the location of the imidazolium cation was similar to that of simple electrolyte.

The scientific contribution of this work has a great impact on the understanding of micellization thermodynamics of anionic surfactants with aromatic rings in pure water and in the presence of ILs with aromatic structures.

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Capítulo 4: Solvophobic effect of 1-alkyl-3-methylimidazolium halides on the thermodynamic of complexation between β-cyclodextrin and dodecylpiridinium cation

Abstract

Preferential solvation participate in various supramolecular self-assembly processes, whose thermodynamic properties can be modulated by the addition of ionic liquids (ILs). However, the effects of these liquids on the thermodynamics of the host-guest complexation process remain unexplored. In this study, the thermodynamic properties of the complexation between 1-dodecylpiridinium cations ($C_{12}Py^+$) and β-cyclodextrin (βCD) species in aqueous solutions with different concentrations of 1-alkyl-3-methylimidazolium halides ($C_n$mimX) were investigated by isothermal titration calorimetry. In water, $C_{12}Py^+$ and βCD form a 1:1 inclusion complex, which is enthalpically ($\Delta H_{ic}^o = -9.2 \pm 0.1$ kJ mol^{-1}) and entropically ($T\Delta S_{ic}^o = 16.1 \pm 0.2$ kJ mol^{-1}) favorable. However, in IL aqueous solutions, all the βCD–$C_{12}Py^+$ thermodynamic parameters of the complexation change and this IL effect is dependent on the carbon chain length of $C_n$mim$^+$ cations. ILs with shorter alkyl chains ($C_n$mim$^+$, $n \leq 4$) decreases the system entropy, while ILs with longer alkyl chains ($C_n$mim$^+$, $n \geq 6$) reduce the enthalpy values. These effects are attributed to *i)* preferential solvation of surfactant tails by ILs; *ii)* ability of the ILs to modify the 3D water structure and *iii)* inclusion of IL molecules into the inner cavity of βCD.

Keywords: inclusion complex, ionic liquids, isothermal titration calorimetry, β-cyclodextrin, 1-dodecylpyridinium chloride.

4.1. *Introduction*

Various molecular associations of chemical species held together by weak intermolecular interactions constitute the basis for the molecular recognition field, and most of these aggregates include host-guest inclusion complexes.[1,2] Cyclodextrins are one of the most important classes of host molecules in supramolecular chemistry, whose main representative is β-cyclodextrin (βCD) due its easy synthesis and low cost.[3] βCD is a macrocyclic compound formed by seven glucopyranose units bound via α-1,4-glycosidic linkages with hydrophobic cavity, which can accommodate a large number of guest molecules forming supramolecular inclusion complexes.[4] Among various molecules capable of interacting with βCD, surfactants possess special chemical structures, which make it a strategic molecular probe to study thermodynamic inclusion process with βCD.[5,6]

The formation of βCD–surfactant complexes is driven by weak non-covalent interactions such as hydrophobic, electrostatic, and van der Waals ones as well as by hydrogen bonds (in addition to steric effects).[7] They involve not only direct interactions between the surfactant and βCD species, but also their interaction with the solvent.[8–12] Thus, changing the solvent property by the addition of cosolutes or cosolvents can affect the thermodynamics of the complexation interaction between βCD and surfactants.[13,14] Ionic liquids (ILs) are a class of cosolutes that are able to strongly modify the process of molecular association such as surfactant micellization,[15–17] polymer-surfactant association,[18] and protein aggregation.[19,20] They consist of organic and semi-organic salts existing in the liquid state at temperatures below the water boiling point.[21] The versatility of ILs results from the high number of possible combinations of cation-anion pairs (over 10 trillion), allowing their application in different fields of science.[22]

Among various types of ILs, the imidazolium cation-based ILs containing an imidazolium ring with delocalized charge are the most widely studied due to their easy synthesis, high stability under oxidative and reductive conditions,[23] and ability to affect several cooperative processes. For instance, 1-alkyl-3-methylimidazolium cation ($C_n mim^+$) has been examined as a cosolute in different systems containing molecules with high aggregation capacity. Ferreira et al.[18] investigated the effect of 1-butyl-3-methylimidazolium halides on the interactions between poly(ethylene oxide) (PEO) and sodium dodecyl sulfate (SDS) in aqueous solutions using fluorescence spectroscopy, electrical conductivity, and isothermal titration calorimetry (ITC) techniques. The authors

showed that the presence of IL affected the relative stabilities of SDS–PEO nanostructures and SDS micelles, making the former aggregates less stable. Ali et al.[24] studied quaternary systems containing water, 1-decyl-3-methylimidazolium chloride, cetylpyridinium chloride, and cetylpyridinium bromide by the conductimetry and tensiometry techniques and concluded that the IL cations acted as a co-surfactant forming micelles containing cetylpyridinium. Figueiredo et al.[25] evaluated the effects of 1-ethyl- and 1-butyl-3-methylimidazolium chlorides on the denaturation temperature (T_d) of alpha-helix Im7 protein by differential scanning calorimetry. They found that the IL with a larger lateral chain promoted more intense decrease in protein denaturation process, indicating the existence of favorable protein-IL interactions.

Because of the strong effect of imidazole ILs on various cooperative processes, we expect that their presence can modulate possible interactions between βCD and the 1-dodecylpiridinium cation, $C_{12}Py^+$. The formation of βCD–$C_{12}Py^+$ complexes has been reported in only 3 works. Satake et al.[26] obtained the binding constant (K_{ic} = 4980 L mol^{-1}) for βCD–$C_{12}Py^+$ complexation by performing conductivity measurements in pure water at 298.15 K. Czapkiewicz et al.[27] evaluated the effect of 0.1 mol L^{-1} NaCl solution on the βCD–$C_{12}Py^+$ complex formation by conducting surface tension studies at 298.15 K. The calculated K_{ic} magnitude was equal to 15600 L mol^{-1}, which was approximately three times greater than the value obtained by Satake et al. in the absence of NaCl (the observed phenomenon was explained by the so-called NaCl salting-out effect). Beiginejad et al.[28] conducted a complete thermodynamic evaluation of the formation of βCD–$C_{12}Py^+$ complex in pure water by performing conductivity measurements at 298.15 K. The obtained K_{ic}, standard enthalpy change (ΔH_{ic}^o), standard Gibbs free energy change (ΔG_{ic}^o), and standard entropy change (ΔS_{ic}^o) values were 17220 L mol^{-1}, –41.59 kJ mol^{-1}, –24.18 kJ mol^{-1}, and –58 J mol^{-1} K^{-1}, respectively. Due to the discrepancies between the K_{ic} values reported in the literature, the indirect determination of ΔH_{ic}^o (van't Hoff approach) and the existence of only one[28] study focused on the thermodynamic properties of βCD–$C_{12}Py^+$ inclusion, more thermodynamics studies is need to understand this molecular association process.

In this work, using ITC, all thermodynamic parameters associated with the formation of βCD–$C_{12}Py^+$ inclusion complex (K_{ic}, ΔG_{ic}^o, ΔH_{ic}^o, and $T\Delta S_{ic}^o$) were directly determined, as well as, the effect of 1-alkyl-3-methylimidazolium halides (whose cation

structure is depicted in Figure B.S1 in the supporting information) on this inclusion process.

4.2. *Experimental*

4.2.1. Materials

βCD (≥ 97.0%), 1-dodecylpyridinium chloride hydrate (98.0%), 1-methylimidazolium chloride (95.0%), 1-ethyl-3-methylimidazolium chloride (≥ 95.0%), 1-butyl-3-methylimidazolium chloride (≥ 98.0%), 1-butyl-3-methylimidazolium bromide (> 97.0%), 1-butyl-3-methylimidazolium iodide (99.0%), 1-hexyl-3-methylimidazolium chloride (≥ 97.0%), and 1-octyl-3-methylimidazolium chloride (≥ 97.0%) were purchased from Sigma-Aldrich (USA). Sodium chloride (99.0%) was supplied by Isofar (Brazil). All chemicals were used without further purification. Milli-QII water (Millipore, USA) was used to prepare all solutions.

4.2.2. ITC studies

Titration calorimetric measurements were performed using a thermal activity monitor (TAM III) equipped with 4.0 mL sample and reference stainless steel cells. This setup manufactured by TA instruments (USA) was controlled by the TAM assistant™ software. Experiments consisted of 50 and 100 consecutives 5 μL injections of concentrated C_{12}PyCl solution into the sample cell containing 2.7 mL of solvent (pure water or water + IL mixture) with or without 2.1 mmol L^{-1} βCD at 298.15000 ± 0.00001 K. The reference cell was filled with 2.7 mL of the solvent present in the sample cell. Surfactant solution was prepared from the same solvent, and titrations were performed using a Hamilton syringe (500 μL) controlled by the 3810 syringe pump of the TA Instruments setup. The time interval between two consecutive injections was equal to 900 s, and the solution inside the sample cell was stirred at a speed of 180 rpm (3.0 s^{-1}) during all experiments using a gold helix stirrer. Data were prepared for analysis by subtracting the heat of dilution (measured during the titration of the C_{12}PyCl solution into a solvent) from the heat released or absorbed during the titration of the C_{12}PyCl solution into the βCD solution (the titration of the solvent into the βCD solution produced negligible thermal effects).

Figure 4.1 shows the thermogram obtained from the addition of the C_{12}PyCl solution into the βCD aqueous solution at 298.15 K.

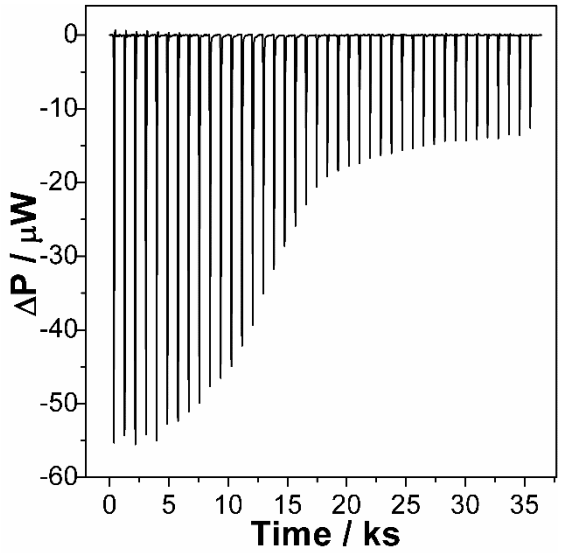

Figure 4.1. Thermal power as a function of time plotted for the sequential 5 μL injections of 66.4 mmol L^{-1} C_{12}PyCl into 2.7 mL of 2.1 mmol L^{-1} βCD aqueous solution at 298.15 K.

The enthalpy change during each injection was obtained by dividing the corresponding peak area by the number of moles of added C_{12}PyCl.

4.2.3. Analysis of ITC results

The association process can be described by the following chemical equation:

$$\beta CD + C_{12}Py^{+} \rightleftharpoons \beta CD - C_{12}Py^{+} \qquad 4.1$$

where βCD and $C_{12}Py^{+}$ are free chemical species, and $\beta CD - C_{12}Py^{+}$ is the complex formed at the equilibrium. The obtained experimental data were fitted with the binding of a solute to one macromolecule with identical and independent sites model[29] using the equation:

$$Q = \frac{V_c \Delta H_{ic}^o}{2K_{ic}}\left[1 + K_{ic}[C_{12}Py^+]_T + nK_{ic}[\beta CD]_T - \sqrt{(1 + K_{ic}[C_{12}Py^+]_T + nK_{ic}[\beta CD]_T)^2 - 4n{K_{ic}}^2[\beta CD]_T[C_{12}Py^+]_T}\right] \quad 4.2$$

where Q is the total heat content; V_c represents the effective volume inside the sample calorimetric cell; $[\beta CD]_T$ and $[C_{12}Py^+]_T$ are the total concentrations of βCD and $C_{12}Py^+$ species in the system, respectively; and n is the stoichiometric ratio. By fitting the ITC data to equation 4.2, values of K_{ic} and ΔH_{ic}^o were obtained for the formation of βCD−$C_{12}Py^+$ inclusion complex. The magnitude of K_{ic} was used to calculate the ΔG_{ic}^o values for the complex formation according to equation 4.3:

$$\Delta G_{ic}^o = -RTlnK_{ic} \quad 4.3$$

where R is the universal gas constant, and T is the absolute temperature of the system. From the magnitudes of ΔG_{ic}^o and ΔH_{ic}^o, $T\Delta S_{ic}^o$ values were calculated using the Gibbs fundamental relation:

$$\Delta G_{ic}^o = \Delta H_{ic}^o - T\Delta S_{ic}^o \quad 4.4$$

4.3. *Results and Discussion*

4.3.1. βCD−$C_{12}Py^+$ complexation in pure water

To elucidate the effect of ILs on the formation of βCD−$C_{12}Py^+$ inclusion complex, it is necessary to investigate the thermodynamics of βCD−$C_{12}Py^+$ complexation in the absence of those cosolutes. ITC is a very accurate and reproducible experimental technique, which is able to determine all thermodynamic parameters ($n, K_{ic}, \Delta G_{ic}^o, \Delta H_{ic}^o$, and $T\Delta S_{ic}^o$) associated with the formation of βCD−$C_{12}Py^+$ inclusion complex by conducting a single experiment.[30]

Figure 4.2 shows the enthalpy change observed (ΔH_{obs}) during the titration of 66.4 mmol L^{-1} $C_{12}PyCl$ aqueous solution in pure water and in 2.1 mmol L^{-1} βCD aqueous solution, which is plotted as a function of the molar ratio between $C_{12}Py^+$ and βCD species ($r = [C_{12}Py^+]/[\beta CD]$).

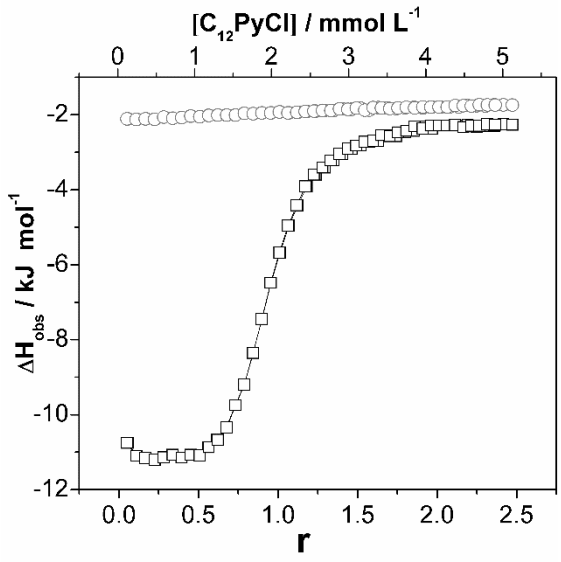

Figure 4.2. ΔH_{obs} values measured during the titration of 66.4 mmol L^{-1} $C_{12}PyCl$ in (○) pure water and (□) 2.1 mmol L^{-1} βCD aqueous solution at a temperature of 298.15 K and various molar ratios between $C_{12}Py^{+}$ and βCD species. The $C_{12}PyCl$ concentration is denote by the top axis.

For both titrations, the ΔH_{obs} values remained negative at all r magnitudes; however, the resulting curves exhibited different profiles, demonstrating that the surfactant interacted with the macrocyclic compound. When the solution of $C_{12}PyCl$ was added to the calorimetric sample cell containing the βCD solution, the resulting values of ΔH_{obs} reflected the existence of a complex energetic balance involving the disruption and formation of various intermolecular interactions. Whereas the surfactant concentration in the sample cell was even lower than the critical micelle concentration (cmc) of the $C_{12}PyCl$ solution (16.2 mmol L^{-1}),[31] the ΔH_{obs} values obtained for the surfactant titration into the βCD solution can be expressed by the following sum:

$$\Delta H_{obs} = \Delta H_{mic}^{dil} + \Delta H_{demic} + \Delta H_{int} \qquad 4.5$$

where ΔH_{mic}^{dil} is the enthalpy change for the micelle dilution; ΔH_{demic} is the enthalpy change associated with the demicellization process, in which micelles are broken to produce free monomers of surfactant in solution; and ΔH_{int} is the enthalpy change due to βCD−$C_{12}Py^{+}$ interactions. The same terms of equation 4.5 (except for ΔH_{int}) were applicable when $C_{12}PyCl$ was added to pure water. Therefore, the subtraction of each ΔH_{obs} magnitude obtained for the surfactant dilution in pure water from the ΔH_{obs} values

determined in the presence of βCD produced an apparent enthalpy change due to the interaction between βCD and $C_{12}Py^+$ species (ΔH_{ap-int}). We used the term "apparent" here because the amount of monomers complexing with βCD after each injection was unknown. Figure 4.3 shows the ΔH_{ap-int} values plotted as a function of r for the βCD−$C_{12}Py^+$ complexation in pure water at 298.15 K.

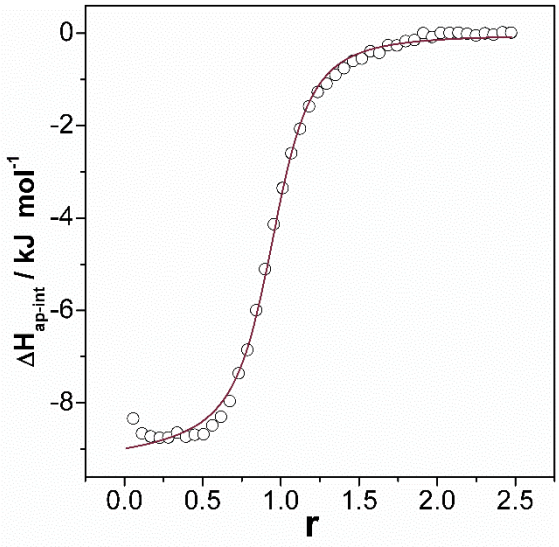

Figure 4.3. ΔH_{ap-int} values plotted as a function of r for the βCD−$C_{12}Py^+$ complexation in pure water at 298.15 K. (-) The non-linear fitting model used to obtain the n, K_{ic} and ΔH_{ic}^o parameters.

The resulting ΔH_{ap-int} versus r curve with a sigmoidal profile was fitted to the binding model (equation 4.2) to obtain the values of n, K_{ic}, ΔG_{ic}^o, ΔH_{ic}^o, and $T\Delta S_{ic}^o$.

The negative inclusion standard Gibbs free energy change (ΔG_{ic}^o = −25.2 ± 0.2 kJ mol^{-1}) indicates that in the aqueous mixture of βCD and $C_{12}PyCl$ formed at the equilibrium conditions, the concentration of βCD−$C_{12}Py^+$ complex dominated over those of the dissociated reactants, with the association process being enthalpically (ΔH_{ic}^o = −9.2 ± 0.1 kJ mol^{-1}) and entropically ($T\Delta S_{ic}^o$ = 16.1 ± 0.2 kJ mol^{-1}) favorable.

The observed changes in the thermodynamic parameters of the inclusion process are affected by three different subprocesses that occur simultaneously during βCD−$C_{12}Py^+$ association. Therefore, the system enthalpy change due to the association process, ΔH_{ic}^o, can be written as the following sum:

$$\Delta H_{ic}^o = \Delta H_{des}^{o,Surf} + \Delta H_{des}^{o,\beta CD} + \Delta H_{int}^{o,\beta CD-Surf} \quad 4.6$$

where $\Delta H_{des}^{o,Surf}$ and $\Delta H_{des}^{o,\beta CD}$ are the standard enthalpy changes associated with the desolvation of $C_{12}Py^+$ and βCD species, respectively; and $\Delta H_{int}^{o,\beta CD-Surf}$ is the standard enthalpy change related to the formation of βCD−$C_{12}Py^+$ complex. In previous studies, it was found that during the complexation between macrocyclic compounds (such as βCD) and surfactants, the hydrophobic chain of the surfactant was inserted inside the βCD hydrophobic cavity, promoting the release of water molecules from the hydrophobic regions of both βCD and surfactant species to the solution bulk.[32] The water molecules solvating the surfactant hydrophobic chains interact more intensely with one another via hydrogen bonds as compared to those in the bulk, resulting in a positive value of $\Delta H_{des}^{o,Surf}$.[33] On the other hand, although the βCD cavities exhibit hydrophobic properties, the magnitude of $\Delta H_{des}^{o,\beta CD}$ is negative because the geometric confinement of water molecules inside the cavities decreases their orbital orientation degree, causing a reduction in the strengths of water-water hydrogen bonds as compared with those in the bulk.[34] The term $\Delta H_{int}^{o,\beta CD-Surf}$ is negative mainly due to the existence of dispersive (van der Waals) forces, hydrogen bonds, and hydrophobic interactions between βCD and $C_{12}Py^+$ species. Because the association process in this study is exothermic, it can be concluded that $\left|\Delta H_{des}^{o,\beta CD} + \Delta H_{int}^{o,\beta CD-Surf}\right| > \left|\Delta H_{des}^{o,Surf}\right|$, indicating that the balance between the energy expended to remove water molecules from the surfactant tail solvation shell and the energies released during the formation of βCD−$C_{12}Py^+$ complex and desolvation of βCD species determines the sign and magnitude of ΔH_{ic}^{o}.

A similar approach can be applied to discussing entropic contributions to the βCD−$C_{12}Py^+$ complexation process. The corresponding $T\Delta S_{ic}^{o}$ values are expressed as the sum of the following three terms:

$$T\Delta S_{ic}^{o} = T\Delta S_{rel}^{o,H2O-Surf} + T\Delta S_{rel}^{o,H2O-\beta CD} + T\Delta S_{tras-rota-conf}^{o,\beta CD-Surf} \quad 4.7$$

where $T\Delta S_{rel}^{o,H2O-Surf}$ and $T\Delta S_{rel}^{o,H2O-\beta CD}$ are the standard entropy changes caused by releasing water molecules from the solvation shells of $C_{12}Py^+$ chain and βCD cavity to the solution bulk, respectively; and $T\Delta S_{tras-rota-conf}^{o,\beta CD-Surf}$ is the standard entropy change due to the variations of the number of translational-rotational-conformational degrees of freedom observed for both the surfactant and desolvated βCD molecules after the

formation of βCD−$C_{12}Py^+$ complex. The hydration of the nonpolar parts of surfactants and βCD result in different entropic contributions: while the water molecules solvating the non-polar surfactant chain are more structured than those in the bulk,[35] the water molecules inside the βCD cavity are structured to a lower degree. The βCD cavity volume is large enough to store up to eleven water molecules; however, it typically accommodates only six or seven of them, providing larger translational and rotational degrees of freedom.[36] As a consequence, the water molecules inside the βCD cavity have a higher partial molar entropy as compared to that of the water molecules present in the solution bulk. Thus, the term $T\Delta S_{rel}^{o,H20-Surf}$ is a positive one, whereas the sign of $T\Delta S_{rel}^{o,H2O-\beta CD}$ is negative. $C_{12}Py^+$ and βCD species have higher magnitudes of the translational, rotational, and conformational entropies in their free state as compared with those in their complexed state, making $T\Delta S_{tras-rota-conf}^{o,\beta CD-Surf}$ values negative. Therefore, the favorable entropic contribution for the βCD−$C_{12}Py^+$ association process mainly resulted from the entropic gain of the water molecules released from the surfactant tail.

The thermodynamic parameters of the association process between βCD and $C_{12}Py^+$ at 298.15 K were previously obtained by Beiginejad et al.[28] using a conductometric method and van't Hoff approach. While the ΔG_{ic}^o value calculated in Ref.[28] (−24.2 kJ mol^{-1}) was very close to that obtained in this work (−25.2 ± 0.2 kJ mol^{-1}), their ΔH_{ic}^o (−41.6 kJ mol^{-1}) and $T\Delta S_{ic}^o$ (−17.4 kJ mol^{-1}) magnitudes were very different from ours (ΔH_{ic}^o = −9.2 ± 0.1 kJ mol^{-1} and $T\Delta S_{ic}^o$ = 16.1 ± 0.2 kJ mol^{-1}). In contrast to our assessment, the authors of the cited study concluded that the driving force for the βCD−$C_{12}Py^+$ association process was only enthalpic. The large discrepancy between the ΔH_{ic}^o and $T\Delta S_{ic}^o$ values obtained in both works can be explained by the fact that in the Beiginejad paper's, the ΔH_{ic}^o magnitude was determined indirectly using the van't Hoff approach, whereas in our work, the enthalpy change was measured directly by ITC. Generally, for the thermodynamic processes that occur in multiple steps with very close entropies of different states, $\Delta H_{ic}^{o,ITC} \neq \Delta H_{ic}^{o,van't\ Hoff}$; On the other hand, their magnitudes are equal only when the system change proceeds in one step with a finite entropy difference between the two thermodynamic states.[37]

4.3.2. Effect of imidazolium ILs on βCD−$C_{12}Py^+$ complexation

The thermodynamic parameters of βCD−$C_{12}Py^+$ association determined in this work showed the important role of the $C_{12}Py^+$ tail solvation by water molecules during the complexation process, in which the release of water increased both the enthalpy and entropy of the system. Several cosolutes and cosolvents,[38–41] including imidazolium ILs,[42,43] can modify the three-dimensional structure of water molecules solvating the hydrophobic regions and/or solution bulk, thus promoting changes in the enthalpic and entropic contributions associated with the host-guest complexation processes. To elucidate the effect of imidazolium ILs on the βCD−$C_{12}Py^+$ association, we have determined the thermodynamic parameters of the supramolecular structure formation process in the presence of ILs with different concentrations.

Figure 4.4a shows the effect of 1-butyl-3-methylimidazolium chloride (C_4mimCl) on the ΔH_{obs} values plotted as functions of r at IL concentrations of 10.0 and 100.0 mmol L^{-1}.

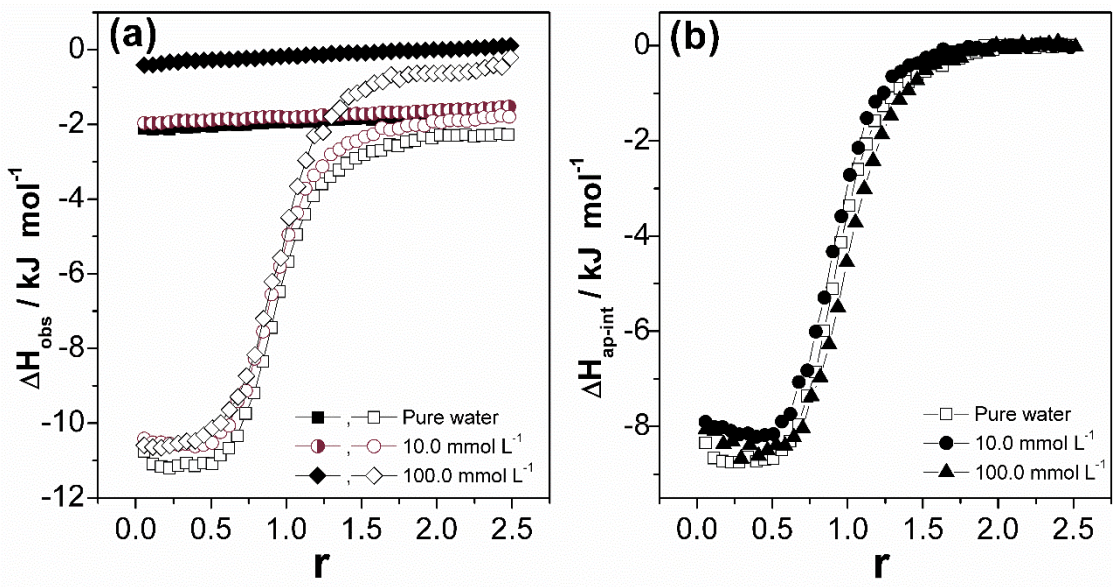

Figure 4.4. **(a)** ΔH_{obs} plotted as a function of r during the titration of C_{12}PyCl in different solvents: *i*) pure water or water + IL (full symbols) and *ii*) 2.1 mmol L^{-1} βCD + water or 2.1 mmol L^{-1} βCD + water + IL (open symbols). **(b)** Plots of ΔH_{ap-int} *versus* r obtained for the βCD−$C_{12}Py^+$ complexation in different solvents. The surfactant dilution curves are depicted as functions of the molar ratio instead of the surfactant concentration because this format is more suitable for the comparison with the interaction curves.

The increase in the C_4mimCl concentration up to 100.0 mmol L^{-1} did not affect the surfactant dilution and formation of βCD−$C_{12}Py^+$ complexes as indicated by the similarities between the obtained sigmoidal and linear curves (see also the curves of ΔH_{ap-int} *versus* r depicted in Figure 4.4b). On the other hand, at C_4mimCl concentrations above 200.0 mmol L^{-1}, a new process was detected during surfactant dilution (Figure 4.5).

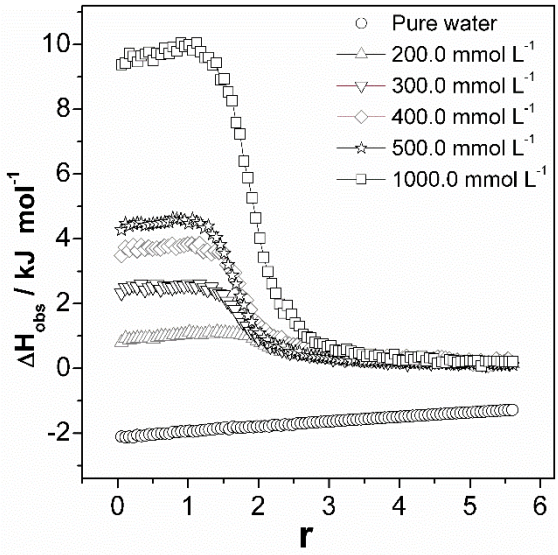

Figure 4.5. Plots of ΔH_{obs} *versus* r obtained for the dilution of C_{12}PyCl at various C_4mimCl concentrations and temperature of 298.15 K.

While the enthalpy change measured for the dilution of C_{12}PyCl in pure water exhibited a linear increase with r, at C_4mimCl concentrations above 200.0 mmol L^{-1}, the dilution curves possessed sigmoidal profiles, indicating that the IL addition induced the cooperative interaction process (micellization) of C_{12}PyCl species. This ability of ILs to reduce the *cmc* of surfactants has been investigated previously in the literature.[24,44] At C_4mimCl contents of 0, 10.0, and 100.0 mmol L^{-1}, this process occurred at C_{12}PyCl concentrations higher than the value used for evaluating the complexation reaction of C_{12}PyCl with βCD and, therefore, was not discussed.

The C_{12}PyCl dilution curves obtained at C_4mimCl concentrations higher than 100.0 mmol L^{-1} consist of three distinct regions: the pre-micellar region where only free monomers exist ($r < 1.3$). ΔH_{obs} magnitudes measured in the pre-micellar region increase with increasing C_4mimCl content, and it's values are affected by the following three processes: *i*) micelle dilution, *ii*) breakage of micelles, *iii*) solvation of surfactant

monomers. The process *iii* produces the largest effect on the ΔH_{obs} magnitude at various IL concentrations since the solvation of the surfactant chain depends on the solvent composition. In pure water, the pre-micellar region exhibited negative ΔH_{obs} values (around −2 kJ mol^{-1}) because the hydrogen bonds formed between the water molecules surrounding the surfactant tail are stronger than those between the water molecules in the solution bulk. When the solvent was a mixture of water with IL, the ΔH_{obs} values in the pre-micellar region were positive (around 1 and 10 kJ mol^{-1} at C_4mimCl contents of 200.0 and 1000.0 mmol L^{-1}, respectively), indicating that C_4mimCl species competed with water molecules during the solvation of the surfactant tail. As the interactions between C_4mimCl and the surfactant tail are weaker than the water−IL and IL−IL ones, the solvation of $C_{12}Py^+$ hydrophobic chains by C_4mimCl increases the enthalpy of the system, making the magnitudes of ΔH_{obs} positive. Therefore, since the increase in the IL concentration decreased the number of water molecules in the surfactant solvation shell, the ΔH_{obs} values in the pre-micellar region became more positive. At $1.3 < r < 3$, an abrupt decrease in ΔH_{obs} occurred due to the onset of the C_{12}PyCl micelles formation process. At $r > 3$ corresponding to the post-micellar region, in which micelles were diluted, the ΔH_{obs} values became constant again. In this area, the system is characterized by a dynamic equilibrium, at which monomers and micelles coexist with each other. The standard enthalpy change for the micellization process (ΔH^o_{mic}) was obtained from the difference $\Delta H^o_{mic} = \Delta H^{o,pos}_{obs} - \Delta H^{o,pre}_{obs}$, where the $\Delta H^{o,pos}_{obs}$ and $\Delta H^{o,pre}_{obs}$ terms were calculated by the forward extrapolation of the pre-micellar and post-micellar linear curves to the inflection point (cmc), respectively, as was suggested by Loh et al.[45] ΔH^o_{mic} is the thermodynamic parameter that represents the net energy of the intermolecular interactions that occur when a mole of surfactant changes its state from the monomeric state to the aggregate one. In pure water, the ΔH^o_{mic} values of C_{12}PyCl obtained between 293.15 and 303.15 K are positive, and the main molecular process that affects their magnitudes is the rupture of water-water interactions around the surfactant tail in the monomeric state.[46] In the water–IL mixtures, the ΔH^o_{mic} values became more negative with increasing IL concentration and reached the magnitudes of –0.9 ± 0.1 and –9.8 ± 0.4 kJ mol^{-1} at C_4mimCl concentrations of 200.0 and 1000.0 mmol L^{-1}, respectively (Table 4.1). These results indicate that the tails of free monomers were solvated by IL molecules since the breakage of water-water interactions in the hydrophobic regions represented a more endothermic process as compared to breaking the hydrophobic interactions between the

surfactant and C_4mimCl species. Figure 4.6 shows the effect of the C_4mimCl concentration on the curves of ΔH_{obs} *versus* r obtained by the titration of 66.4 mmol L^{-1} C_{12}PyCl into 2.1 mmol L^{-1} βCD solution.

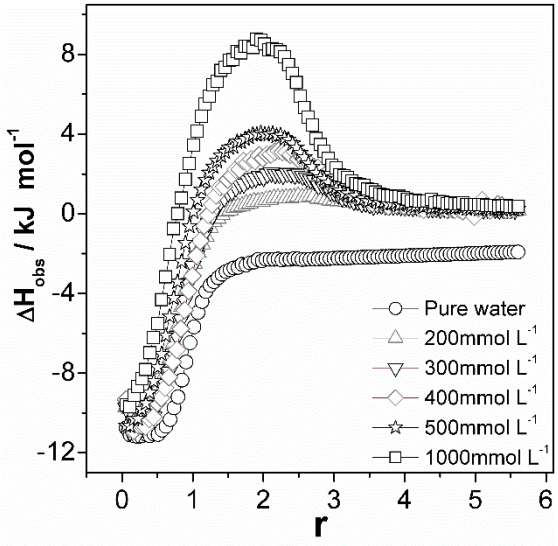

Figure 4.6. Plots of ΔH_{obs} *versus* r obtained by the titration of C_{12}PyCl into 2.1 mmol L^{-1} βCD + water + C_4mimCl mixture at IL concentrations above 200 mmol L^{-1} and temperature of 298.15 K.

For the titration of C_{12}PyCl into βCD + C_4mimCl + water mixtures, the corresponding ΔH_{obs} *versus* r plots consist of two sigmoidal parts: a growing and decreasing one. Such complex profiles indicate that at higher C_4mimCl contents, the inclusion of $C_{12}Py^+$ ions into βCD cavities and auto-aggregation of surfactant molecules overlap at some concentrations during titration. However, both these processes can be evaluated independently.

To test our hypothesis, the standard enthalpy change (ΔH_{2nd}^{o}) associated with the second process observed at higher r in Figure 4.6 was calculated from the difference between two ΔH_{obs} values ($\Delta H_{2nd}^{o} = \Delta H_{obs}^{o,2} - \Delta H_{obs}^{o,1}$, where the magnitudes of $\Delta H_{obs}^{o,1}$ and $\Delta H_{obs}^{o,2}$ were obtained by extrapolating r in the decreasing sigmoidal part to 0 and ∞, respectively). Table 4.1 lists the ΔH_{2nd}^{o} values determined at C_4mimCl concentrations between 200.0 and 1000.0 mmol L^{-1}, which are very close to the ΔH_{mic}^{o} magnitudes calculated using the surfactant dilution curves depicted in Figure 4.5, confirming that the second process in the titration curve involving βCD was the micellization of C_{12}PyCl

surfactant that occurred after the formation of the inclusion complex from $C_{12}Py^+$ and βCD species.

Table 4.1. Micellization enthalpy changes obtained for C_{12}PyCl in the presence and absence of βCD at various C_4mimCl concentrations.

C_4mimCl	**surfactant**	**after inclusion**
mmol L^{-1}	$\Delta \boldsymbol{H}^{o}_{\boldsymbol{mic}}$ / kJ mol^{-1}	$\Delta \boldsymbol{H}^{o}_{\boldsymbol{2nd}}$ / kJ mol^{-1}
200.0	−0.9 ± 0.1	−0.9 ± 0.2
300.0	−2.2 ± 0.1	−2.3 ± 0.3
400.0	−3.3 ± 0.1	−3.2 ± 0.2
500.0	−4.3 ± 0.1	−4.3 ± 0.2
1000.0	−9.8 ± 0.4	−9.9 ± 0.2

When the βCD–$C_{12}Py^+$ complexation occurs in the surfactant concentration range corresponding to the micellization of C_{12}PyCl without βCD, it is necessary to take into account the real ΔH_{ap-int} *versus* r curves associated with only complexation process. Otherwise, the direct subtraction of the ΔH_{obs} values obtained via surfactant dilution from the ΔH_{obs} magnitudes determined from the C_{12}PyCl titration for each r in the presence of βCD does not produce the correct ΔH_{ap-int} parameters. The result of this direct subtraction is affected by the energy change caused by the micelle formation in the absence of βCD (Figure 4.5) at surfactant concentrations corresponding to the inclusion process (Figure 4.6).

To obtain the correct values of ΔH_{ap-int}, it should be considered that $\Delta H_{ap-int} = \Delta H^{\beta CD}_{obs\,(i)} - \Delta H^{*,dil}_{obs\,(i)}$, where $\Delta H^{\beta CD}_{obs\,(i)}$ is the observed enthalpy change during the i-th injection of C_{12}PyCl into the βCD solution, and $\Delta H^{*,dil}_{obs\,(i)}$ is the enthalpy change calculated via the following equation:

$$\Delta H^{*,dil}_{obs\,(i)} = \Delta H^{dil}_{obs\,(i)} + \left(|\Delta H^{o}_{mic}|.\, X_{mic_{(i)}}\right) \qquad 4.8$$

where $\Delta H^{dil}_{obs\,(i)}$ is the observed enthalpy change after the i-th injection of C_{12}PyCl into the solvent (see the dilution curves in Figure 4.5), and $X_{mic_{(i)}}$ is the fraction of C_{12}PyCl monomers that formed micelles in the calorimetric cell after the same injection. The magnitude of $X_{mic_{(i)}}$ was calculated as follows:

$$X_{mic_{(i)}} = \frac{\Delta H_{obs\,(i)} - \Delta H^{pre}_{obs\,(i)}}{\left|\Delta H^{o}_{mic}\right|} \quad 4.9$$

where $\Delta H^{pre}_{obs\,(i)}$ is the enthalpy change obtained by extrapolating each r value of the linear fit to the points on the surfactant dilution curve in the pre-micellar region (see Figure B.S2 in the supporting information).

Using this method, curves of ΔH_{ap-int} *versus* r were obtained, and the thermodynamic parameters determined at 298.15 K for the formation of the inclusion complex from βCD and $C_{12}Py^+$ species in the presence of C_4mimCl were calculated (Table 4.2).

Table 4.2. Thermodynamic parameters of the βCD–$C_{12}Py^+$ complexation process obtained at different C_4mimCl concentrations.

C_4mimCl	$\boldsymbol{K_{ic}}$	$\boldsymbol{\Delta G^o_{ic}}$	$\boldsymbol{\Delta H^o_{ic}}$	$\boldsymbol{T\Delta S^o_{ic}}$	$\boldsymbol{n}$
mmol L^{-1}	10^4. L mol^{-1}	kJ mol^{-1}			
0.0	2.6 ± 0.2	−25.2 ± 0.2	−9.2 ± 0.1	16.1 ± 0.2	0.94 ± 0.01
2.50	3.3 ± 0.3	−25.8 ± 0.2	−8.1 ± 0.1	17.7 ± 0.2	0.94 ± 0.01
10.0	3.1 ± 0.3	−25.6 ± 0.2	−8.5 ± 0.0	17.1 ± 0.2	0.91 ± 0.01
100.0	3.0 ± 0.2	−25.6 ± 0.2	−9.3 ± 0.1	16.2 ± 0.2	0.93 ± 0.01
200.0	1.6 ± 0.1	−24.0 ± 0.1	−10.5 ± 0.1	13.5 ± 0.1	0.85 ± 0.01
300.0	1.8 ± 0.2	−24.3 ± 0.3	−10.7 ± 0.2	13.6 ± 0.4	0.87 ± 0.01
400.0	1.9 ± 0.2	−24.4 ± 0.2	−11.7 ± 0.1	12.7 ± 0.3	0.92 ± 0.01
500.0	0.66 ± 0.03	−22.4 ± 0.1	−16.3 ± 0.1	6.1 ± 0.1	0.91 ± 0.01
1000.0	0.56 ± 0.04	−21.4 ± 0.2	−23.2 ± 0.4	−1.8 ± 0.4	0.87 ± 0.01

At all examined C_4mimCl concentrations, the obtained n values were near unity, indicating that the 1:1 stoichiometry of βCD–$C_{12}Py^+$ complex was maintained in the presence of C_4mimCl. After the C_4mimCl concentration increased to 1000.0 mmol L^{-1}, the corresponding ΔG^o_{ic} values became less negative, indicating a decrease in the number of complex species formed at the same contents of βCD and $C_{12}Py^+$ within the system. Despite the high enthalpic energy released during the complex formation with increasing IL concentration, the observed decrease in the complex stability was due to the entropic reduction that made the inclusion process entropically unfavorable at a C_4mimCl content of 1000.0 mmol L^{-1}.

The ability of C_4mimCl to make the enthalpy change of the βCD–$C_{12}Py^+$ complex formation process more negative and the related entropy change less positive can be attributed to the kosmotropic effect of C_4mim^+ and Cl^- ions and preferential solvation of

$C_{12}Py^+$ species by the IL. It is well known that Cl^- and C_4mim^+ are kosmotropic ions that intensify the hydrogen bonds formed between water molecules, making them more ordered in the solution bulk.[47,48] Thus, the kosmotropic effect reduces both the partial molar entropy and partial molar enthalpy of the water molecules in the bulk, which decreases the contribution of the water released from the solvation shell of $C_{12}Py^+$ and βCD to the entropic gain (as indicated by less positive values of $T\Delta S_{ic}^o$) and enthalpic loss (corresponding to more negative values of ΔH_{ic}^o) during the inclusion process. At the same time, in the solvent composed of C_4mimCl and water, the species that can form supramolecular complexes (such as $C_{12}Py^+$ and βCD) undergo direct intermolecular interactions with either water or IL molecules (or both) in the first layers of the solvation shell. The results obtained in this study suggest that when the IL concentration in the system increased, the competition between IL and water molecules was promoted by solvating the surfactant tail; as result, a smaller number of structured water molecules were released from the $C_{12}Py^+$ hydrophobic region to the solution bulk after the formation of βCD–$C_{12}Py^+$ complex. Thus, the terms $\Delta H_{des}^{o,Surf}$ and $T\Delta S_{rel}^{o,H20-Surf}$ in the equations 4.6 and 4.7, respectively, became less positive in the presence of the IL, contributing to more negative ΔH_{ic}^o and $T\Delta S_{ic}^o$ values.

4.3.3. Effect of cosolute cation nature on βCD–$C_{12}Py^+$ association

After examining how the C_4mimCl concentration affected the thermodynamic properties (ΔG_{ic}^o, ΔH_{ic}^o, and $T\Delta S_{ic}^o$) of the formation of βCD–$C_{12}Py^+$ inclusion complex, we decided to investigate the effects produced by the cation nature (C_nmim^+ or Na^+) and length of the IL alkyl chain ($C_n$mimCl, n = 0, 2, 4, 6, or 8) on the relative contributions of the electrostatic and hydrophobic interactions to βCD–$C_{12}Py^+$ association. To determine the contribution of cosolutes to this process, the approach used in previous sections was applied, and the thermograms (Figures B.S3 – B.S13) and tables (Tables B.S1 and B.S2) containing the obtained thermodynamic parameters of inclusion can be found in the supporting information.

Figure 4.7 shows the ΔG_{ic}^o values determined for the association process of βCD and $C_{12}Py^+$ species as a function of the cosolute concentration. For the C_6mimCl and C_8mimCl ILs, only concentrations below or equal to 100.0 mmol L^{-1} and 10.0 mmol L^{-1}, respectively, were investigated because it was not possible to apply the theoretical model

described in equation 4.2 to the obtained experimental curves at higher concentrations of these ILs.

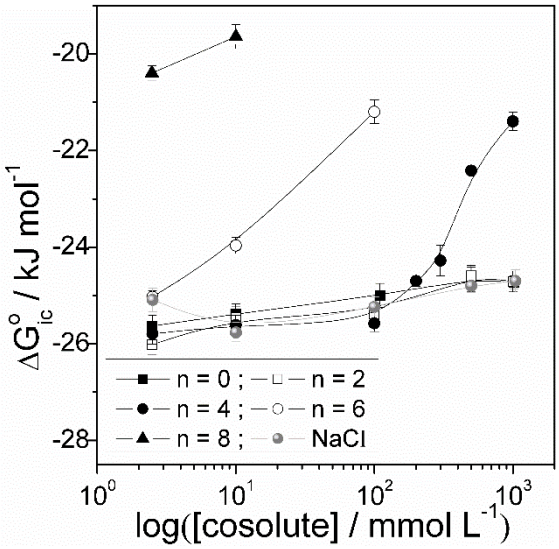

Figure 4.7. Plots of ΔG^o_{ic} *versus* cosolute concentration obtained for the inclusion of $C_{12}Py^+$ ions into βCD cavities in the aqueous solution of $C_n$mimCl (n = 0, 2, 4, 6 or 8) and NaCl at 298.15 K. The cosolute concentration is plotted on a logarithmic scale for better visualization of the results.

Regardless of the cation composition, the ΔG^o_{ic} values were negative at all cosolute concentrations, indicating that the βCD–$C_{12}Py^+$ inclusion complex dominated over free $C_{12}Py^+$ and βCD species concentration in the equilibrium state. However, the thermodynamic stability of βCD–$C_{12}Py^+$ complex decreased with an increase in the IL concentration (this effect was more pronounced at higher lengths of the alkyl chain of the $C_n$mimCl IL). At all cosolute concentrations, the effect of NaCl on the stability of βCD–$C_{12}Py^+$ inclusion complex was very similar to those produced by the C_0mimCl and C_2mimCl ILs. However, a similarity between the effects of NaCl and C_4mimCl was observed only at a cosolute concentration of 100 mmol L^{-1}, whereas the complex stability decreased in the presence of C_6mimCl and C_8mimCl much more intensely than in the presence of NaCl, C_0mimCl, C_2mimCl, and C_4mimCl. These results suggested that the effects of ILs on the $C_{12}Py^+$ inclusion into βCD cavities were modulated by hydrophobic interactions.

While we were unable to locate any studies on the effects of the ILs used in this work on the inclusion of surfactant species into βCD cavities, their contributions to other cooperative processes (such as the micellization of surfactants and denaturation of

proteins) were examined previously. In general, the increase in the IL concentration (in the same or similar range) decreases the denaturation temperature (T_d) of proteins[49] and cmc of ionic surfactants.[50–52] This effect is more pronounced in the presence of ILs than in the presence of inorganic salts (such as LiCl, NaBr, or NaCl) due to the simultaneous electrostatic and hydrophobic interactions occurring in the aqueous solutions of ILs, which promote preferential solvation of the surfactant tail or protein surface.

To elucidate the effects of cosolutes on the $C_{12}Py^+$ inclusion into βCD cavities, the entropic and enthalpic contributions to ΔG_{ic}^o must be evaluated. Figure 4.8 shows the ΔH_{ic}^o and $T\Delta S_{ic}^o$ values obtained for the inclusion of $C_{12}Py^+$ ions into βCD cavities and plotted as functions of the cosolute concentration.

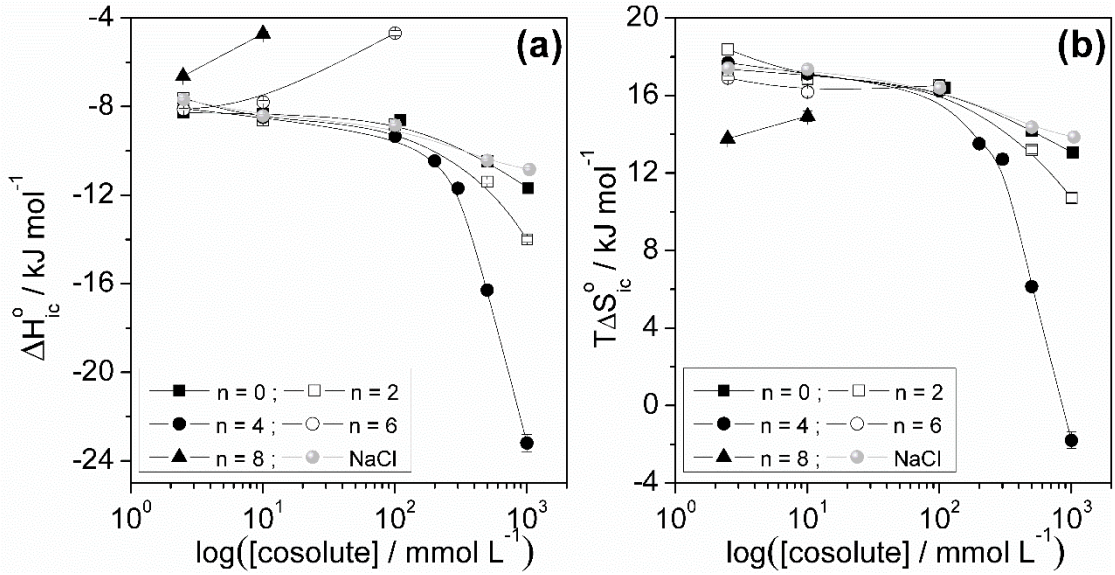

Figure 4.8. (a) ΔH_{ic}^o and **(b)** $T\Delta S_{ic}^o$ values obtained for the formation of βCD–$C_{12}Py^+$ inclusion complex in the *i*) water + $C_n$mimCl (n = 0, 2, 4, 6, or 8) and *ii*) water + NaCl mixtures at various cosolute concentrations and temperature of 298.15 K.

For all evaluated systems containing different cosolutes, the formation process of βCD–$C_{12}Py^+$ inclusion complex remained enthalpically and entropically favorable, except for the system with a C_4mimCl solvent concentration of 1000 mmol L^{-1}, which resulted in a small entropy decrease. Interestingly, when the concentration of the NaCl and $C_n$mimCl ($n \leq 4$) mixture increased to 100 mmol L^{-1}, the ΔH_{ic}^o and $T\Delta S_{ic}^o$ values remained almost constant regardless of the cosolute composition. However, at cosolute concentrations higher than 100 mmol L^{-1}, the βCD–$C_{12}Py^+$ inclusion process became abruptly more exothermic and less entropically favorable with ΔH_{ic}^o and $T\Delta S_{ic}^o$

magnitudes decreasing in the order of C_4mimCl > C_2mimCl > C_0mimCl > NaCl. On the other hand, the same process was less exothermic in the presence of larger amounts of C_6mimCl and C_8mimCl with ΔH_{ic}^{o} (C_8mimCl) > ΔH_{ic}^{o} (C_6mimCl) and became more entropically favorable only for the C_8mimCl IL (even at low contents).

As has been previously shown in equation 4.6, the ΔH_{ic}^{o} values represented the sums of three subprocesses occurring simultaneously ($\Delta H_{des}^{o,Surf}$, $\Delta H_{des}^{o,\beta CD}$, and $\Delta H_{int}^{o,\beta CD-Surf}$). According to our hypothesis, the increase in the C_4mimCl concentration promoted the decrease in ΔH_{ic}^{o} due to the preferential solvation of the surfactant tail by IL molecules, which made the $\Delta H_{des}^{o,Surf}$ values less positive. The larger decrease in ΔH_{ic}^{o} in the presence of C_4mimCl (as compared to those observed for the C_0mimCl and C_2mimCl ILs) can be attributed to the higher hydrophobicity of its longer alkyl chain, which promoted the solvation of the surfactant tail with a lower amount of energy required to break the IL–$C_{12}Py^+$ bonds.

Additionally, the kosmotropic effect of the $C_n$mimCl ILs also depends on the number of carbons in the lateral chain of the imidazolium cation. More hydrophobic ILs ($C_n$mimCl with higher n values) exhibit higher kosmotropicity in the three-dimensional structure of water molecules.[48] Thus, the presence of C_4mimCl promotes a more intense strengthening of the hydrogen bonds between water molecules in the bulk as compared with the effects produced by the C_2mimCl and C_0mimCl ILs. For this reason, the molecules released by the $C_{12}Py^+$ solvation shell due to the preferential solvation of the surfactant alkyl tail and from the interior of βCD cavity during the desolvation process make the $\Delta H_{des}^{o,Surf}$ and $\Delta H_{des}^{o,\beta CD}$ values less positive and more negative, respectively, when the C_4mimCl IL is used.

After taking into account the effect of the ILs discussed above and higher hydrophobicity of C_6mimCl and C_8mimCl, which promote the solvation of the surfactant tail, a more abrupt decrease in ΔH_{ic}^{o} values with increasing IL concentration can be expected. However, a completely opposite result was obtained because, in addition to the preferential solvation of the surfactant tail, the IL cations competed with water molecules for occupying βCD cavity due to the formation of βCD–C_6mim^+ (K_{ic} = 90 $mmol^{-1}$ L) and βCD–C_8mim^+ (K_{ic} = 624 $mmol^{-1}$ L) inclusion complexes in pure water.[53] Therefore, the $\Delta H_{des}^{o,\beta CD}$ value became less negative because the enthalpic balance associated with the

removal of C_6mim^+ or C_8mim^+ ions from βCD cavities is less favorable than the energetic balance related to the removal of water molecules from the same cavities.

A similar analysis procedure was performed to determine the effect of ILs on $T\Delta S_{ic}^{o}$ values. The trends depicted in Figure 4.8b supports our hypothesis about the preferential solvation and kosmotropic effect on the inclusion process. The increase in the degree of IL hydrophobicity reduced the number of water molecules solvating the $C_{12}Py^+$ tail, causing a decrease in $T\Delta S_{rel}^{o,H2O-Surf}$ as compared to the value obtained for pure water. In addition, increasing the amount of kosmotropic agents in the solution bulk strengthens the hydrogen bonds between water molecules, making the three-dimensional water structure more rigid, intensifying the decrease in $T\Delta S_{rel}^{o,H2O-\beta CD}$, and reducing the magnitude of $T\Delta S_{rel}^{o,H2O-Surf}$.

4.3.4. Effect of cosolute anion composition on βCD–$C_{12}Py^+$ association

The effect of the IL anion nature and composition on the formation of βCD–$C_{12}Py^+$ complexes was also evaluated in this work. For this purpose, the C_4mimBr and C_4mimI ILs with concentrations of 2.5, 10.0, and 100.0 mmol L^{-1} were used, and the determined thermodynamic parameters were compared with the results obtained in the presence of C_4mimCl. It was found that changing the composition of IL anions did not affect the thermodynamic parameters of the complexation reaction, indicating that the observed effects of the ILs on the βCD–$C_{12}Py^+$ complex formation were mainly attributed to the presence of imidazolium cations.

4.4. *Conclusions*

In this study, the ITC technique was successfully used to determine the driving forces of the βCD–$C_{12}Py^+$ inclusion complex formation in aqueous solutions of various cosolutes. The effects of the imidazolium ILs with different alkyl chain lengths on the thermodynamic parameters of the inclusion process were evaluated, revealing that $C_{12}Py^+$ and βCD species formed inclusion complexes with a 1:1 stoichiometry in pure water due to both the hydrophobic and dispersive interactions between the host and guest species.

The thermodynamic parameters of the formation of βCD–$C_{12}Py^+$ supramolecular complex were strongly dependent on the IL concentration and alkyl chain length of the

C_nmim^+ cation (n = 0, 2, 4, 6, or 8). In general, the presence of imidazolium cations in the system destabilized the βCD–$C_{12}Py^+$ inclusion complex: for the ILs with shorter lateral chains (n ≤ 4), this destabilization resulted from the decreased entropic contribution, whereas for the ILs with longer lateral chains (n ≥ 6), it was caused by the reduction of the enthalpic contribution. These phenomena originated from the balance existing between the following main processes: *i)* the preferential solvation of the $C_{12}Py^+$ tail (n = 0, 2, 4, 6 or 8), during in which C_nmim^+ cations competed with the water molecules occupying the βCD cavities (n = 6 or 8); and *ii)* the kosmotropic effect of ILs ions altering the three-dimensional water structure in the solution bulk. The obtained data can be used to achieve a better understanding of the influence of ILs on various aggregation processes.

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Capítulo 5: Considerações finais

Os resultados obtidos para os dois processos de agregação em água pura foram surpreendentes, uma vez que o avanço tecnológico em técnicas e métodos instrumentais permitiu desvendar as limitadas informações e discussões termodinâmicas sobre a formação de micelas de SDBS e formação de complexos de inclusão βCD–$C_{12}Py^+$ reportadas ao longo dos últimos 50 anos.

Adicional à informação termodinâmica da micelização do SDBS já reportada na literatura, conseguimos determinar, usando ITC e espectroscopia de fluorescência, que o processo de micelização ocorre em duas etapas, uma associada à transição de monômeros de SDBS para dímeros, transição fundamental para o entendimento das interações intermoleculares responsáveis pela formação de micelas de surfactantes aniônicos com anéis aromáticos no grupo cabeça, e uma segunda associada à transição de dímeros para micelas. Para a formação de complexos de inclusão βCD–$C_{12}Py^+$, foi determinado, em contraste ao reportado na literatura cientifica, que o processo é dirigido entálpica e entropicamente. Este fato modifica a perspectiva das interações não covalentes que são predominantes na formação do complexo supramolecular.

Em misturas aquosas de LIs, verificou-se que a termodinâmica de micelização e de formação de complexos de inclusão foi fortemente alterada, dependendo esta da estrutura molecular e concentração do cátion imidazólio em solução. A estabilidade termodinâmica das micelas de SDBS (ΔG^o_{mic}) aumentou na ordem $C_4mim^+ > C_0mim^+ > C_2mim^+$, enquanto que a entalpia (ΔH^o_{mic}) e a entropia ($T\Delta S^o_{mic}$) de micelização variaram na ordem $C_0mim^+ > C_2mim^+ > C_4mim^+$ e $C_4mim^+ > C_2mim^+ > C_0mim^+$, respectivamente. Nós sugerimos que os diferentes cátions C_nmim^+ interagem diferentemente com os agregados de SDBS, sendo que para LIs com $n \geq 2$, a interação ocorre com a superfície do agregado e as cadeias de SDBS dentro da micela, consequentemente aumentando o R_H do agregado, enquanto que para n = 0, embora o R_H das micelas tenha aumentado na presença do LI, a localização do cátion imidazólio foi similar à de um eletrólito simples.

No caso da formação de complexos de inclusão podemos indicar que em geral a presença dos cátions imidazólios desestabilizou o complexo βCD–$C_{12}Py^+$. Para LIs de cadeia curta ($n \leq 4$), a desestabilização resultou da diminuição na contribuição entrópica, enquanto que para LIs com cadeia longa ($n \geq 6$) ela foi causada pela redução da contribuição entálpica. Nós atribuímos este resultado ao balanço existente entre os

seguintes processos: solvatação preferencial da cauda do $C_{12}Py^+$ pelo C_nmim^+ (n = 0, 2, 4, 6 ou 8), competição entre o C_nmim^+ (n = 6 ou 8) e moléculas de água pela cavidade da βCD e o efeito cosmotrópico dos LIs que alteram a rede tridimensional das moléculas de água no *bulk* da solução.

Finalmente, podemos dizer que esta tese oferece contribuições científicas significativas no que diz respeito ao entendimento da influência dos LIs sobre processos de agregação molecular, principalmente sob um ponto de vista de ciência básica, guardando sempre a esperança de que algum dia possa ser vista como uma contribuição relevante para a sociedade.

Apêndice A

Supporting information of the article 1:

Aggregation of sodium dodecylbenzene sulfonate: weak molecular interactions modulated by the alkyl chain length of imidazolium cation species

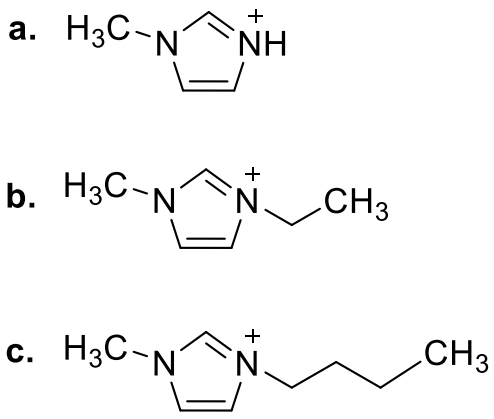

Figure A.S1. Imidazolium cations used in this work. **a.** 1-methylimidazolium, **b.** 1-ethyl-3-methylimidazolium, and **c.** 1-butyl-3-methylimidazolium.

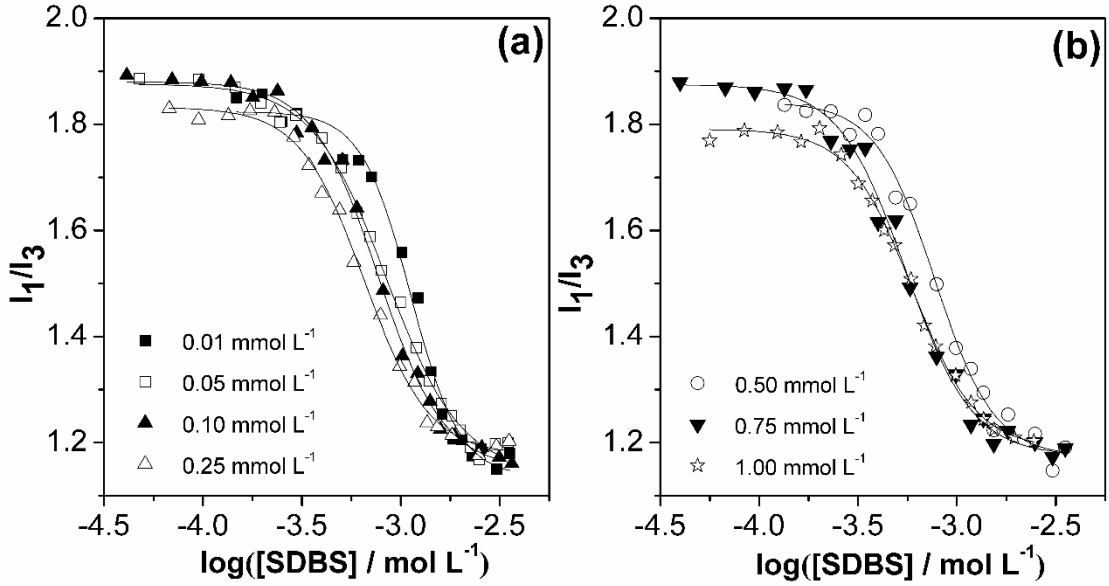

Figure A.S2. Quotient of pyrene vibrational band intensities (I_1/I_3) *versus* log[SDBS] obtained in C_0mimCl + water mixtures. **(a)** [C_0mimCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [C_0mimCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. (−) Fitted model to determine the cmc.

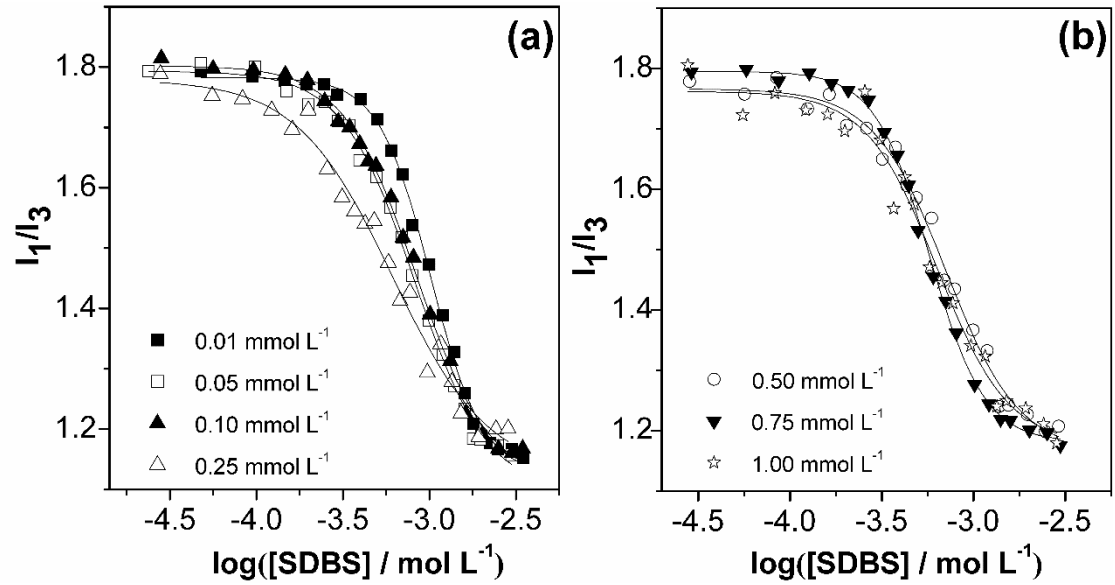

Figure A.S3. Quotient of pyrene vibrational band intensities (I_1/I_3) *versus* log[SDBS] obtained in C_2mimCl + water mixtures. **(a)** [C_2mimCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [C_2mimCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. (−) Fitted model to determine the cmc.

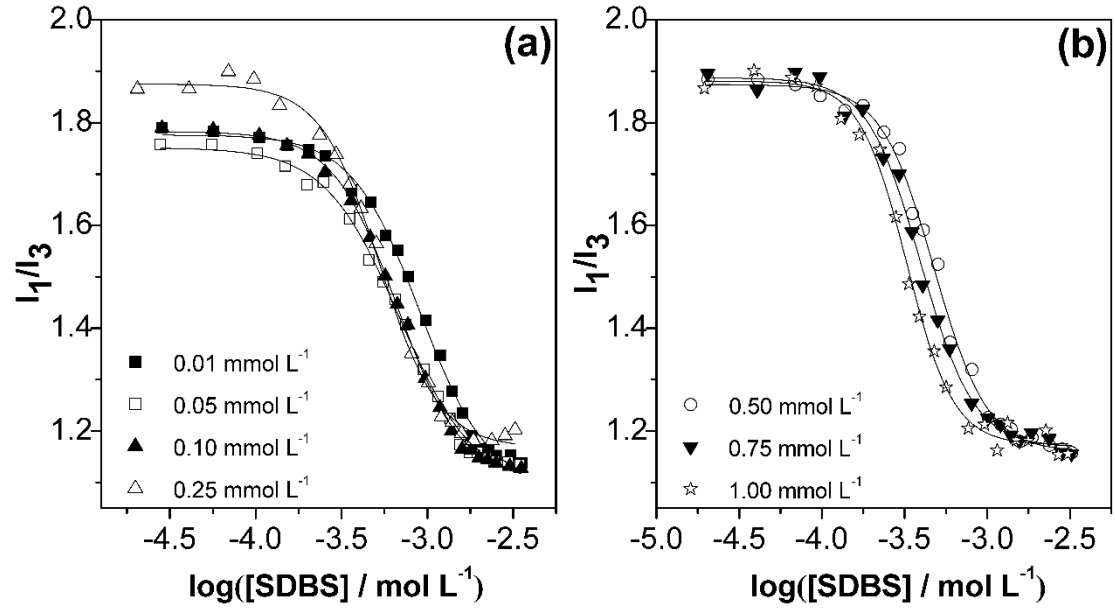

Figure A.S4. Quotient of pyrene vibrational band intensities (I_1/I_3) *versus* log[SDBS] obtained in C_4mimCl + water mixtures. **(a)** [C_4mimCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [C_4mimCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. (−) Fitted model to determine the cmc.

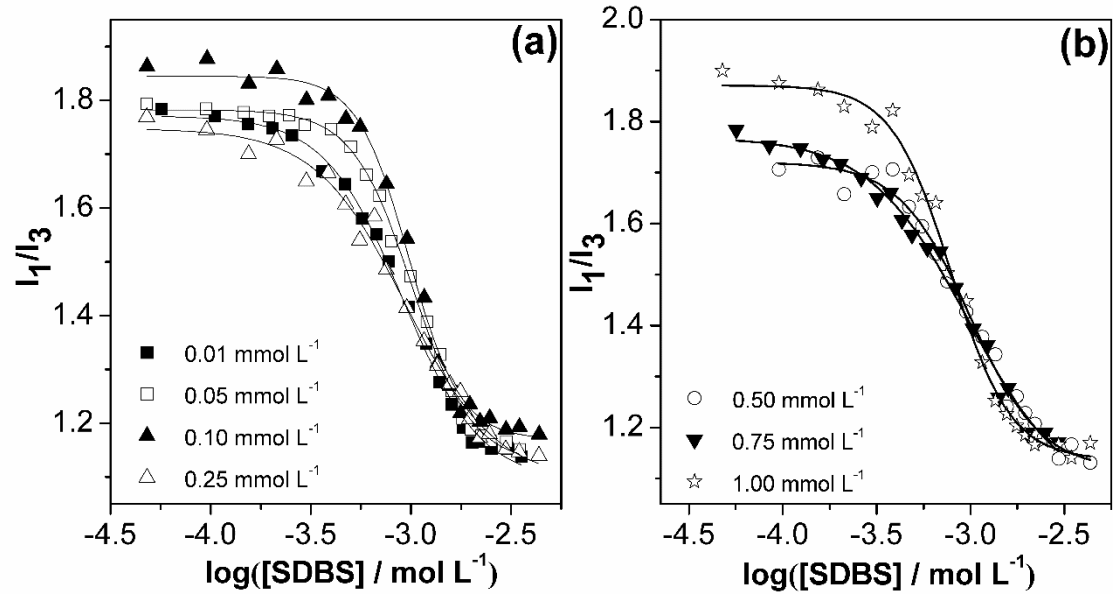

Figure A.S5. Quotient of pyrene vibrational band intensities (I_1/I_3) *versus* log[SDBS] obtained in NaCl + water mixtures. **(a)** [NaCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [NaCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. (−) Fitted model to determine the *cmc*.

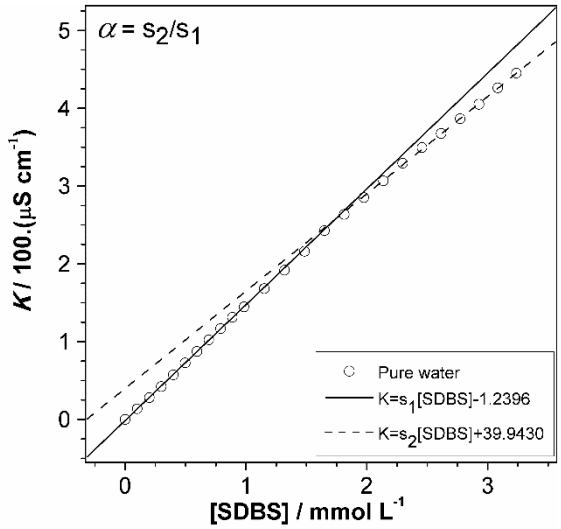

Figure A.S6. Conductivity *versus* [SDBS] obtained in pure water, at 298.2 K. For all measurements, the specific conductivity of pure water was discounted. (–,--) linear fitting to obtain the slopes s_1 (148.82) and s_2 (125.07).

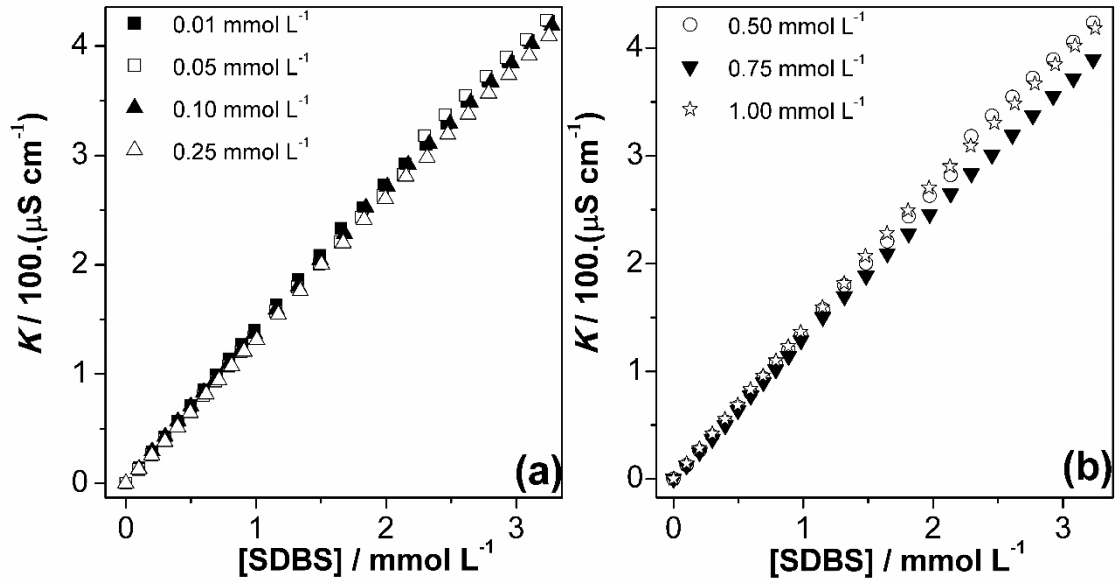

Figure A.S7. Conductivity *versus* [SDBS] obtained in C_0mimCl + water mixtures. **(a)** [C_0mimCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [C_0mimCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. For all measurements, the specific conductivity of the solvent was discounted.

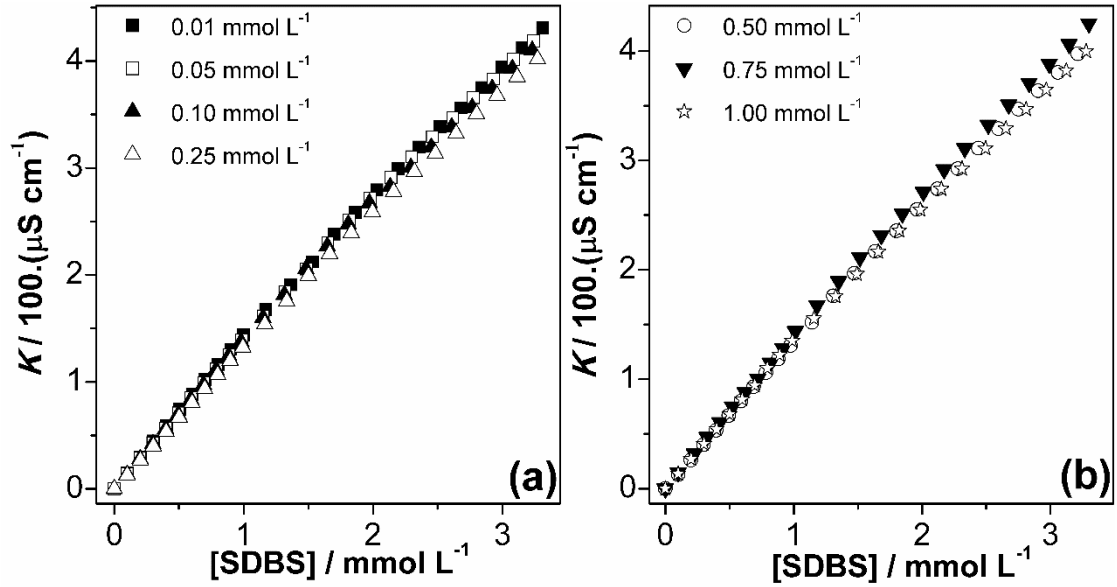

Figure A.S8. Conductivity *versus* [SDBS] obtained in C_2mimCl + water mixtures. **(a)** [C_2mimCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [C_2mimCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. For all measurements, the specific conductivity of the solvent was discounted.

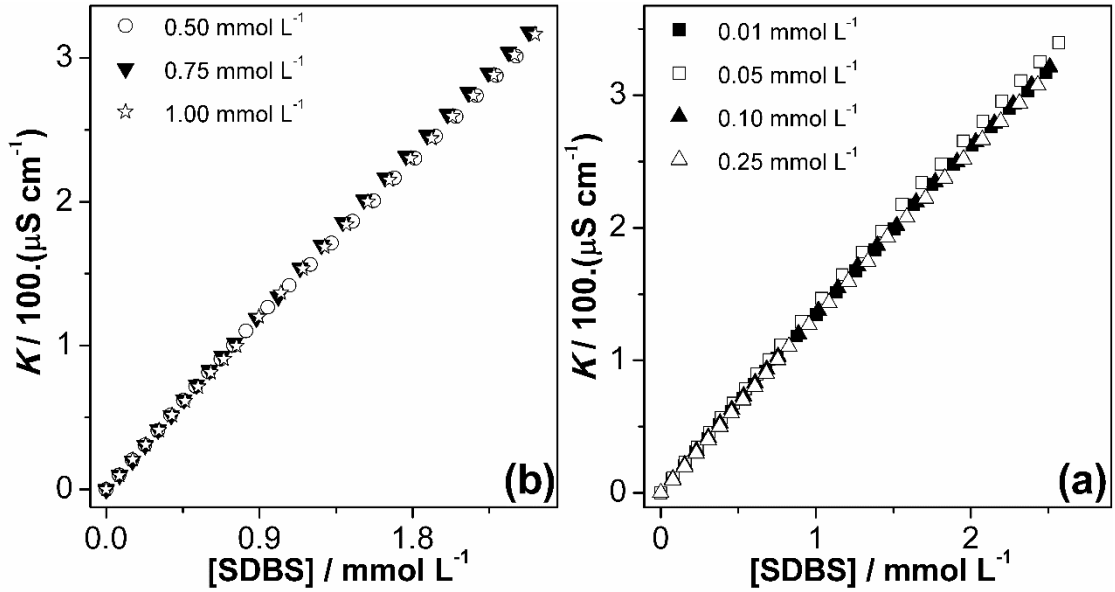

Figure A.S9. Conductivity *versus* [SDBS] obtained in C_4mimCl + water mixtures. **(a)** [C_4mimCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [C_4mimCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. For all measurements, the specific conductivity of the solvent was discounted.

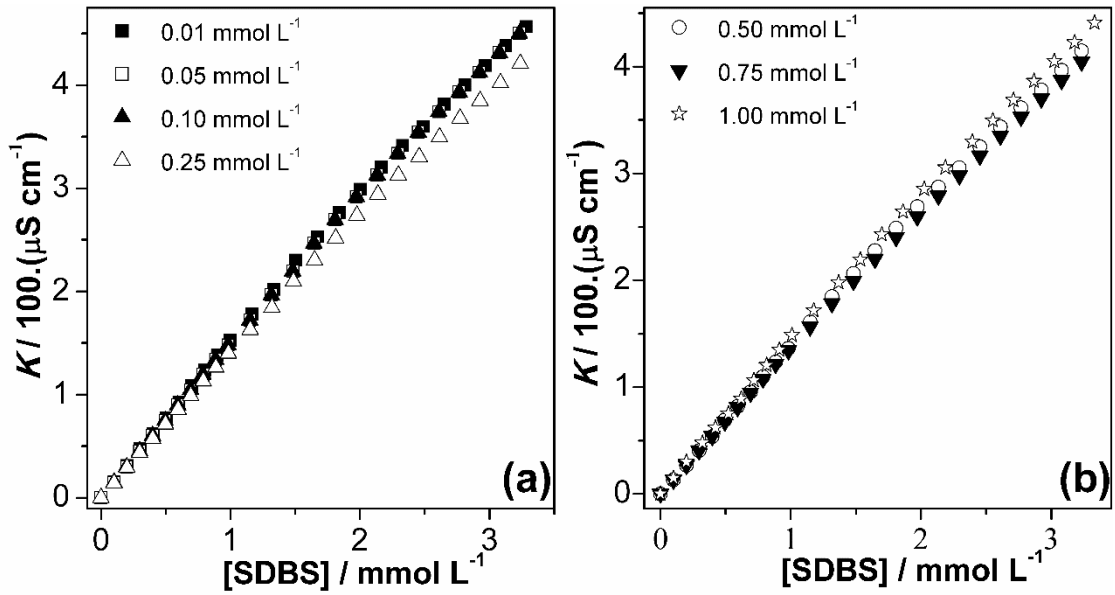

Figure A.S10. Conductivity *versus* [SDBS] obtained in NaCl + water mixtures. **(a)** [NaCl] ≤ 0.25 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} < [NaCl] ≤ 1.00 mmol L^{-1}, at 298.2 K. For all measurements, the specific conductivity of the solvent was discounted.

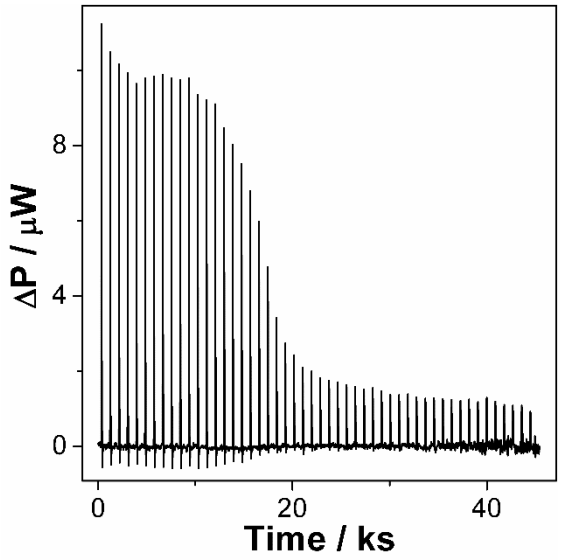

Figure A.S11. Thermal power as a function of time for sequential injections (5 μL) of 25 mmol L^{-1} SDBS into 2.7 mL of pure water at 298.15 K.

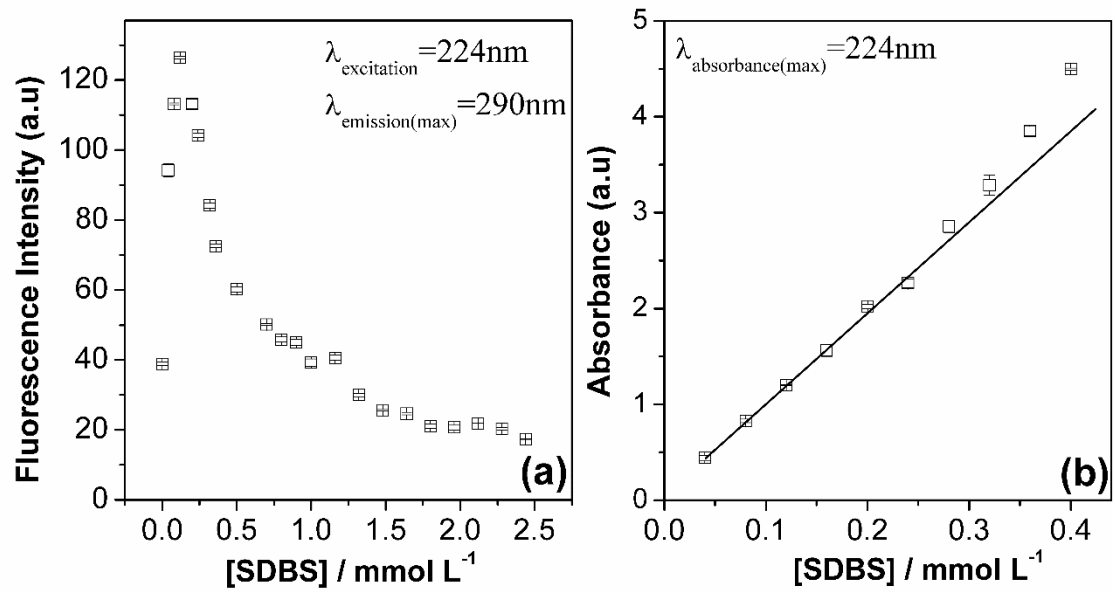

Figure A.S12. (**a**) Fluorescence intensities and (**b**) absorbance of SDBS solutions as a function of surfactant concentration at 298.2 K. The black line in (**b**) represents the linear fit associated with the absorbance of SDBS monomers ([SDBS] ≤ 0.25 mmol L^{-1}). The excitation, emission maximum, and absorbance maximum wavelengths were 224, 290 and 224 nm, respectively.

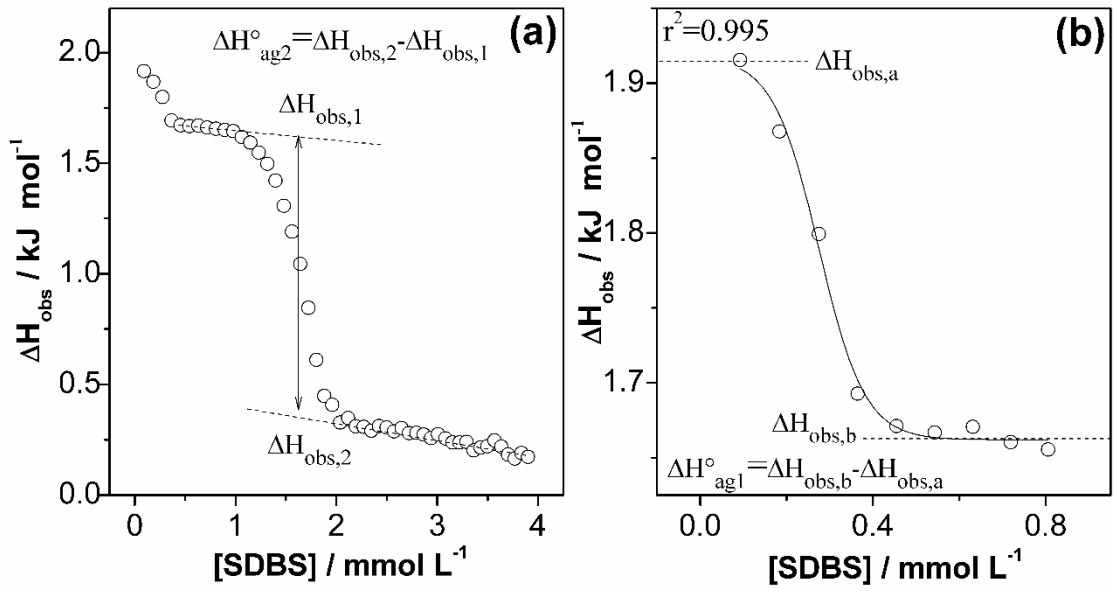

Figure A.S13. ΔH_{obs} as a function of [SDBS] into pure water at 298.15 K. **(a)** ΔH^{o}_{ag2} and **(b)** ΔH^{o}_{ag1} determination. (−) Boltzmann fitting.

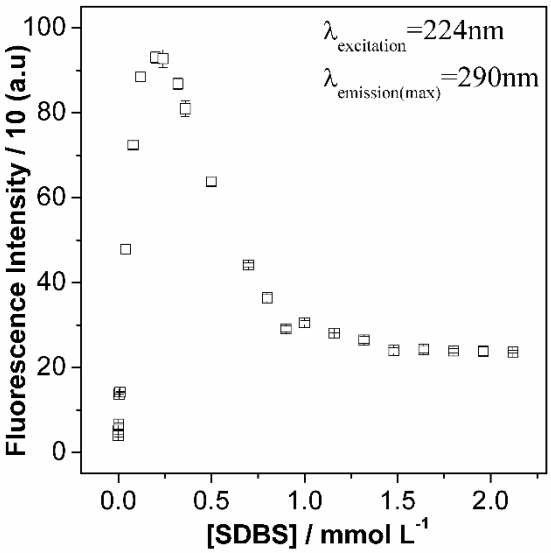

Figure A.S14. Fluorescence intensities of SDBS solutions *versus* surfactant concentration at 298.2 K. [C_4mimCl] = 1 mmol L^{-1}. The excitation, emission maximum wavelengths were 224 and 290 nm, respectively.

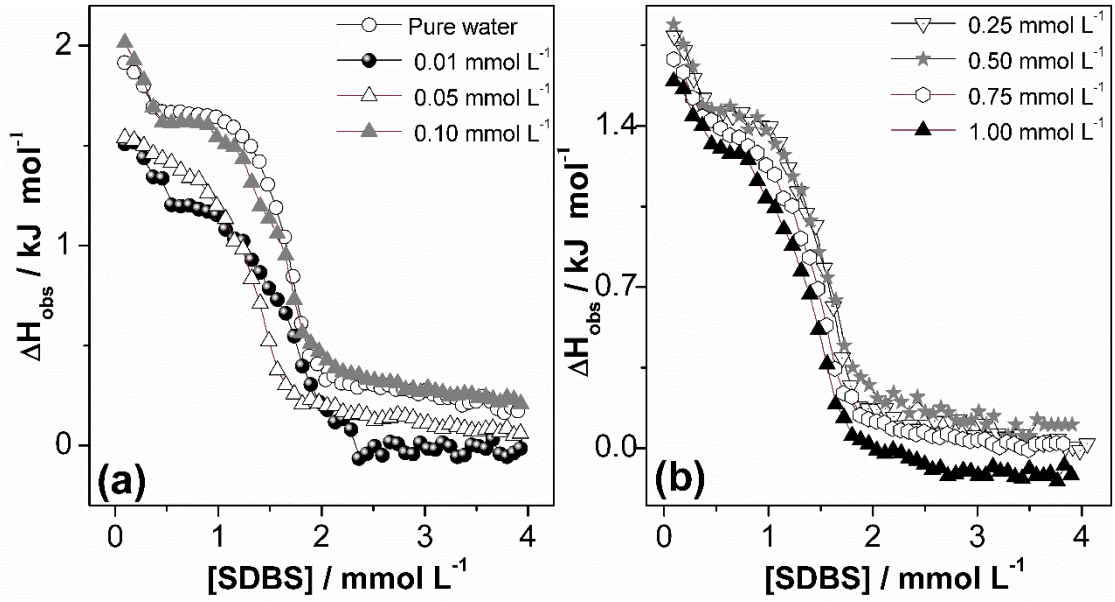

Figure A.S15. ΔH_{obs} *versus* [SDBS] curves obtained by dilution of SDBS solutions in NaCl + water mixtures. **(a)** [NaCl] ≤ 0.10 mmol L^{-1} and **(b)** 0.25 mmol L^{-1} ≤ [NaCl] ≤ 1.00 mmol L^{-1}, at 298.15 K.

Table A.S1. *cmc* (obtained by ITC measurements), standard aggregation parameters (ΔH^o_{ag1}, ΔH^o_{ag2}, ΔH^o_{mic}, $T\Delta S^o_{mic}$ and ΔG^o_{mic}), and α for SDBS micellization process in different water + IL mixtures, at 298.15 K.

[IL]	cmc	ΔH^o_{ag1}	ΔH^o_{ag2}	ΔH^o_{mic}	$T\Delta S^o_{mic}$	ΔG^o_{mic}	α
mmol L^{-1}	mmol L^{-1}	kJ mol^{-1}			kJ mol^{-1}	kJ mol^{-1}	
C_0mimCl							
0.00	1.65 ± 0.01	−0.26 ± 0.02	−1.26 ± 0.03	−1.52 ± 0.04	18.3 ± 0.2	−19.8	0.84
0.01	1.66 ± 0.01	−0.53 ± 0.01	−1.37 ± 0.01	−1.90 ± 0.02	18.1 ± 0.1	−20.0	0.83
0.05	1.44 ± 0.01	−0.39 ± 0.02	−2.04 ± 0.03	−2.43 ± 0.03	17.4 ± 0.3	−19.8	0.87
0.10	1.39 ± 0.01	−1.09 ± 0.04	−1.9 ± 0.1	−3.0 ± 0.1	16.9 ± 0.1	−19.9	0.88
0.25	1.11 ± 0.02	−	−	−3.45 ± 0.08	17.0 ± 0.2	−20.4	0.88
0.50	1.05 ± 0.01	−	−	−3.50 ± 0.05	17.1 ± 0.2	−20.6	0.88
0.75	0.92 ± 0.01	−	−	−3.66 ± 0.06	17.2 ± 0.2	−20.9	0.88
1.00	0.93 ± 0.01	−	−	−3.75 ± 0.07	17.1 ± 0.2	−20.9	0.88
C_2mimCl							
0.01	1.58 ± 0.01	−0.53 ± 0.04	−1.16 ± 0.03	−1.69 ± 0.05	18.7 ± 0.2	−20.4	0.81
0.05	1.43 ± 0.01	−0.60 ± 0.02	−1.37 ± 0.08	−1.97 ± 0.08	18.7 ± 0.2	−20.7	0.81
0.10	1.39 ± 0.01	−0.52 ± 0.05	−1.5 ± 0.1	−2.0 ± 0.1	18.7 ± 0.2	−20.7	0.81
0.25	1.32 ± 0.01	−1.1 ± 0.1	−1.52 ± 0.02	−2.7 ± 0.1	18.4 ± 0.4	−21.1	0.82
0.50	1.20 ± 0.01	−1.6 ± 0.1	−1.46 ± 0.03	−3.1 ± 0.1	18.0 ± 0.2	−21.1	0.83
0.75	1.10 ± 0.01	−1.64 ± 0.08	−1.56 ± 0.07	−3.2 ± 0.1	17.5 ± 0.4	−20.7	0.85
1.00	1.12 ± 0.01	−1.7 ± 0.1	−1.4 ± 0.1	−3.1 ± 0.2	17.7 ± 0.3	−20.8	0.86
C_4mimCl							
0.01	1.58 ± 0.01	−0.51 ± 0.05	−1.32 ± 0.02	−1.65 ± 0.05	18.2 ± 0.1	−19.9	0.86
0.05	0.98 ± 0.03	−0.53 ± 0.06	−1.19 ± 0.07	−1.62 ± 0.04	18.8 ± 0.2	−20.4	0.87
0.10	0.84 ± 0.02	−	−	−1.76 ± 0.06	18.7 ± 0.1	−20.5	0.87
0.25	0.54 ± 0.02	−	−	−2.66 ± 0.02	18.6 ± 0.1	−21.2	0.88
0.50	0.48 ± 0.02	−	−	−2.81 ± 0.02	18.4 ± 0.2	−21.2	0.88
0.75	0.43 ± 0.02	−	−	−2.97 ± 0.01	18.7 ± 0.1	−21.7	0.89
1.00	0.38 ± 0.02	−	−	−2.69 ± 0.01	19.3 ± 0.3	−22.0	0.90

Table A.S2. *cmc* (obtained by ITC measurements), standard aggregation parameters (ΔH^o_{ag1}, ΔH^o_{ag2}, ΔH^o_{mic}, $T\Delta S^o_{mic}$ and ΔG^o_{mic}), and α for SDBS micellization process in different water + NaCl mixtures, at 298.15 K.

[NaCl]	cmc	ΔH^o_{ag1}	ΔH^o_{ag2}	ΔH^o_{mic}	$T\Delta S^o_{mic}$	ΔG^o_{mic}	α
mmol L^{-1}	mmol L^{-1}	kJ mol^{-1}			kJ mol^{-1}	kJ mol^{-1}	
0.00	1.65 ± 0.01	−0.26 ± 0.02	−1.26 ± 0.03	−1.52 ± 0.04	18.3 ± 0.2	−19.8	0.84
0.01	1.64 ± 0.01	−0.36 ± 0.03	−1.24 ± 0.08	−1.60 ± 0.09	18.7 ± 0.1	−20.3	0.82
0.05	1.64 ± 0.01	−0.37 ± 0.04	−1.27 ± 0.03	−1.64 ± 0.05	18.7 ± 0.1	−20.3	0.82
0.10	1.63 ± 0.01	−0.41 ± 0.03	−1.37 ± 0.02	−1.78 ± 0.03	18.8 ± 0.2	−20.6	0.82
0.25	1.61 ± 0.01	−0.37 ± 0.03	−1.40 ± 0.02	−1.77 ± 0.03	18.7 ± 0.1	−20.5	0.81
0.50	1.61 ± 0.01	−0.37 ± 0.04	−1.37 ± 0.02	−1.74 ± 0.04	18.7 ± 0.2	−20.4	0.81
0.75	1.56 ± 0.01	−0.29 ± 0.01	−1.34 ± 0.02	−1.63 ± 0.02	18.8 ± 0.1	−20.4	0.82
1.00	1.44 ± 0.01	−0.23 ± 0.02	−1.24 ± 0.02	−1.47 ± 0.03	19.2 ± 0.2	−20.7	0.81

Apêndice B

Supporting information of the article 2:

Solvophobic effect of 1-alkyl-3-methylimidazolium halides on the thermodynamic of complexation between β-cyclodextrin and dodecylpiridinium cation

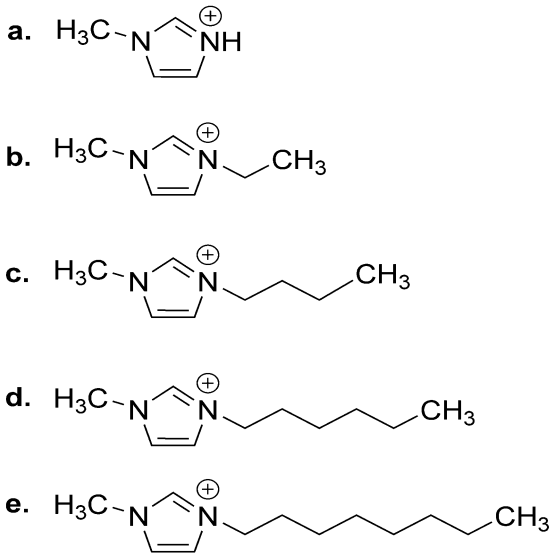

Figure B.S1. Imidazole cations used in this work. **a.** 1-methylimidazolium, **b.** 1-ethyl-3-methylimidazolium, **c.** 1-butyl-3-methylimidazolium, **d.** 1-hexyl-3-methylimidazolium and **e.** 1-octyl-3-methylimidazolium.

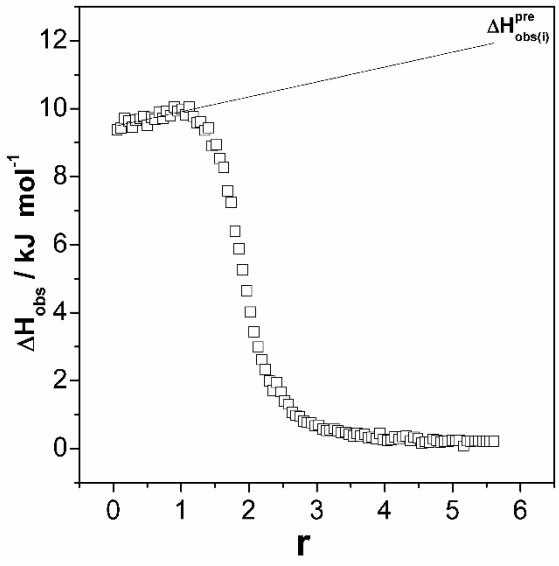

Figure B.S2. ΔH_{obs} as a function of r obtained by dilution of C_{12}PyCl in presence of 1000 mmol L^{-1} C_4mimCl, at 298.15 K. The black line represents a linear extrapolation of $\Delta H_{obs\,(i)}^{pre}$ values

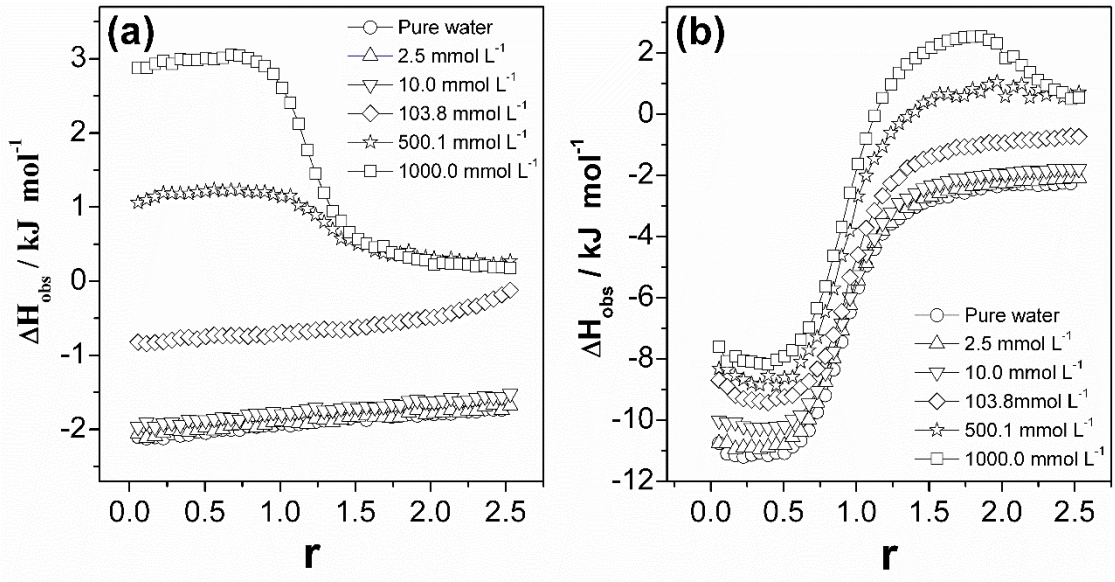

Figure B.S3. ΔH_{obs} as a function of r obtained by titration of C_{12}PyCl in different solvents: **(a)** water + C_0mimCl and **(b)** 2.1 mmol L^{-1} βCD + water + C_0mimCl at 298.15 K.

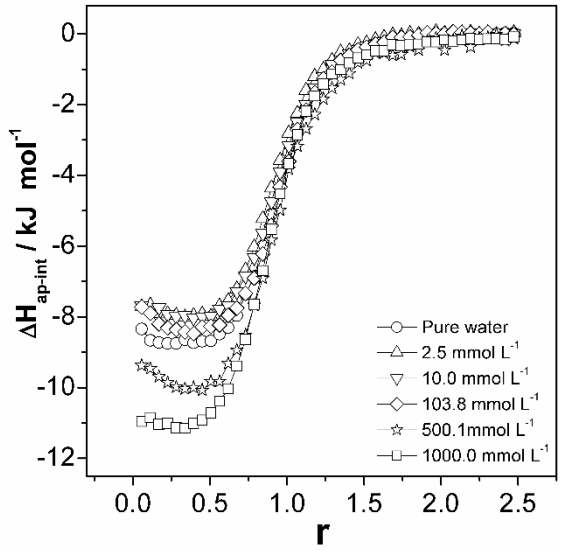

Figure B.S4. ΔH_{ap-int} *versus* r for βCD−$C_{12}Py^+$ complexation processes in C_0mimCl + water mixtures.

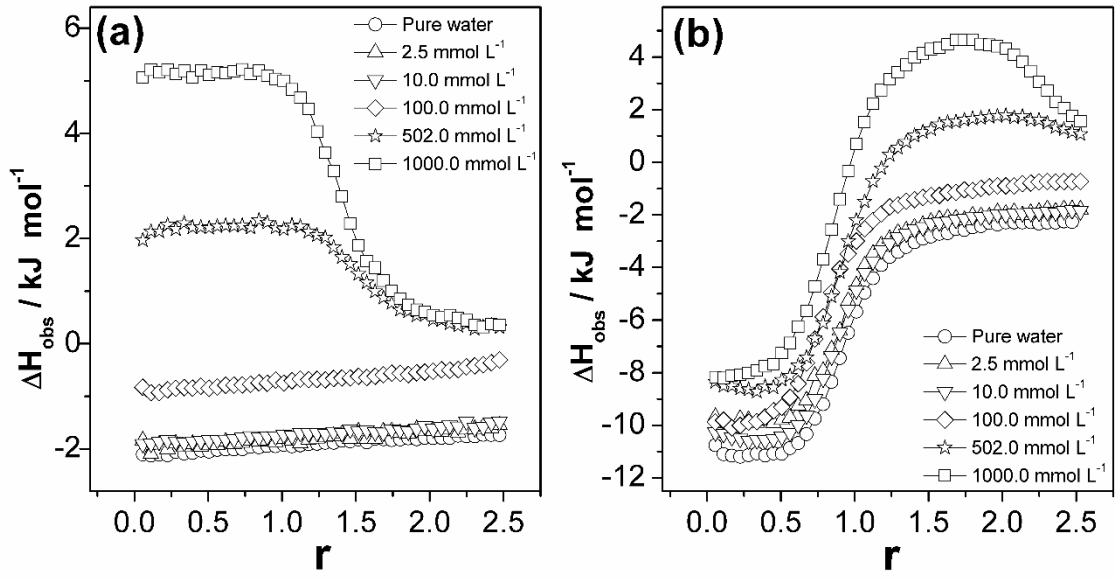

Figure B.S5. ΔH_{obs} as a function of r obtained by titration of C_{12}PyCl in different solvents: **(a)** water + C_2mimCl and **(b)** 2.1 mmol L^{-1} βCD + water + C_2mimCl at 298.15 K.

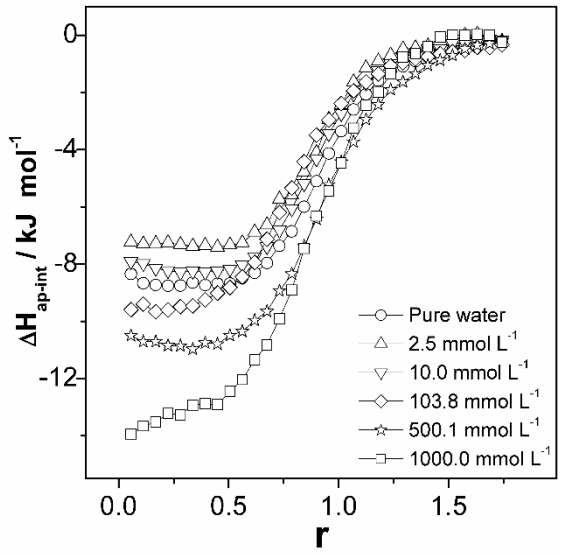

Figure B.S6. ΔH_{ap-int} *versus* r for βCD−$C_{12}Py^+$ complexation processes in C_2mimCl + water mixtures.

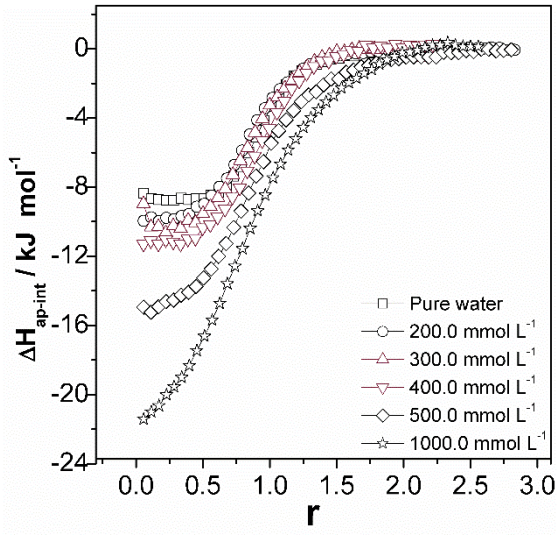

Figure B.S7. ΔH_{ap-int} *versus* r for βCD–$C_{12}Py^+$ complexation processes in C_4mimCl + water mixtures.

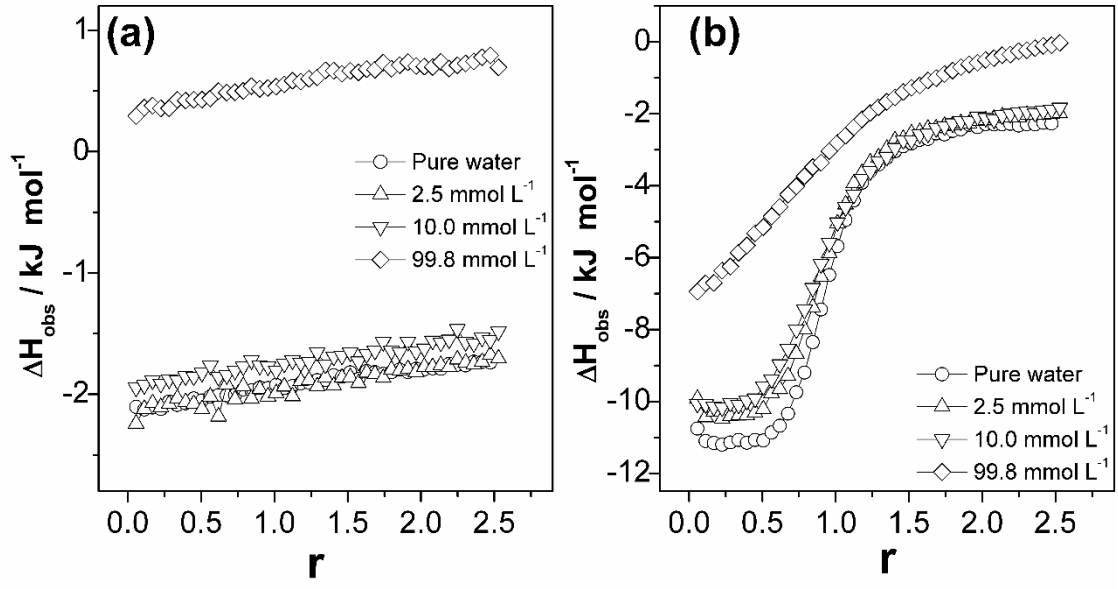

Figure B.S8. ΔH_{obs} as a function of r obtained by titration of C_{12}PyCl in different solvents: **(a)** water + C_6mimCl and **(b)** 2.1 mmol L^{-1} βCD + water + C_6mimCl at 298.15 K.

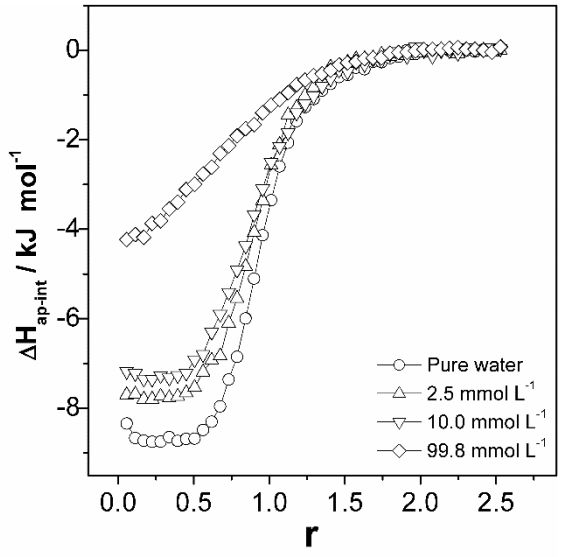

Figure B.S9. ΔH_{ap-int} *versus* r for βCD−$C_{12}Py^+$ complexation processes in C_6mimCl + water mixtures.

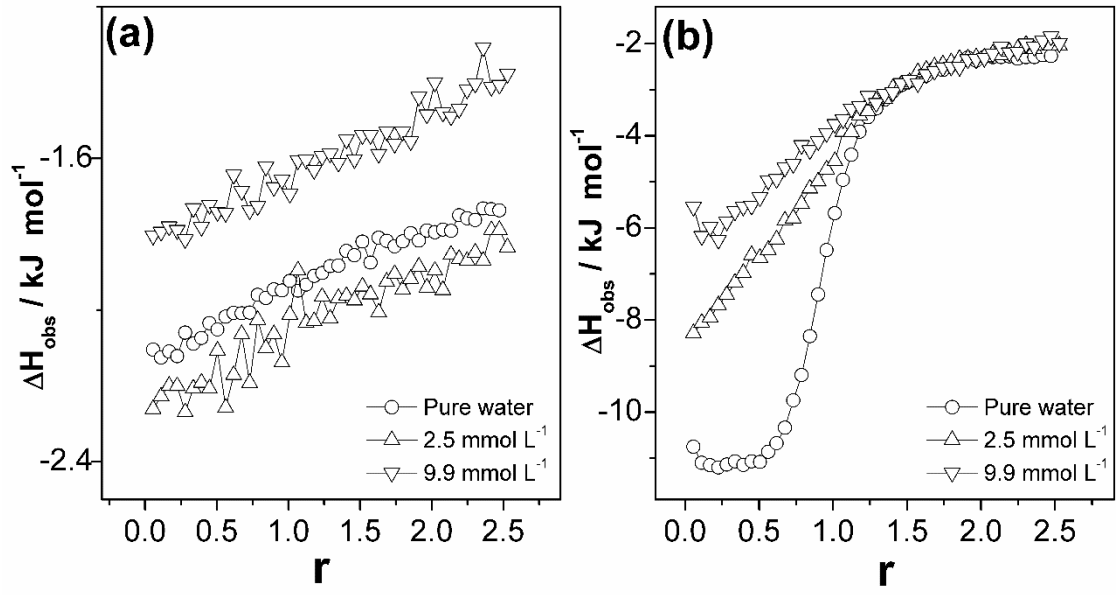

Figure B.S10. ΔH_{obs} as a function of r obtained by titration of C_{12}PyCl in different solvents: **(a)** water + C_8mimCl and **(b)** 2.1 mmol L^{-1} βCD + water + C_8mimCl at 298.15 K.

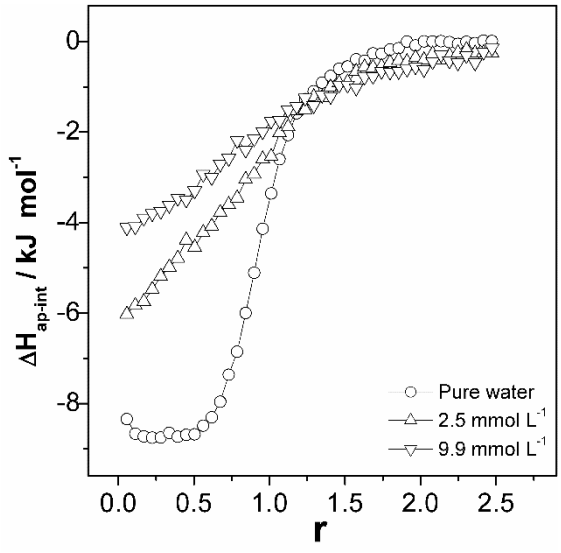

Figure B.S11. ΔH_{ap-int} *versus* r for βCD−$C_{12}Py^+$ complexation processes in C_8mimCl + water mixtures.

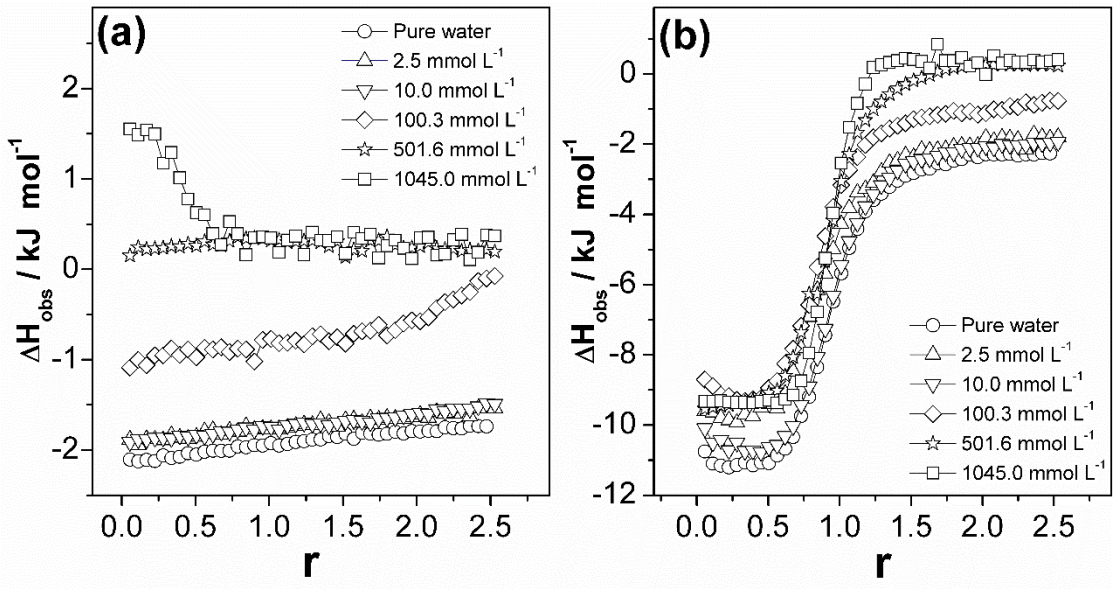

Figure B.S12. ΔH_{obs} as a function of r obtained by titration of C_{12}PyCl in different solvents: **(a)** water + NaCl and **(b)** 2.1 mmol L^{-1} βCD + water + NaCl at 298.15 K.

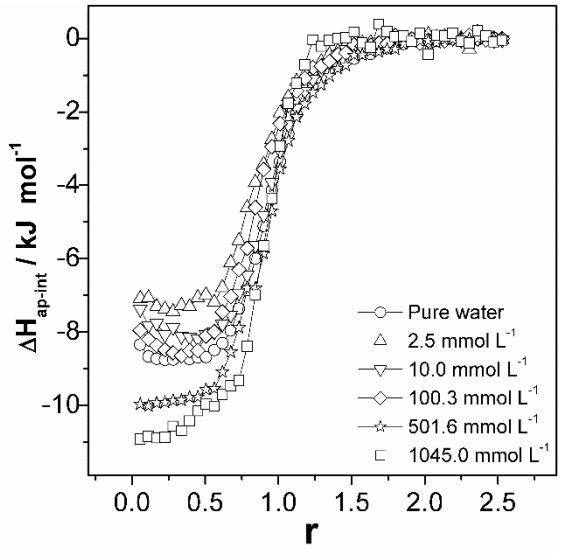

Figure B.S13. ΔH_{ap-int} *versus* r for βCD−$C_{12}Py^+$ complexation processes in NaCl + water mixtures.

Table B.S1. Inclusion thermodynamic parameters for βCD−$C_{12}Py^+$ complexation in different water + ILs mixtures at 298.15 K.

$C_n$mimCl	K_{ic}	ΔG^o_{ic}	ΔH^o_{ic}	$T\Delta S^o_{ic}$	n
mmol L^{-1}	10^4 x L mol^{-1}	kJ mol^{-1}			
C_0mimCl					
0.0	2.6 ± 0.2	−25.2 ± 0.2	−9.2 ± 0.1	16.1 ± 0.2	0.94 ± 0.01
2.5	3.1 ± 0.3	−25.6 ± 0.2	−8.3 ± 0.1	17.4 ± 0.2	0.93 ± 0.01
10.0	2.8 ± 0.2	−25.4 ± 0.2	−8.3 ± 0.1	17.1 ± 0.2	0.96 ± 0.01
110.0	2.4 ± 0.3	−25.0 ± 0.3	−8.6 ± 0.1	16.4 ± 0.3	0.96 ± 0.01
500.0	2.1 ± 0.2	−24.7 ± 0.3	−10.5 ± 0.1	14.2 ± 0.3	0.94 ± 0.01
1021.0	2.2 ± 0.1	−24.7 ± 0.1	−11.7 ± 0.1	13.1 ± 0.2	0.91 ± 0.01
C_2mimCl					
2.5	3.6 ± 0.3	−26.0 ± 0.2	−7.6 ± 0.1	18.4 ± 0.2	0.89 ± 0.01
10.0	2.9 ± 0.3	−25.5 ± 0.2	−8.6 ± 0.1	16.6 ± 0.2	0.91 ± 0.01
100.0	2.8 ± 0.3	−25.4 ± 0.2	−8.8 ± 0.1	16.6 ± 0.2	1.01 ± 0.01
502.0	2.0 ± 0.2	−24.6 ± 0.2	−11.4 ± 0.1	13.2 ± 0.2	0.96 ± 0.01
1002.0	2.1 ± 0.2	−24.7 ± 0.2	−14.0 ± 0.1	10.7 ± 0.3	0.88 ± 0.01
C_6mimCl					
2.5	2.4 ± 0.1	−25.0 ± 0.1	−8.21 ± 0.04	16.9 ± 0.1	0.90 ± 0.01
10.0	1.6 ± 0.1	−24.0 ± 0.2	−7.8 ± 0.1	16.2 ± 0.2	0.90 ± 0.01
99.8	0.52 ± 0.1	−21.0 ± 0.2	−4.7 ± 0.1	16.5 ± 0.3	0.77 ± 0.01
C_8mimCl					
2.5	0.37 ± 0.02	−20.4 ± 0.2	−6.6 ± 0.1	13.8 ± 0.2	0.91 ± 0.01
9.9	0.28 ± 0.03	−19.6 ± 0.3	−4.7 ± 0.1	14.9 ± 0.7	1.01 ± 0.02

Table B.S2. Inclusion thermodynamic parameters for βCD−$C_{12}Py^+$ complexation in different water + NaCl mixtures at 298.15 K.

NaCl	K_{ic}	ΔG^o_{ic}	ΔH^o_{ic}	$T\Delta S^o_{ic}$	n
mmol L^{-1}	10^4 x L mol^{-1}		kJ mol^{-1}		
0.0	2.6 ± 0.2	−25.2 ± 0.2	−9.2 ± 0.1	16.1 ± 0.2	0.94 ± 0.01
2.5	2.5 ± 0.2	−25.1 ± 0.2	−7.7 ± 0.1	17.4 ± 0.3	0.88 ± 0.01
10.0	3.3 ± 0.3	−25.8 ± 0.2	−8.4 ± 0.1	17.4 ± 0.2	0.95 ± 0.01
100.3	2.6 ± 0.2	−25.2 ± 0.2	−8.8 ± 0.1	16.4 ± 0.2	0.88 ± 0.01
501.6	2.2 ± 0.1	−24.8 ± 0.1	−10.4 ± 0.1	14.4 ± 0.1	0.93 ± 0.01
1045.0	2.1 ± 0.6	−24.7 ± 0.7	−10.8 ± 0.1	13.9 ± 0.7	0.90 ± 0.01