Lithium adsorption by acid and sodium amberlite

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Abstract

In this paper we used a previously reported model for examining the adsorption of nonelectrolytes in solution by solid adsorbents to study the adsorption of lithium(I) cations by acid and sodium amberlites, which is an ion-exchange process. Based on the results, both are equilibrium processes and obey a kinetic law with a unity partial order in the Li\(^+\) concentration. The kinetic results were used to calculate the specific rate constants and thermodynamic activation functions involved. Also, equilibrium isotherms were used to determine the corresponding ion-exchange capacities, the individual equilibrium constants, and the thermodynamic functions for the overall process.

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1. Introduction

The adsorption of dissolved nonelectrolytes by solid adsorbents and the exchange of dissolved ions with others of the same sign retained at sites of the opposite sign in a solid exchanger are both phase-equilibrium processes. Consequently, at least macroscopically, the laws governing them and their kinetics must be similar, as are the experimental methods used to examine them. However, the models used to fit the results obtained for these processes differ considerably in both the approach and the equations employed.

To the authors’ minds, the adsorption of nonelectrolytes and the exchange of ions in solution can be assumed to be similar and, to a certain extent, equivalent at least at the formal and macroscopic levels. This assertion is supported by qualitative inferences from the cycles of Figs. 1 and 2, where the adsorption was process assumed to involve the exchange of solvent molecules, \(D\), retained at the active sites of the solid, with molecules of the adsorbate.

The above-described processes can be represented, in simplified form, by the equations

\[ S + A \rightarrow S - A, \]

or, if the ions \(A\) and \(B\) possess the same charge,

\[ S - BzA + zBA \rightarrow S - AzB + zAB, \]

Formally, Eqs. (1) and (3) will be equivalent if the active sites of the solid adsorbent are assumed to play the role of \(B\). Based on the previous reasoning, the models used to examine the adsorption of nonelectrolytes and the exchange of ions in solution can be employed indifferently with both types of process. The former greatly facilitate the determination of the adsorption and exchange capacities and allow one to detect significant changes in the adsorbate or the retained ion such as the formation of oligomers or ions with different charges frequently via hydroxyaqua complexes. On the other hand, the latter are especially useful with a view to calculating the thermodynamic equilibrium constant.

The foregoing was considered in developing a simple model for examining various processes occurring at the solid–liquid interface. The use of this model to study the adsorption of a nonelectrolyte in solution provided the results reported in two previous papers by Valenzuela-Calahorra et al. [1,2]. The present paper reports the results obtained by using the model to examine the exchange of dissolved lithium(I) ions for hydrogen ions or sodium(I) ions retained on a synthetic cation-exchange column.

S - BzA + zBA \(\rightarrow\) S - AzB + zAB,

\[ S - B + A \rightarrow S - A + B. \]
In addition to checking the previous hypothesis, this work was of practical interest in that lithium(I) resins can be used as the starting materials to prepare controlled-release oral pharmaceutical formulations of this cation, which are currently being used in various medical fields, particularly in psychiatry, where they have proved effective in the treatment of bipolar disorders [3].

2. Materials and methods

We used two different cation-exchange resins, namely Amberlite IR 120 plus in its sodium form (A–Na, Sigma–Aldrich Reference No. 22.435-9) and hydrogen form (A–H, Sigma–Aldrich Reference No. 21.653-4). Both products were granular and supplied in the form of roughly spherical particles ranging from 0.297 to 1.168 mm in diameter. The exchange capacity of both A–Na and A–H is ca. \(4.4 \times 10^{-3}\) eq/g dry resin (supplied by Sigma–Aldrich).

The aqueous solutions of lithium ion used in the exchange process were prepared from lithium chloride (Merck, pro-analysis reagent) and deionized, bidistilled water.

The kinetic law governing each resin(s)/Li\(^+\)(d) exchange process was determined in two experiments. In the first, the partial order of the process in the Li\(^+\) concentration in the dissolved phase was calculated; in the second, the specific rate of adsorption of lithium ions by the amberlite was determined. The procedures used in both series were similar and involved placing 0.1 g of acid or sodium amberlite in a 250-ml glass flask furnished with a threaded cap and adding 100 ml of a lithium chloride solution of known concentration; the flask was then stirred at 150 rpm at a temperature constant to within \(\pm 0.1\) °C in a Gallenkamp apparatus for a preset time, after which the Li\(^+\) concentration in solution was determined spectrophotometrically using a Perkin–Elmer 1100B instrument at 670.8 nm. The first series of experiments was conducted at 20 °C, using five solutions of concentrations between \(7 \times 10^{-4}\) and \(2 \times 10^{-3}\) M; the second, using a single LiCl solution ([Li\(^+\)] = \(1.8 \times 10^{-3}\) M) and four different temperatures (viz. 10, 20, 30, and 40 °C).

It should be noted that the amberlites were stored in a moisture-saturated atmosphere at 25 °C; under these conditions, their moisture content was in the region of 50%, so the exchange capacity for these samples can be assumed to be about \(2.2 \times 10^{-3}\) eq/g wet resin.

The data required to run the equilibrium isotherms were obtained similarly, using 0.25 g of A–H or A–Na and 100 ml of an aqueous solution of Li\(^+\) ranging from \(7.2 \times 10^{-4}\) to \(0.14\) M at a temperature of 10, 20, 30, or 40 °C; the phases were allowed to remain in contact for 7 days (i.e., much longer than the time strictly needed for equilibrium to be reached).

3. Results and discussion

3.1. Kinetics

The results obtained in the above-described experiments were used to run the \(C–t\) plots shown in Figs. 3 and 4.

As noted earlier, the exchange of Na\(^+\) or H\(^+\) ions in the solid with Li\(^+\) in solution can be schematized by Eq. (3) and its kinetic law represented by

\[
\frac{dC}{dt} = k_a C^{n_1}(1 - \theta)^{n_2} - k_d \theta^{n_3},
\]

where \(n_1\), \(n_2\), and \(n_3\) are partial reaction orders in \(C\), \((1 - \theta)\), and \(\theta\), respectively, \(k_a\) = specific rate of adsorption of Li\(^+\) ions by the solid, \(k_d\) = specific rate of desorption of Li\(^+\) ions, \(\theta\) = fraction of anionic sites in the exchanger occupied by Li\(^+\) ions, \(1 - \theta\) = fraction of anionic sites in the exchanger occupied by Na\(^+\) or H\(^+\) ions.

At \(t = 0\), \(C = C_0\) and \(\theta = 0\), so the initial rate of the process will be given by

\[
\left(-\frac{dC}{dt}\right)_0 = k_a C_0^{n_1} = k_a C^n.
\]

The data in Figs. 3a and 4a were used to calculate the initial rates, which, based on Eq. (5), were plotted as
\[ \ln(\frac{\theta}{\theta_e}) = \ln(\frac{C_0}{C_0 + k_a C_e t}) \] vs \( \ln C \). Based on the fits obtained, \( n = n_1 = 1 \) (with \( R^2 = 0.999 \) for A–Na and \( R^2 = 0.990 \) for A–H). This conclusion was confirmed by fitting, with highly satisfactory results, the same experimental data to the equation

\[ C = \frac{C_0 + k_a C_e t}{1 + k_a t}, \quad (6) \]

which was obtained by integrating the kinetic law, Eq. (4), with \( n_1 = n_2 = n_3 = 1 \), and assuming \( \theta_e = 1 \) at equilibrium, in which case Eq. (5) becomes

\[ -\frac{dC}{dt} = k_a (C - C_e)(1 - \theta) \quad (7) \]

with

\[ \theta = \frac{C_0 - C}{C_0 - C_e}. \quad (8) \]

For various reasons, the experimental value of \( \theta_e \) tends to be less than unity in practice, so the \( k_a \) value obtained from Eq. (6) does not coincide exactly with the specific rate of the adsorption process concerned. However, such a \( k_a \) value, which shall henceforward be denoted by \( k'_a \), can be related to the actual \( k_a \) value for the process via the following equation:

\[ k_a = \frac{k'_a C_e^n}{C_0^n - C_e^n (1 - \theta_e)}. \quad (9) \]

By fitting the experimental values shown in Figs. 3b and 4b to Eq. (6), the \( C_e \) and \( k'_a \) values of Table 1 were obtained. The table also includes the \( k_a \) values calculated from Eq. (9) and those of the corresponding activation enthalpies (\( \Delta H^* \)), free energies (\( \Delta S^* \)), and energies of activation (\( \Delta G^* \)), which were calculated on the assumption that they remained constant throughout the temperature ranges considered and which, based on the theory of the permanent regime of the transition state, should be given by

\[ k = \frac{RT}{Nh} e^{-\Delta G^*/RT} = \frac{RT}{Nh} e^{\Delta S^*/R} e^{-\Delta H^*/RT}, \quad (10) \]

where \( R = \) gas constant, \( N = \) Avogadro’s number, \( h = \) Planck’s constant, and \( T = \) temperature (K).

As can be seen from Table 1, the specific rates obtained for A–Na and A–H are of the same order of magnitude, which is unsurprising taking into account that the textural properties of the two resins and the nature of their exchange sites are virtually identical. Also, this suggests that the cation seemingly exerts no substantial effect on the rate of exchange with amberlite; however, the differences in the \( \Delta H^* \) and \( \Delta S^* \) values suggest that the mass, the electrostatic potential, and the degree of solvation of the initially retained cations (both in the solid and in the activated species through which the process develops) dictate the values of these quantities. Specifically, the differences in \( \Delta H^* \) can be ascribed to significant differences in the degree of hydration of the ions of the activated species. Thus, if solvation of the cations retained by the solid is assumed to be negligible, then the process by
Table 1
Specific rates and thermodynamic activation functions for the A–Na and A–H systems

<table>
<thead>
<tr>
<th>SCE</th>
<th>T (°C)</th>
<th>$C_0 \times 10^3$ (mol/l)</th>
<th>$C_r \times 10^4$ (mol/l)</th>
<th>$k_a' \times 10^3$ (mol/l)</th>
<th>$R^2$</th>
<th>$\theta_e$ ($= n_e^2/n_0^2$)</th>
<th>$k_a \times 10^3$ (mol/l)</th>
<th>$\Delta H^*$ (kJ/mol)</th>
<th>$\Delta S^*$ (J/Kmol)</th>
<th>$R^2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>A–Na</td>
<td>10</td>
<td>1.44</td>
<td>4.59</td>
<td>0.38</td>
<td>0.996</td>
<td>0.69</td>
<td>0.42</td>
<td>41.52</td>
<td>-0.16</td>
<td>0.982</td>
</tr>
<tr>
<td></td>
<td>20</td>
<td>5.49</td>
<td>0.82</td>
<td>0.994</td>
<td>0.58</td>
<td>0.97</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>30</td>
<td>6.39</td>
<td>1.25</td>
<td>0.992</td>
<td>0.45</td>
<td>1.66</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td></td>
<td>40</td>
<td>7.29</td>
<td>1.74</td>
<td>0.993</td>
<td>0.36</td>
<td>2.56</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>A–H</td>
<td>10</td>
<td>1.39</td>
<td>3.92</td>
<td>0.75</td>
<td>0.999</td>
<td>0.98</td>
<td>0.76</td>
<td>18.13</td>
<td>-0.24</td>
<td>0.995</td>
</tr>
<tr>
<td></td>
<td>20</td>
<td>4.19</td>
<td>0.96</td>
<td>0.999</td>
<td>0.88</td>
<td>0.99</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>30</td>
<td>4.45</td>
<td>1.20</td>
<td>0.998</td>
<td>0.77</td>
<td>1.29</td>
<td></td>
<td></td>
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<tr>
<td></td>
<td>40</td>
<td>4.69</td>
<td>1.58</td>
<td>0.998</td>
<td>0.68</td>
<td>1.77</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Notes. SCE = Solid cation exchanger. $k_a$ = specific rate of adsorption of dissolved Li$^+$ ions by the solid. $k_a'$ = specific rate of adsorption of Li$^+$ ions by the solid as calculated using $\theta_e = 1$ as reference. $n_0^2$ = exchange capacity of the solid referred to Li$^+$ ions at the working temperature as calculated from the corresponding equilibrium isotherm. $\Delta H^*$ = enthalpy of activation of the exchange process. $\Delta S^*$ = entropy of activation of the exchange process.

which the activated species is formed can be schematized as

$$A–H + [\text{Li(H}_2\text{O)}_n]^+ + p\text{H}_2\text{O} \rightarrow A–\text{Li[H(H}_2\text{O)}_x]^+, \tag{11}$$

$$A–Na + [\text{Li(H}_2\text{O)}_n]^+ + q\text{H}_2\text{O} \rightarrow A–\text{Li[Na(H}_2\text{O)}_y]^+, \tag{11}$$

where $y < x$ and, possibly, $y < n$, in which case $q < 0$.

The enthalpy of solvation of hydrogen ion offsets that of its release—which is more endothermic than that for Na$^+$ ion—to a greater extent; this results in the A–Na/Li$^+$ system possessing a bigger $\Delta H^*$ value than the A–H/Li$^+$ system. On the other hand, the stronger interaction—via hydrogen bonds—between water molecules in the coordination spheres of the cations involved in the activated species, and also between such cations and the anionic sites in the solid, decrease the number of degrees of freedom, which results in a smaller $\Delta S^*$ value for the A–H/Li$^+$ system.

The previous $k_a$ values are somewhat smaller than those obtained by Navarrete–Casas [4] in the exchange of Na$^+$ ions retained on a synthetic zeolite (Z-10) with an average channel diameter of 10 per Li$^+$ cation; this may have been the result of more markedly hindered motion of the cations in the bulk of the solid resin. In principle, this differential behavior cannot be ascribed to poorer diffusion via a concentration gradient, as the structure of amberlite resins is much more open (i.e., their pores are larger), but rather to stronger interaction between the cations and the anionic sites in the resinate. These assumptions are consistent with the calculated diffusion coefficients for both systems, as well as with the corresponding equilibrium constants [4].

Diffusion coefficients, $D$, can be calculated from kinetic data using various methods such as those of Banerjee et al. [5], Carman and Haul [6] and Aharoni and Suzin [7]; all are approximate in nature and provide qualitative information that is purely estimative and virtually equivalent. In this work, we used that of Carman and Haul [6] and obtained the $D$ values shown in Fig. 5. As can be seen, $D$ was similar for the A–Na/Li$^+$ and A–H/Li$^+$ systems and decreased as Li$^+$ cations were gradually adsorbed; this suggests that the kinetics of the process is governed to a great extent by the diffusion of $[\text{Li(H}_2\text{O)}_n]^+$ cations. These $D$ values are much greater than those found by the authors for the above-
mentioned Z-10/Li$^+$ system [4], which is consistent with the textural differences between Z-10 solids and amberlite. Finally, contrary to the typical findings for diffusion exclusively due to a concentration gradient, $D$ decreased with increasing temperature. This can be ascribed to an increased energy of interaction between the anionic sites in the solid and [Li(H$_2$O)$_n$]$^+$ ions the degree of solvation of which decreases with increasing temperature; this energy increased with increasing temperature, consistent with an endothermic equilibrium process.

3.2. Equilibrium

The experimental results obtained in the study of equilibrium in the ion-exchange processes of the A–Na/Li$^+$ and A–H/Li$^+$ systems at 10, 20, 30, and 40 °C were plotted in the form of $n^s$ vs $C/C_0$ to obtain the isotherms of Fig. 6.

Based on shape, all the $n^s$ vs $C/C_0$ isotherms of Fig. 6 appear to be of the S type in the classification of Giles and Smith [8], which, according to this author, suggests a low relative affinity of the anionic sites of amberlite for lithium ions and vice versa. Also, the mere visual inspection of the isotherms reveals that the exchange process is seemingly more favorable in the A–H/Li$^+$ system. Conversely, the exchange process appears to go further (at equilibrium) in the A–Na/Li$^+$ system. Also, the number of ions exchanged in both processes—at an identical $C/C_0$ ratio—increases with increasing temperature. All these assertions require clarification and justification, however.

If the system and its components (viz., adsorbate, adsorbed, and adsorbed) undergo no substantial changes, then the overall (reversible) adsorption–desorption process must be a single step, the kinetic and equilibrium laws for which must be consistent. In such a case, based on Eq. (4) (where $n_1 = n_2 = n_3 = 1$) and the fact that $–dC/dt = 0$ at equilibrium, the isotherms of Fig. 6 must fit the equation

$$\theta = \frac{n^s}{n^e} = \frac{K_{ci}C_e}{1 + K_{ci}C_e},$$

$$C_e = \frac{1}{K_{ci}n_0^e} + \frac{C}{n_0^e},$$

where $\theta$ = fraction of anionic sites in the exchanger occupied by Li$^+$ cations, $C$ = molar concentration of lithium ions in the dissolved phase, $n^s$ = amount of lithium (mol) retained per gram of amberlite, $n_0^e$ = exchange capacity of amberlite referred to Li$^+$ ion under the working conditions, $K_{ci}$ = kinetic equilibrium constant (= $k_a/k_d$).

Based on Eq. (13), the experimental results of Fig. 6 were also plotted in the form $C/n^s$ vs $C$. The corresponding analytical fitting provided the results given in Table 2.

Our $n_0^e$ values were all less than the previous $2.2 \times 10^{-3}$ eq/g wet resin, which indicates that the ion-exchange process was incomplete under the experimental conditions used, as is usually the case with equilibrium processes. However, the $n_0^e$ values for the A–Na/Li$^+$ system were slightly greater than those for the A–H/Li$^+$ system. This suggests that the exchange of Na$^+$ ions adsorbed on the resin with Li$^+$ ions in solution is more favorable than that with H$^+$ ions; however, based on the $K_{ci}$ values and, especially, the $K$ values obtained—which are commented on below—this assumption is unwarranted on thermodynamic grounds, as H$^+$ ions should be easier to exchange with dissolved Li$^+$ ions. As a result, the fact that $n_0^e$ is no greater for A–Na than for A–H must have a kinetic origin—probably steric hindrances to the diffusion of hydrated cations. In any case, one cannot easily pinpoint the specific origin as the differences in both the size and the free energy of dissolution of the hydrated cations result in easier release and increased mobility of hydrogen ions relative to sodium ions. The reported radii [9] for the hydrated cations involved are as follows: 282 pm for H$^+$, 382 pm for Li$^+$, and 358 pm for

<table>
<thead>
<tr>
<th>Solid</th>
<th>$T$ (°C)</th>
<th>Slope</th>
<th>Intercept</th>
<th>$R^2$</th>
<th>$n_0^e \times 10^{+3}$ (mol/g)</th>
<th>$K_{ci}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>A–Na</td>
<td>10</td>
<td>705.21</td>
<td>1.0222</td>
<td>0.995</td>
<td>1.4</td>
<td>689</td>
</tr>
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<td></td>
<td>20</td>
<td>649.02</td>
<td>1.0454</td>
<td>0.997</td>
<td>1.5</td>
<td>620</td>
</tr>
<tr>
<td></td>
<td>30</td>
<td>563.16</td>
<td>1.1326</td>
<td>0.998</td>
<td>1.8</td>
<td>497</td>
</tr>
<tr>
<td></td>
<td>40</td>
<td>504.46</td>
<td>1.2010</td>
<td>0.998</td>
<td>2.0</td>
<td>420</td>
</tr>
<tr>
<td>A–H</td>
<td>10</td>
<td>987.90</td>
<td>1.1562</td>
<td>0.999</td>
<td>1.0</td>
<td>854</td>
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<tr>
<td></td>
<td>20</td>
<td>913.27</td>
<td>0.9834</td>
<td>0.997</td>
<td>1.1</td>
<td>928</td>
</tr>
<tr>
<td></td>
<td>30</td>
<td>826.73</td>
<td>0.7417</td>
<td>0.996</td>
<td>1.2</td>
<td>1114</td>
</tr>
<tr>
<td></td>
<td>40</td>
<td>740.72</td>
<td>0.5742</td>
<td>0.996</td>
<td>1.4</td>
<td>1290</td>
</tr>
</tbody>
</table>
Na\(^+\). The free energies of dissolution for these ions calculated from the Born equation [10] are \(-648\) kJ/mol for H\(^+\), \(-479\) kJ/mol for Li\(^+\), and \(-375\) kJ/mol for Na\(^+\).

Although the kinetic equilibrium constant, \(K_{ci}\), lacks predictive value for the thermodynamic equilibrium constant, \(K\), it does provide qualitative information for comparison purposes. As can be seen from Table 2, the largest \(K_{ci}\) values were those for the A–H/Li\(^+\) system, where \(K_{ci}\) increased with increasing temperature; this suggests that the ion-exchange process is endothermal. On the other hand, \(K_{ci}\) for the A–Na/Li\(^+\) system decreases with increasing temperature, so the process must be exothermal.

The thermodynamic equilibrium constants, \(K\), for each process were calculated using the simplified equation of Gaynes and Thomas [11],

\[
\ln K = (z_B - z_A) + \ln K_s dN_A,
\]

where \(z_B\) and \(z_A\) = charges of the adsorbed cations (\(B\)) and Li\(^+\), respectively (unity in both cases), \(N_A\) = equivalent fraction of Li\(^+\) in the solid ion exchanger, \(K_s\) = selectivity coefficient, defined as

\[
K_s = \frac{X_A m_B \gamma_B}{X_B m_A \gamma_A}
\]

with \(X_A\) and \(X_B\) = mole fractions of the cations adsorbed in the solid exchanger, \(m_A\) and \(m_B\) = molal concentrations of the cations in the dissolved phase, and \(\gamma_A\) and \(\gamma_B\) = individual activity coefficients for the cations in the dissolved phase.

The mean activity coefficients for the electrolytes HCl, NaCl, and LiCl in an aqueous solution were calculated from the modified Debye–Hückel equation [12]. There values were substituted into the Gluekauf equation [13] to calculate the mean activity coefficients for the electrolytes in the HCl and LiCl aqueous solutions and in the NaCl and LiCl solutions. Finally, the ratio between the individual activity coefficients, \(\gamma_B/\gamma_A\), was calculated from such coefficients:

\[
\frac{\gamma_B}{\gamma_A} = \frac{[\gamma(X_B A_B) \pm \gamma_B^2]/\gamma_X}{[\gamma(X_A A_A) \pm \gamma_A^2]/\gamma_X}.
\]

The \(K\) values thus calculated are shown in Table 3. As can be seen, \(K\) was greater for the cation-exchange process A–H/Li\(^+\) than for the A–Na/Li\(^+\) system. Also, \(K\) increased markedly with increasing temperature in the former and decreased gradually with it in the latter.

Taking into account that

\[
K = e^{-\Delta G^0/RT} = e^{\Delta S^0/R} e^{-\Delta H^0/RT},
\]

the analytical fitting of the regression line obtained by plotting \(\ln K\) vs \((1/T)\) allowed the standard enthalpies and entropies for the studied processes over the working temperature range (10–40\(^\circ\)C) to be calculated. The resulting \(\Delta H^0\) and \(\Delta S^0\) values are shown in Table 3, which also includes the corresponding free energies.

Based on the data in Table 3, the ion exchange in the A–H/Li\(^+\) system is endothermal in nature, whereas in the A–Na/Li\(^+\) system is slightly exothermal. Also, based on Eq. (3) and on the fact that the ion-exchange process involves breakage of the electrostatic bonds between sulfonate anions in the resin and hydrated Na\(^+\) or H\(^+\) cations, as well as the formation of sulfonate–Li\(^+\) bonds, the foregoing can be ascribed to the sulfonate–H\(^+\) bond energy being much higher than that of the sulfonate–Li\(^+\) bond—and to the opposite situation in the case of the initial sulfonate–Na\(^+\) bond with respect to the sulfonate–Li\(^+\) bond. This is consistent with the electrostatic potentials of the corresponding cations.

If his hypothesis be confirmed, it might also account for the decreased \(n_0^0\) values obtained for A–H.

Both processes are endoentropic, particularly that involving the A–H/Li\(^+\) system. The entropy increase arises from the fact that the ion exchange cannot be assigned, especially in the A–H/Li\(^+\) system, to major changes in the number of water molecules present in the hydration spheres of the ions; otherwise, the exchange of H\(^+\) with Li\(^+\) ions would be markedly exoentropic (\(\Delta S^0 < 0\)) and that of Na\(^+\) with Li\(^+\) slightly exoentropic. Probably, the entropy increase observed in both processes arises from the increased disorder resulting from the higher relative mobility of the cations released by the exchanger with respect to the Li\(^+\)/aqua complex initially present in solution [14].

However the results and specific aspects of this work are accounted for, from the foregoing it follows that the proposed model is useful for studying ion-exchange processes at the solid/liquid interface using the same methodology as in adsorption/desorption processes.

References
