Structures of Orthoborate Anions and Physical Properties of Their Lithium Salt Non-aqueous Solutions

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Abstract

We compare the physical properties and solution conductivities of three new lithium orthoborate salts with those of the well known slat lithium bis(trifluoromethanesulfonyl)imide LiTFSI. The three new lithium salts are lithium bis(perfluoropinacolato)borate (LiBPFPB), lithium bis(oxalato)borate (LiBOB) and lithium bis(malonato)borate (LiBMB). Computational models of the three orthoborate anions show that the borate oxygens in BPFPB\textsuperscript{−} anion are the least exposed. The oxygens are electronically identical in BPFPB\textsuperscript{−}, but not in the other anions. The three new lithium salts show conductivities that closely approach those of LiTFSI but show surprising and solvent-dependent orderings. The conductivity is nearly independent of the salt content in the salt concentration range of 0.5 ~ 1M, which is advantageous for their applications. Self-diffusivity measurements for \textsuperscript{7}Li, \textsuperscript{19}F and \textsuperscript{1}H are presented and are consistent with the very high ionic dissociation levels proposed for these salts on the basis of other evidence. Lithium ion

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transport number at room temperature for LiBPFPB in non-aqueous solutions exceeds 0.5 and for LiBMB is about 0.4.

**Introduction**

In recent publications [1-3] we have described some new aromatic-free variants of the orthoborate-based class of electrolytes pioneered by Barthel, Gores and co-workers [4-6]. These salts, in particular those based on dicarboxylate capping of the orthoborate oxygens, show particular promise for replacing the fluorinated lithium salts, such as LiPF$_6$ and Li bis(trifluoromethanesulphonyl) imide (LiTFSI), that are in widespread use at the moment. Not only do the new salts yield non-aqueous solutions of exceptional conductivity, but also they have wide electrochemical windows. In at least two cases they are also non-corrosive and have relatively benign decomposition products. We have given evidence that the bis(oxalato) orthoborate (BOB$^-$) anion is the most weakly associating electrolyte anion so far tested [2]. The exceptional performance of LiBOB in carbonate solutions under testing in commercial type rechargeable Li ion cells is currently being reported elsewhere [7].

The origin of the favorable electrochemical characteristics must lie in the charge delocalization in these anions due to the presence of strongly electron-withdrawing groups attached to the charge centers. To date there have not been any attempts to model this charge distribution. In this paper we present the results of initial attempts to rectify this situation using density functional calculations, and then use these results to interpret observations on the physical properties of these salts in a variety of solvents. We have also extended our recent studies of cation and anion diffusivities of the giant bis(perfluoropinnacolato) orthoborate (BPFPFB$^-$) to the solutions of the smaller bis-malonato-orthoborate (BMB$^-$), using the malonato group protons to follow the motion. Lithium diffusivity results are also presented for the case of LiBOB solutions.

**Experimental**

*Materials.*—Lithium bis(perfluoropinnacolato)borate (LiBPFPB), lithium bis(oxalato)borate (LiBOB) and lithium bis(malonato)borate (LiBMB) were synthesized and characterized
following the methods described previously [1, 2]. The solvents used in this work, such as propylene carbonate (PC), 1,2-dimethoxyethane (DME), acetonitrile (AN), dimethylsulfoxide (DMS), ethylmethylsulfoxide (EMS), dimethylsulfoxide (DMSO) and \textit{N,N}-dimethylformamide (DMF), were dried over calcium hydride or phosphorous pentoxide with heating overnight, distilled at normal or reduced pressure. Solutions of lithium salts in different solvents were prepared in a VAC dry box filled with purified nitrogen, by dissolving the lithium salt in the corresponding solvents.

Attempts were made to determine the stoichiometries of solvates that form in these systems and which limit the ambient temperature composition ranges that can be explored. Solid materials were obtained by slow evaporation of solvents DME and acetonitrile from solutions of LiBPFPB and LiBOB. However the stoichiometries obtained were not rational and the study was not pursued.

\textit{Measurements.}—DSC measurements were carried out using a Perkin-Elmer DSC-7 in sub-ambient mode with liquid nitrogen reservoir. The instrument temperature scale was calibrated with crystal-crystal transition of cyclopentane (\(-151.16^\circ\text{C}\)) and melting of indium (\(+156.60^\circ\text{C}\)). Samples were sealed in aluminum pans, purged with helium gas, and heated at a rate of 10K min\(^{-1}\). The values of glass transition temperature (\(T_g\)) reported are the onset temperatures of the DSC curves during up-scans.

Conductivities were determined by a.c. impedance measurements using a HP 4192A LF Impedance Analyzer in a frequency range from 5 Hz to 13 MHz, using home-made dip-type cells containing two paralleled platinum discs. The cell constants were from 0.7 to 1.3 cm\(^{-1}\), calibrated by 0.1m KCl aqueous solution. The temperature range varied with samples studied, and was increased slowly and continuously during the run. Test measurements at constant temperature yielded the same results within the data noise.

Self-diffusivities were measured using the pulsed field gradient spin echo diffusion method applied in our recent study of dissociation and transport of LiBPFPB in PC and DME [3]. Measurements of \(^7\text{Li}\) diffusions were made for the three salts LiBOB, LiBMB and LiBPFPB dissolved in the solvents PC, DME and DMSO, and \(^{19}\text{F}\) diffusion was made for LiBPFPB in DMSO. Details of the measurements have been given in ref. [3]. Attempts to obtain the self-diffusivity values of the BOB\(^-\) anion in different solutions using the \(^{11}\text{B}\) resonance were not
successful. However the BMB\(^-\) anion could be studied straightforwardly by measuring \(^1\)H NMR of the methylene protons in this anion when the salt was dissolved in perdeuterated DMSO (\textit{i.e.} DMSO-\(d_6\)). The spectra show two separated peaks, the second one due to a small proportion of undeuterated DMSO. This favorable impurity permitted the solvent diffusivity to be compared with that of the anion. All the measurements were based on the calibration of water diffusivity of magnetic fields using the \(2.31 \times 10^{-9}\) m\(^2\)s\(^{-1}\) at 25\(^\circ\)C from Holz and Weingärtner [8]. In our previous study [3] we used the earlier and lower value of the diffusivity of water in ref. [9], \(2.13 \times 10^{-9}\) m\(^2\)s\(^{-1}\) at 25\(^\circ\)C, which is now considered too low. In our tabulated results we have recalculated the earlier diffusivities on the new calibration standard.

**Computations.**—Anion models were created using standard computational methods found in commercially available software. Anion geometries were optimized using B3LYP/6-31+G** pseudospectral calculations (accuracy level: accurate) in Jaguar 4.1 [10]. \(D_2\) symmetry was assumed for BOB\(^-\) and BPFPB\(^-\), but no symmetry was assumed for BMB\(^-\) (four conformers of BMB\(^-\) were created using an automatic MMFF94 search in PC Spartan Pro, v. 1.0.5 [11], but B3LYP/6-31+G** geometry optimization of these conformers always yielded the same structure). The electronic structures of the geometry-optimized anions were then calculated using B3LYP/6-311+G** calculations in Linux Spartan’02 [12]. The B3LYP computational method is a hybrid density functional method. The 6-311+G** basis set is a split-valence+polarization basis set that adds a set of diffuse valence functions to all atoms in order to model the effects of increased electron-electron repulsion encountered in anions. Electrostatic potential maps of 0.002 au isoelectron density surfaces were calculated using Linux Spartan’02.

**Results and Discussion**

**Computations results: Structure and electrostatic maps.**—One way to compare the electronic character of these ions is to compare electrostatic potential maps of the ions’ 0.002 au isoelectron density surfaces (these surfaces typically have the same dimensions as a standard space-filling model). Maps for the three orthoborate anions, BPFPB\(^-\), BOB\(^-\) and BMB\(^-\), are shown in Figure 1 (BOB\(^-\) upper left, BMB\(^-\) upper right, BPFPB\(^-\) bottom). The maps are
displayed using the same color-energy scale so that colors in different maps can be used to assess and compare map potentials (See supplementary materials. Colored maps are also available from the authors on request). The map potentials are coded red ($\leq -106$ kcal/mol) < orange < yellow < green ($= -73$ kcal/mol) < blue ($\geq -40$ kcal/mol). Although the potential is negative everywhere on each map, the most negative potentials (red and orange) are found around, and in between, oxygens. The fluorines in BPFPB$^-$, on the other hand, are considerably less negative. There are some subtle variations between oxygens in BOB$^-$ and BMB$^-$. The most negative potentials in BOB$^-$ are found near the terminal oxygens. The most negative potentials in BMB$^-$ are found near the terminal oxygens, and around the two boron-bound oxygens that approach each other closely (these oxygens appear to lie in a row with two terminal oxygens). The oxygens in BPFPB$^-$ generate potentials similar to those found in the other anions. However, in BPFPB$^-$ these oxygens are less exposed.

**Figure 1.** Electrostatic potential maps of all three anions (BOB$^-$ upper left, BMB$^-$ upper right, BPFPB$^-$ bottom). The oxygens in BPFPB$^-$ are least exposed. The oxygens bonded to boron are slightly less negative in BOB$^-$ (yellow-orange); the most negative regions (red) in this anion lie between pairs of terminal oxygens.
The optimized geometries of the three orthoborate anions are listed in Table I. Note that the italicized O-B-O bond angles are between oxygens in the same bidentate ligand.

**Table I.** Optimized geometry parameters of the three orthoborate anions

<table>
<thead>
<tr>
<th>Anion</th>
<th>B-O bond distance (Å)</th>
<th>O-B-O bond angle (°)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>intra-bidentate ligand</td>
</tr>
<tr>
<td>BPFPB⁻</td>
<td>1.482</td>
<td>102.9</td>
</tr>
<tr>
<td>BOB⁻</td>
<td>1.476</td>
<td>105.6</td>
</tr>
<tr>
<td>BMB⁻</td>
<td>1.458, 1.459, 1.490, 1.490</td>
<td>111.9, 111.9</td>
</tr>
</tbody>
</table>

Information on electron density distribution, which may correlate with both cohesive energy and ion dissociation in the solutions, is also available from the $^{11}$B NMR chemical shifts for these salts. From our previous reports [1, 2], the $^{11}$B chemical shifts of LiBPFPB, LiBOB and LiBMB in DMSO-$d_6$ (reference to BF$_3$·Et$_2$O in CDCl$_3$) is $\delta$ 16.14ppm, 12.20ppm and 8.26ppm, respectively. Thus, the charge density at the boron nucleus is lowest in the case of LiBPFPB. Note that trends in charge density at the nucleus cannot be inferred from the electrostatic potential data shown in Figure 1 because potentials are derived from the total charge density distribution. The ionic dissociation in solutions, judged by the ionic transport numbers given below, qualitatively falls in the same order.

**Glass Transition Temperatures of solutions.**—As described previously [1, 2], LiBPFPB and LiBOB are very soluble in common organic solvents. It is known that high solubility frequently correlates with low melting point through their common dependence on a low salt lattice energy. Unfortunately, however, the melting point of LiBOB is not low enough to be observed before decomposition occurs at 300ºC. We have also been unable to observe fusion in any binary system either of type LiBOB-LiX, or LiBOB-MBOB where X is a second anion and M is a second alkali ion or even a large fusible organic cation. So no solvent-free glasses containing LiBOB or LiMOB have been obtained so far.
Glass transitions can however be observed in the binary solutions with solvents such as DMSO in which the solubility LiBOB is quite high, reaching almost 30 mol%. By comparison with data on LiClO$_4$ in the same solvent from ref. [13], it can be seen (in Figure 2a) that the glass transition temperatures ($T_g$) of LiBOB solutions are high, and indeed higher than we would have expected from its large anion size and correspondingly expected low coulomb energy – e.g. relative to LiClO$_4$. This will lead to correspondingly viscous solutions, and lower ionic mobilities, so LiBOB will be found most suitable as an electrolyte at lower salt contents where the very small ionic association characteristic will be most advantageous. The perfluorinated orthoborate, on the other hand, yields solutions in PC whose $T_g$ values are only slightly higher than those of BF$_4^-$ and TFSI$^-$ anions Li salts (Figure 2b), despite its much greater molar mass. A full accounting of the factors determining the value of $T_g$ in these salt-containing systems still remains to be formulated.

![Graphs showing glass transition temperatures of LiBOB and LiBPFPB in binary solutions](image)

**Figure 2.** Plots of glass transition temperatures of LiBOB and LiBPFPB in binary solutions with a common component, compared with those of other common salts. The extrapolation of the plots help reach the glass transition temperatures for individually non-glassforming substances.

Figure 2b also shows that beyond 42 mol% LiBPFPB, the value of $T_g$ does not increase. It is possible that this is due to experimental problems but it seems equally likely that at this
composition the anions have begun to cluster in anticipation of the quasi-molecular state that exists in the pure state (we recall that the salt distills over at about 170°C under reduced pressure [1]). The reasons for this must be basically steric but are not well understood by us. It may be related to the manner in which the small Li⁺ cation can get “tucked away” in the negatively charged crevice of the anion. (See Figure 1).

_Ion conductivity._—The temperature dependence of ionic conductivity of LiBPFPB-PC and LiBOB-PC binary systems are shown in Figure 3 and Figure 4, respectively. All the plots show non-Arrhenius type of ion transport behavior, as expected. The 0.5M PC solutions of both salts exhibit the highest conductivity at ambient and sub-ambient temperatures. The isothermal conductivities of these systems between −25 and 75°C based on Figure 3 and Figure 4 data, are plotted against salt concentration in Figure 5 and Figure 6.

**Figure 3.** Temperature dependence of ionic conductivity of LiBPFPB-PC binary system. The salt compositions given in mole percentage in the Figure legend correspond to the following molar concentrations: 0.08M, 0.5M, 1M, 1.6M, 1.97M, 2.15M and 2.57M (the latter is for the pure molten salt).
Figure 4. Temperature dependence of ionic conductivity of LiBOB-PC binary system.

Figure 5. Isothermal conductivity of LiBPFPB-PC system with salt concentration at different temperatures.
The plots show the usual temperature dependent maximum with increase of salt concentration. The maximum conductivity appears at salt content between 0.5M and 1M. Such variation of conductivity with salt concentration is determined by the balance of opposing effects on conductivity of free ion number and ion mobility. The maximum conductivity moves to higher concentrations for higher isothermal temperatures. This phenomenon is expected from the fact that mobility is strongly temperature dependent, but free ion concentration is not. For fully dissociated salts, such as LiBPFPB, LiBOB and LiBMB in PC, it can be simply modeled using a VFT equation with linear composition dependence of $T_o$ [14].

It is seen that when the temperature is at and above 25°C, the conductivity of both solutions becomes almost independent of concentration in the range of 0.5 ~ 1M, a feature that could be advantageous for the applications of such salts. The same behavior is also found for the variation of the isothermal conductivity of LiBMB-DMSO solutions with salt concentration in the range of 0.5 to 1.3M as shown in Figure 7. Increasing the concentration of LiBMB in DMSO decreases

**Figure 6.** Isothermal conductivity of LiBOB-PC system with salt concentration at different temperatures.
the temperature at which solvate crystals form in LiBMB-DMSO solutions, from around 18°C for 0.2M solution, to 10°C for 0.5M solution, 6°C for 0.7M and 0.9M solutions, and to 0°C for 1.1M and 1.3M solutions.

Figure 7. Variation of isothermal conductivity of LiBMB-DMSO solutions with molar salt concentration.

Figure 8 shows the temperature dependence of conductivity of LiBPFPB-DME solutions at different salt concentrations. The breaks at $\frac{1000}{T} \approx 3.60$ (i.e. $t \approx 5^\circ C$) for 0.6M solution and $4.17$ (i.e. $t \approx -33^\circ C$) for 0.4M solution are due to the crystallization of solvate during cooling. There is no such break for 0.2M solution in the temperature studied (−60 to 60°C). Fortunately, this solvate crystallization phenomenon does not appear in the LiBOB-DME binary systems at any salt concentration up to 1.2M. Results for these interesting solutions are shown in Figure 9.
Figure 8. Temperature dependence of ionic conductivity of LiBPFPB-DME solutions with different salt concentrations.

Figure 9. Temperature dependence of ionic conductivity of LiBOB-DME solutions with different salt concentrations.
Figure 10. Comparison of ionic conductivity of LiBPFPB, LiBOB, LiBMB and LiTFSI in different solvents.

Figure 10 gives comparisons of the ionic conductivity of LiBPFPB, LiBOB and LiBMB with that of the commonly used lithium salt LiTFSI, in 0.5M EMS-DMS solutions (a), 0.5M DMSO (b) and 0.08M PC (c).
DMSO solutions (b) and in 0.08M PC solutions (c), respectively. EMS-DMS is a room temperature stable sulfone solvent of exceptional anodic stability which has been developed in this laboratory [15]. It contains 15mol% of DMS to lower the melting point of EMS to ambient. Salt solutions remain uncrystallized to much lower temperatures. It is seen that LiTFSI yields the highest conductivities regardless of solvent, though in the dilute PC solutions the difference becomes marginal. LiBMB and LiBOB are equally conductive in DMSO but rather different in EMS-DMS, for reasons that are not obvious at this time. Except in DME all salts appear to be completely dissociated, according to the diffusivity conductance relations that we now examine.

Self-diffusivity and lithium ion transport number.—The diffusivities are obtained from the slopes of the Stejskal-Tanner plots shown in Figure 11. Except for the data for LiBPFPB in PC and DME reported earlier [3], the self-diffusivities have only been determined at ambient temperature. Our main object was to obtain information on transport numbers for comparison with finding for less dissociating salts in non-aqueous electrolyte solutions [16].

![Figure 11](image-url)

**Figure 11.** Stejskal-Tanner plots for (a) $^7$Li and (b) anion ($^{19}$F and $^1$H) for 0.5M lithium orthoborate salts in different solvents. Note strict linearity over more than one order in log($S/S_o$).

Figure 11 (a) and (b) show respectively the data of lithium and anion species of the three lithium orthoborate salts in different solvents with a concentration of 0.5M at room temperature.
For comparison, the plots for LiBPFPB in PC and DME at 30°C from ref. [3] are also included. The linearity in these semi logarithmic plots over more than one order of magnitude in echo amplitude demonstrates the homogeneity of the pulses. It is seen that the decays of the Stejskal-Tanner plots, $S/S_0$ vs. $\delta^2(\Delta-\delta/3)$, for lithium, fluorinated species and protonated species are all independent of time $\Delta$, the refocusing time interval between the two gradient pulses at the Stejskal-Tanner pulse sequence. This means that self-diffusion coefficient $D$ for these species is indeed constant.

The results calculated from the slopes of the plots shown in Figure 11 (a) and (b), for both the Li cation and the anions (represented by F- and H-containing species), are presented for the salts LiBPFPB and LiBMB. For LiBOB only the cation diffusivity is reported. The self-diffusivity values are listed in Table II, which also contains data for the solution ionic conductivities, both directly measured and calculated from diffusivities. The calculated values are obtained by assuming that the Nernst-Einstein equation should be exact for fully dissociated salts, even though this is known to be incorrect except at infinite dilution of mobile species [17, 18].

It is only for LiBMB in the solvent DMSO-$d_6$ that a complete set of lithium, anion and solvent diffusivities are available. For this case we observe that the diffusivity of the anion remains the larger of the two ionic species while the solvent diffuses faster than either by a factor of ca. 2. The calculated conductivity exceeds the measured value by 50%, which is a larger deviation than for the "giant anion" salt, LiBPFPB, but is still small relative to values encountered for more common salts in moderate dielectric constant solvents ($\varepsilon = 47.2$ at 25°C for normal DMSO). This finding confirms the highly dissociated nature of this salt in moderate dielectric constant solvents. Supporting this conclusion is the similarity of the $\sigma_{Li,NE}/\sigma_{exp}$ values (col. 6 in Table II) for LiBMB and LiBPFPB, 0.66 and 0.68, respectively.

For the lithium ion transport number $t_{Li^+}$, we obtain, using the definition $t_{Li^+} = D_{Li^+} / (D_{Li^+} + D_{BMB-})$, the value of 0.43. Making an exaggerated correction for ion-pairing, we obtain $t_{Li^+} = 0.39$. Both are smaller than found in the perfluorinated “giant anion” case in non-deuterated DMSO, $t_{Li^+} = 0.55$ (Table II).

For LiBOB only $D_{Li}$ is known but its Nernst-Einstein conductivity accounts for 0.76 of the total in DMSO of the measured value and 0.60 in PC, comparable to that for LiBMB (Table II, 6th column, $\sigma_{Li,NE}/\sigma_{exp}$). We can also deduce that $t_{Li^+}$ for LiBOB in non-aqueous solutions is
similar to that for LiBMB because the two anions are similar in structure and size. The implied high degree of dissociation confirms the conclusions of our earlier work based on conductivity vs. temperature data [2].

**Table II.** Self-diffusivities, lithium ion transport number and degree of dissociation of lithium orthoborate salts in different solvents

<table>
<thead>
<tr>
<th>Solvent</th>
<th>T (°C)</th>
<th>$D_{Li}$ (m$^2$s$^{-1}$)</th>
<th>$D_{anion}$ (m$^2$s$^{-1}$)</th>
<th>$\sigma_{exp}$ (mScm$^{-1}$)</th>
<th>$\sigma_{Li,NE}/\sigma_{exp}$</th>
<th>$\sigma_{NE}/\sigma_{exp}$</th>
<th>$\xi a$ (m$^2$s$^{-1}$)</th>
<th>$D_{Li}/(D_{Li} + D_{anion})$</th>
<th>$t_{Li+} b$</th>
<th>$\alpha c$</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.5M LiBPFPB</td>
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</tr>
<tr>
<td>DMSO</td>
<td>25</td>
<td>1.76E-10</td>
<td>1.50E-10</td>
<td>4.86</td>
<td>0.68</td>
<td>1.26</td>
<td>2.59</td>
<td>0.54</td>
<td>0.55</td>
<td>0.78</td>
</tr>
<tr>
<td>PC</td>
<td>30</td>
<td>1.21E-10$^d$</td>
<td>1.09E-10$^d$</td>
<td>3.92</td>
<td>0.57$^d$</td>
<td>1.08$^d$</td>
<td>2.12$^d$</td>
<td>0.53$^d$</td>
<td>0.53$^d$</td>
<td>0.92$^d$</td>
</tr>
<tr>
<td>DME</td>
<td>30</td>
<td>5.37E-10$^d$</td>
<td>4.79E-10$^d$</td>
<td>9.87</td>
<td>1.01$^d$</td>
<td>1.90$^d$</td>
<td>5.34$^d$</td>
<td>0.53$^d$</td>
<td>0.55$^d$</td>
<td>0.50$^d$</td>
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<tr>
<td>0.5M LiBOB</td>
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<tr>
<td>DMSO</td>
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<td>—</td>
<td>5.24</td>
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<td>—</td>
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<tr>
<td>PC</td>
<td>25</td>
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<td>4.14</td>
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<tr>
<td>DME</td>
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<tr>
<td>0.5M LiBMB</td>
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<tr>
<td>DMSO</td>
<td>25</td>
<td>1.74E-10</td>
<td>—</td>
<td>4.99</td>
<td>0.66</td>
<td>—</td>
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<tr>
<td>DMSO-d$_6$</td>
<td>25</td>
<td>1.52E-10</td>
<td>2.04E-10$^e$</td>
<td>4.45</td>
<td>0.64</td>
<td>1.50</td>
<td>2.37</td>
<td>0.43</td>
<td>0.39</td>
<td>0.71</td>
</tr>
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</table>

Notes:

$^a \xi = RT\sigma_{exp}/(cF^2)$

$^b t_{Li+} = (D_{Li} - D_{anion} + \xi)/2\xi$

$^c \alpha = (\xi - D_{Li})/2D_{anion} + 0.5$. The value of $\alpha$ is a lower bound because it is calculated by assuming that the Nernst-Einstein equation is an exact relation for concentrated ionic systems, which is certainly incorrect. Theoretically deduced deviations (ref. 15) due to inter-ionic friction amount to some 30% at $T > T_o$, so any salt with $\alpha > 0.7$ by this analysis should be considered to be fully dissociated.

$^d$ These values differ from those recently published in ref. 3 because of a change in calibration standard. We have recalibrated using the diffusivity of water given by ref. 8 instead of the earlier and lower value of ref. 9.

$^e$ The presence of two proton resonances in the case of LiBMB in DMSO-$d_6$, one due to normal impurities, allowed us to approximately determine the solvent diffusivity, which was ca. $2D_{BMB}$.
Conclusions

The molecular structures of the minimum energy configurations of three new orthoborate anions, BPFPB\(^-\), BOB\(^-\) and BMB\(^-\), have been calculated using density functional method. The comparisons of the physical properties and solution conductivities of their lithium salts with those for LiTFSI show that these three salts have solution conductivities closely approaching LiTFSI but show solvent-dependent orderings. The conductivity is nearly independent of the salt content in the salt concentration range of 0.5 ~ 1M, which is advantageous for their applications. Self-diffusivity of \(^7\)Li, \(^19\)F and \(^1\)H for these three salts are consistent with the very high ionic dissociation levels proposed for these salts on the basis of other evidence.

Acknowledgment

This work was supported by the Department of Energy under contract no. DEFG0393ER14378-003, supplemented by contract no. DEFG0395ER45541. The authors would like to thank Dr. Burkhard Geil in Technische Universität Darmstadt, Germany for useful discussions and Ms. Emily McDaniel in the Department of Physics, ASU for some NMR measurements. We also wish to thank Jason Sievers, a REV student in the department of Physics, who was supported under NSF REV grant number DMR 9988056.

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