

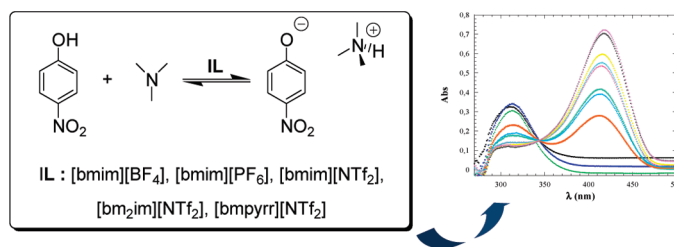
## Determination of Basic Strength of Aliphatic Amines through Ion Pair Formation in Some Ionic Liquid Solutions

Francesca D'Anna,\* Paola Vitale, and Renato Noto\*

Dipartimento di Chimica Organica "E. Paternò", Università degli Studi di Palermo, Viale delle Scienze-Parco d'Orleans II, 90128 Palermo, Italy

fdanna@unipa.it; rnoto@unipa.it

Received June 3, 2009



To have an evaluation of the basic strength of aliphatic amines in ionic liquid solution, the stability constants relevant to the formation of amine/*p*-nitrophenol ion pairs were determined in different ionic liquids at 298 K. In particular, aliphatic (pyrrolidinium) and aromatic (imidazolium) ionic liquids were used. Imidazolium ions,  $\text{bmim}^+$  and  $\text{bm}_2\text{im}^+$ , having different hydrogen bond donor abilities were taken into account. Anions were chosen ( $[\text{BF}_4^-]$ ,  $[\text{PF}_6^-]$ , and  $[\text{NTf}_2^-]$ ; where  $\text{NTf}_2 =$  bis(trifluoromethanesulfonyl)imide) showing different shape, size, and coordination ability. Several primary, secondary (cyclic or not), and tertiary amines were used to study the effect of amine structure on the ion pair stability. The comparison between data collected in this work and those previously obtained in ordinary molecular solvents evidence that ionic liquids are solvent media able to exalt the amine basicity. Furthermore, the experimental trends obtained in different ionic liquids can only be rationalized considering the whole of parameters affecting their tridimensional structure.

### Introduction

Chemistry research has been characterized in the last years by a burgeoning interest in room temperature ionic liquids (RTILs).<sup>1</sup> From an industrial point of view, it has been generally expected that they may be used as substitution solvents for volatile organic solvents. The plethora of papers published up to now on this

subject, have tried to shed light on their peculiar features coming from the large number of cation–anion combinations, giving rise to completely different solute–solvent interactions than those detected in conventional solvents. As a consequence, RTILs are able to significantly affect solute properties and exert particular effects when employed as reaction media.<sup>2</sup>

\*To whom correspondence should be addressed. F.D. and R.N.: Phone: +39091596919. Fax: +39091596825.

(1) (a) Welton, T. *Chem. Rev.* **1999**, *99*, 2071–2083. (b) Wasserscheid, P.; Keim, M. *Angew. Chem., Int. Ed.* **2000**, *39*, 3772–3789. (c) Rogers, R. D.; Seddon, K. R. *Ionic Liquids: Industrial Application to Green Chemistry*; ACS Symp. Ser. No. 818; American Chemical Society: Washington, DC, 2002. (d) Rogers, R. D.; Seddon, K. R. *Ionic Liquids as Green Solvents. Progress and Prospects*; ACS Symp. Ser. No. 856; American Chemical Society: Washington, DC, 2003. (e) Rogers, R. D.; Seddon, K. R.; Volkov, S. *Green Industrial Application of Ionic Liquids*; NATO Science Series II: Mathematics, Physics and Chemistry; Kluwer: Dordrecht, The Netherlands, 2003; Vol. 92. (f) Chiappe, C.; Pieraccini, D. *J. Phys. Org. Chem.* **2005**, *18*, 275–297. (g) Rogers, R. D.; Seddon, K. R. *Ionic Liquids III A: Fundamentals, Progress, Challenges and Opportunities*; ACS Symp. Ser. No. 901; American Chemical Society: Washington, DC, 2005. (h) Harper, J. B.; Kobrak, M. N. *Mini-Rev. Org. Chem.* **2006**, *3*, 253–269. (i) Malhotra, S. V. *Ionic Liquids in Organic Synthesis*; ACS Symp. Ser. No. 950; American Chemical Society: Washington, DC, 2005. (j) Chowdhury, S.; Mohan, R. S.; Scott, J. L. *Tetrahedron* **2007**, *63*, 2363–2389. (k) *Ionic Liquids in Synthesis*; Wasserscheid, P.; Welton, T., Eds.; Wiley-VCH: Weinheim, Germany, 2008. (l) Leclercq, L.; Schmitzer, A. R. *Supramol. Chem.* **2009**, *21*, 245–263. (m) Toma, S.; Meciarová, M.; Sebesta, R. *Eur. J. Org. Chem.* **2009**, 321–327.

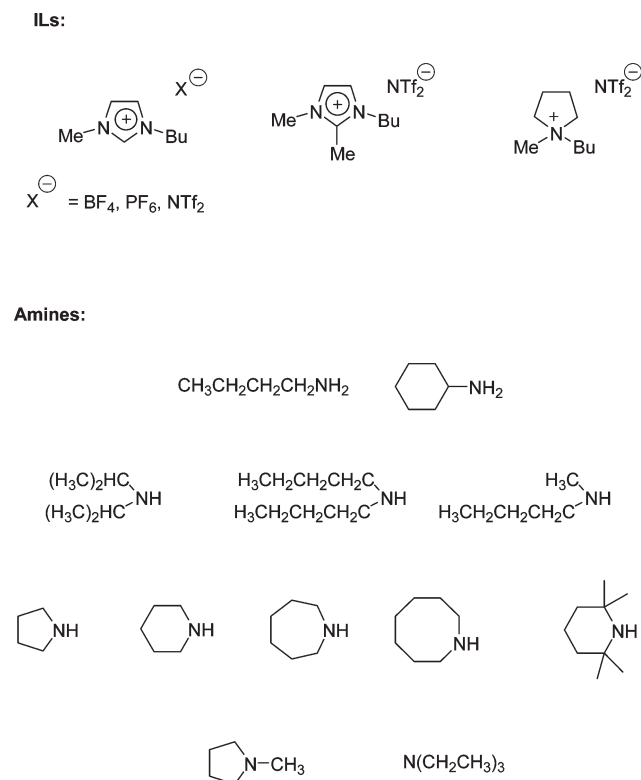
(2) (a) Sheldon, R. *Chem. Commun.* **2001**, 2399–2407. (b) Lancaster, N. L.; Welton, T.; Young, G. B. *J. Chem. Soc., Perkin Trans. 2* **2001**, 2267–2270. (c) Chiappe, C.; Conte, V.; Pieraccini, D. *Eur. J. Org. Chem.* **2002**, 2831–2837. (d) Lancaster, N. L.; Salter, P. A.; Welton, T.; Young, G. B. *J. Org. Chem.* **2002**, *67*, 8855–8861. (e) Chiappe, C.; Pieraccini, D.; Saullo, P. *J. Org. Chem.* **2003**, *68*, 6710–6715. (f) Chiappe, C.; Pieraccini, D. *J. Org. Chem.* **2004**, *69*, 6059–6064. (g) Akiyama, T.; Suzuki, A.; Fuchibe, K. *Synlett* **2005**, 1024–1026. (h) Ranu, B. C.; Jana, R. *J. Org. Chem.* **2005**, *70*, 8621–8624. (i) Conte, V.; Floris, B.; Galloni, P.; Mirruzzo, V.; Scarsio, A.; Sordi, D.; Strukul, G. *Green Chem.* **2005**, *7*, 262–266. (j) Man, B. Y. W.; Hook, J. M.; Harper, J. B. *Tetrahedron Lett.* **2005**, *46*, 7641–7645. (k) Lindén, A. A.; Johansson, M.; Hermanns, N.; Bäckvall, J.-E. *J. Org. Chem.* **2006**, *71*, 3849–3853. (l) Zhao, X.; Alper, H.; Yu, Z. *J. Org. Chem.* **2006**, *71*, 3988–3990. (m) Chiappe, C.; Piccioli, P.; Pieraccini, D. *Green Chem.* **2006**, *8*, 277–281. (n) Tiwari, S.; Khupse, N.; Kumar, A. *J. Org. Chem.* **2008**, *73*, 9075–9083. (o) Wells, T. P.; Hallett, J. P.; Williams, C. K.; Welton, T. *J. Org. Chem.* **2008**, *73*, 5585–5588. (p) Yau, H. M.; Barnes, S. A.; Hook, J. M.; Youngs, T. G. A.; Croft, A. K.; Harper, J. B. *Chem. Commun.* **2008**, 3576–3587. (q) Zicmanis, A.; Katkevica, S.; Mekss, P. *Catal. Commun.* **2009**, *10*, 614–619. (r) Baciocchi, E.; Chiappe, C.; Del Giacco, T.; Fasciani, C.; Lanzalunga, O.; Lapi, A.; Melai, B. *Org. Lett.* **2009**, *11*, 1413–1416. (s) Hallett, J. P.; Liotta, C. L.; Ranieri, G.; Welton, T. *J. Org. Chem.* **2009**, *74*, 1864–1868. (t) Harper, J. B.; Rosella, C. E. *Tetrahedron Lett.* **2009**, *50*, 992–994.

We have been particularly interested in the last topic, and we have taken into consideration some classical organic reactions that are affected by base catalysis as probe reactions.<sup>3</sup> We have frequently used aliphatic amines as base catalysts, and in several cases we have observed a lack in correlation between our kinetic data in ionic liquid (IL) solution and  $pK_a$  values collected for these catalysts in water solution.<sup>3d,g</sup> In agreement with a referee's comment,<sup>4</sup> we believe that lack of  $pK_a$  values determined in IL solution could represent a limitation to the full understanding of data collected in these solvent systems. A recent paper on this issue has pointed out that particular caution is needed in regard to possible mistakes deriving from the use, in IL solution, of  $pK_a$  values collected in aqueous solution.<sup>5</sup> Indeed, it is well-known that thermodynamic parameters of a solute are strongly affected by solvent properties.

In our opinion, the determination of amines basicity in these solvent media could prove interesting also because the analysis of the papers published in the past few years about ILs shows that aliphatic amines have been frequently used both as nucleophiles and as bases.<sup>3b,d,g,6</sup> In particular, in different cases the results obtained evidenced higher basicity of aliphatic amines in IL solution than in conventional organic solvents, outlining the significant differences in the solvation processes occurring in these solvent media.<sup>3a,c-e,6b</sup>

In this light, and bearing in mind some preliminary results recently obtained by us,<sup>4</sup> we chose the equilibrium of amine/ $p$ -nitrophenol ion pair formation as a suitable probe to have a measurement of the basic strength of aliphatic amines in IL solution (Chart 1). By analogy with basicity measurements carried out in organic solvents, such as benzene,<sup>7</sup> 1,4-dioxane,<sup>8</sup> or tetrahydrofuran,<sup>9</sup> we determined the stability constants by spectrophotometric titration at 298 K. We took into account ILs differing both in anion and cation properties, such as [bmim][BF<sub>4</sub>], [bmim][PF<sub>6</sub>], [bmim][NTf<sub>2</sub>], [bm<sub>2</sub>m][NTf<sub>2</sub>], and [bmpyrr][NTf<sub>2</sub>] (Chart 1) [where bmim = 1-butyl-3-methylimidazolium; bm<sub>2</sub>m = 1-butyl-2,3-dimethylimidazolium; bmpyrr = *N*-butyl-*N*-methylpyrrolidinium; and NTf<sub>2</sub> = bis(trifluoromethylsulfon)imide]. Chosen ILs differ in the anion size, symmetry, and coordination ability, but also in the cation ability to give hydrogen bond,  $\pi$ - $\pi$ , and  $\pi$ -cation interactions.

CHART 1. Structures of Used Amines and Ionic Liquids



It is well-known that all these features significantly contribute to determine the properties of ILs. Indeed, changes in the cation–anion combination give rise to significant variations in classical solvent parameters such as  $\alpha$ ,  $\beta$ ,  $\pi^*$ , etc.<sup>10</sup> However, they could also affect the tridimensional structure of these media that are frequently depicted as polymeric supramolecular fluids in which constituent ions interact by means of Coulomb's interactions, hydrogen bond, and  $\pi$ - $\pi$  interactions.<sup>11</sup>

As far as amines are concerned, they were chosen on the grounds of their basicity (see later) in aqueous solution, structure, and steric hindrance (Chart 1). Some different primary, secondary, and tertiary aliphatic amines were used. In particular, for secondary amines both acyclic and cyclic structures with different ring size and conformational freedom were taken into account. All these factors could affect not only the equilibrium position but also the ability of the

(3) (a) D'Anna, F.; Frenna, V.; Noto, R.; Pace, V.; Spinelli, D. *J. Org. Chem.* **2005**, *70*, 2828–2831. (b) D'Anna, F.; Frenna, V.; Noto, R.; Pace, V.; Spinelli, D. *J. Org. Chem.* **2006**, *71*, 5144–5150. (c) D'Anna, F.; Frenna, V.; Noto, R.; Pace, V.; Spinelli, D. *J. Org. Chem.* **2006**, *71*, 9637–9642. (d) D'Anna, F.; Frenna, V.; Pace, V.; Noto, R. *Tetrahedron* **2006**, *62*, 1690–1698. (e) D'Anna, F.; Frenna, V.; La Marca, S.; Noto, R.; Pace, V.; Spinelli, D. *Tetrahedron* **2008**, *64*, 672–680. (f) D'Anna, F.; Frenna, V.; Marullo, S.; Noto, R.; Spinelli, D. *Tetrahedron* **2008**, *64*, 11209–11217. (g) D'Anna, F.; La Marca, S.; Noto, R. *J. Org. Chem.* **2008**, *73*, 3397–3403. (h) D'Anna, F.; Marullo, S.; Noto, R. *J. Org. Chem.* **2008**, *73*, 6224–6228.

(4) D'Anna, F.; Noto, R. *Tetrahedron* **2007**, *63*, 11681–11685.

(5) Johnson, K. E.; Pagni, R. M.; Bartmess, J. *Monatsh. Chem. Chem. Monthly* **2007**, *138*, 1077–1101.

(6) (a) Yadav, J. S.; Reddy, B. V. S.; Basak, A. K.; Navsaiah, A. V. *Tetrahedron Lett.* **2003**, *44*, 2217–2220. (b) Mečiarová, M.; Toma, S. *Chem.—Eur. J.* **2007**, *13*, 1268–1272. (c) Mečiarová, M.; Cigán, M.; Toma, S.; Gáplovský, A. *Eur. J. Org. Chem.* **2008**, 4408–4411. (d) Ranieri, G.; Hallett, J. P.; Welton, T. *Ind. Eng. Chem. Res.* **2008**, *47*, 638–644.

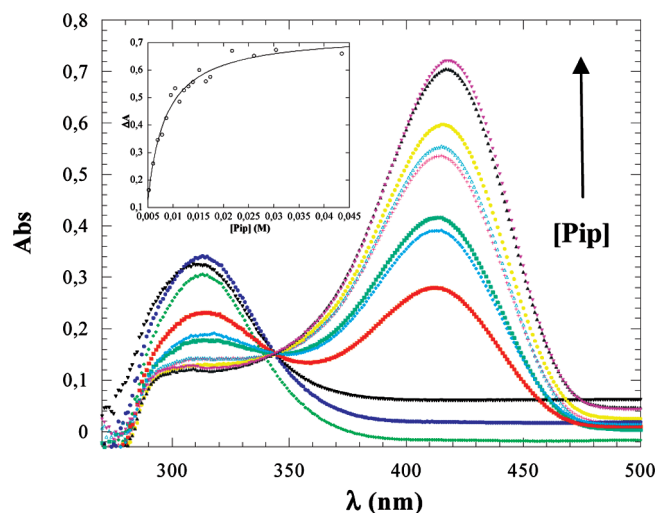
(7) Frenna, V.; Vivona, N.; Consiglio, G.; Spinelli, D. *J. Chem. Soc., Perkin Trans. 2* **1985**, 1865–1868.

(8) Pearson, R. G.; Vogelson, D. C. *J. Am. Chem. Soc.* **1958**, *80*, 1038–1043.

(9) Furlong, J. J. P.; Lewkowicz, E. S.; Nudelman, N. S. *J. Chem. Soc., Perkin Trans. 2* **1990**, 1461–1465.

(10) (a) Carmichael, A. J.; Seddon, K. R. *J. Phys. Org. Chem.* **2000**, *13*, 591–595. (b) Muldoon, M. J.; Gordon, C. M.; Dunkin, I. R. *J. Chem. Soc., Perkin Trans. 2* **2001**, 433–435. (c) Crowhurst, L.; Mawdsley, P. R.; Perez-Arlandis, J. M.; Salter, P. A.; Welton, T. *Phys. Chem. Chem. Phys.* **2003**, *5*, 2790–2794.

(11) (a) Hanke, C. G.; Johansson, A.; Harper, J. B.; Lynden-Bell, R. M. *Chem. Phys. Lett.* **2003**, *374*, 85–90. (b) Holbrey, J. D.; Reichert, W. M.; Nieuwenhuyzen, M.; Sheppard, O.; Hardacre, C.; Rogers, R. D. *Chem. Commun.* **2003**, 476–477. (c) Harper, J. B.; Lynden-Bell, R. M. *Mol. Phys.* **2004**, *102*, 85–94. (d) Dupont, J. *J. Braz. Chem. Soc.* **2004**, *15*, 341–350. (e) Gozzo, F. C.; Santos, L. S.; Augusti, R.; Consorti, C. S.; Dupont, J.; Eberlin, M. *Chem.—Eur. J.* **2004**, *10*, 6187–6193. (f) Deetlefs, M.; Hardacre, C.; Nieuwenhuyzen, M.; Sheppard, O.; Soper, A. K. *J. Phys. Chem. B* **2005**, *109*, 1573–1598. (g) Lachwa, J.; Bento, I.; Duarte, M. T.; Lopes, J. N. C.; Rebelo, L. P. N. *Chem. Commun.* **2006**, 2445–2447. (h) Lynden-Bell, R. M.; Del Pópulo, M. G.; Youngs, T. G. A.; Kohanoff, J.; Hanke, C. G.; Harper, J. B.; Pinilla, C. C. *Acc. Chem. Res.* **2007**, *40*, 1138–1145. (i) Hardacre, C.; Holbrey, J. D.; Nieuwenhuyzen, M.; Youngs, T. G. A. *Acc. Chem. Res.* **2007**, *40*, 1146–1155.



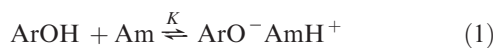
**FIGURE 1.** UV-vis spectra of *p*-nitrophenol in the presence of increasing concentrations of piperidine in [bmim][BF<sub>4</sub>] solution at 298 K.

formed ion pair to get in the thick network of interactions that characterize the solvent media used.

## Results and Discussion

**Method.** A first series of experiments was performed with 2,4-dinitrophenol as proton donor. However, the UV-vis spectrum of this indicator in [bmim][BF<sub>4</sub>] solution revealed that it was actually present in its unprotonated form, which suggested that the acid strength of this indicator is too large. Hence another indicator, having a lower acidity, namely the *p*-nitrophenol, was chosen. In the absence of amine, this molecule, both in IL solution as well as in dioxane, is in its undissociated form ( $\lambda_{\max} = 304$  and 315 nm in dioxane and in IL, respectively). Interestingly, the amines studied were able to induce phenol dissociation only in IL solution, as accounted for by UV-vis spectra (Figure 1).

On increasing base concentration, the absorbance of the unprotonated form of the indicator ( $\lambda_{\max} = 415$  nm) increases. Furthermore, the presence of an isosbestic point at  $\lambda = 344$  nm indicates that in IL solution a single equilibrium, such as the one depicted in eq 1, is operating:



The equilibrium constant value can be defined as:

$$K = \frac{[\text{ArO}^- \text{AmH}^+]}{[\text{ArOH}][\text{Am}]} \quad (2)$$

Considering the mass balance relevant to *p*-nitrophenol and Am, the absorbance variations can be expressed as (eq 3):

$$\Delta A = A_i - A_0 = \frac{S_t \Delta \epsilon K [\text{Am}]}{1 + K [\text{Am}]} \quad (3)$$

where  $A_i$  and  $A_0$  are the absorbances, at a given wavelength of *p*-nitrophenol solutions in the presence and absence of amines, respectively,  $\Delta \epsilon$  is the variation in molar absorptivity as a consequence of the ion pair formation,  $K$  is the equilibrium constant value, and  $S_t$  is the analytical concentration of *p*-nitrophenol. Nevertheless, in agreement with

previous reports and explanations,<sup>3a</sup> the occurrence of a peculiar cation-amine interaction causes a decrease of its effective concentration. Under these circumstances, eq 3 does not correctly fit data trends, which must rather be described by the following modified relationship:

$$\Delta A = A_i - A_0 = \frac{S_t \Delta \epsilon K ([\text{Am}] - a)}{1 + K ([\text{Am}] - a)} \quad (4)$$

where “ $a$ ” is the amine concentration involved in the interaction with the IL cation. A typical fit of experimental data by means of eq 4 is depicted in the inset of Figure 1.

**The Ion Pair Formation in [bm<sub>2</sub>im][NTf<sub>2</sub>].** The behavior just described showed some peculiar features when the ion pair formation was studied in [bm<sub>2</sub>im][NTf<sub>2</sub>] solution. In this case, the UV-vis spectra recorded at a fixed concentration of diisopropylamine, for instance, showed a decrease of the absorbance value at  $\lambda = 424$  nm and an increase at  $\lambda = 315$  nm as a function of time (see Figure 3 of the Supporting Information). To ascertain the causes of this phenomenon, we carried out some investigations, by using different experimental methods (see later), changing both concentration and nature of the amine. In Figure 2 plots of absorbance values as a function of time at different diisopropylamine concentrations (Figure 2a) and in the presence of different amines (Figure 2b) are reported.

Analysis of plots reported in Figure 2 shows that the observed process is affected by both factors considered. In all cases the plots obtained show three different regions: a first one in which absorbance values slightly decrease as a function of time, a second one in which a sudden absorbance variation is recorded, and a third one in which a constant absorbance value is detected. Eventually, the absorbance value stayed constant after 40 min, irrespective of the variable (concentration or amine nature).

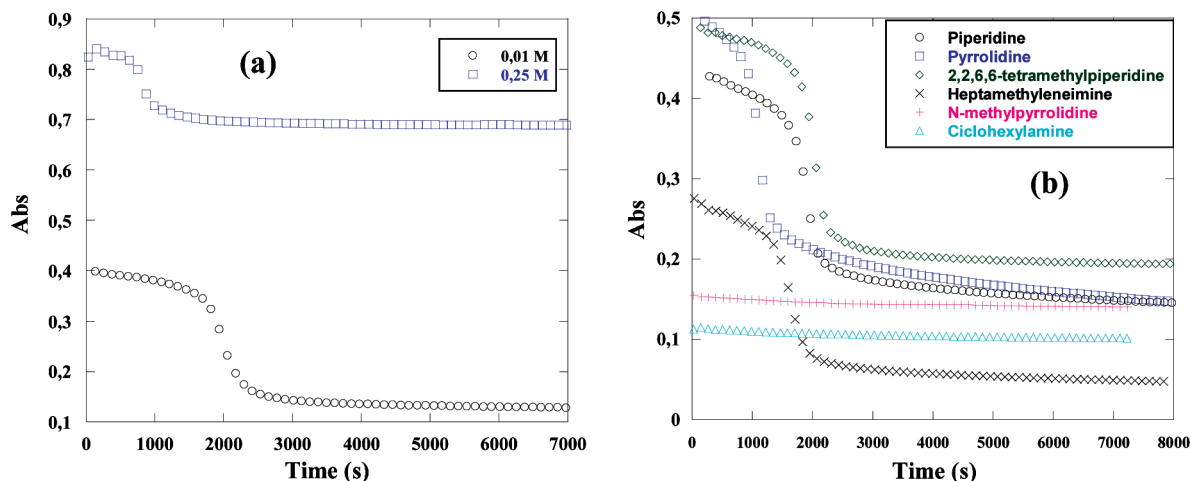
Among the amines used, pyrrolidine is the base able to induce the most significant absorbance variation and the obtained result perfectly recalls the situation previously outlined by us on studying some elimination reactions in IL solution.<sup>3d,g</sup>

To verify if the above process was operating in the other ILs tested in this work, we took piperidine as the model amine (because of the well-defined first region of its plot) and carried out an investigation in the different IL solutions; in order to have comparable absorbance values at  $t = 0$ , different amine concentrations in the various ILs were employed. Obtained data show that, with the exception of [bm<sub>2</sub>im][NTf<sub>2</sub>], in all the other IL solutions the absorbance of *p*-nitrophenate remained constant.

The system [bm<sub>2</sub>im][NTf<sub>2</sub>]/ion pair was also analyzed by using <sup>1</sup>H NMR measurements (<sup>1</sup>H NMR spectra of measurement solutions are reported in Figure 4 of the Supporting Information). Analysis of these spectra shows that the addition of different components of measurement solutions did not induce any change in the chemical shift values of protons corresponding to the 1-butyl-2,3-dimethylimidazolium cation. However, the addition of 75  $\mu\text{L}$  of 1,4-dioxane, as well as the addition of 50  $\mu\text{L}$  of a *p*-nitrophenol solution (0.002 M) in 1,4-dioxane, induced a significant variation in the signals multiplicity. As previously reported,<sup>3d,12</sup> these variations can

(12) D'Anna, F.; Cascino, M.; Lo Meo, P.; Riela, S.; Noto, R. *Arkivoc* 2009, 8, 30–46.





**FIGURE 2.** Plot of absorbance values ( $\lambda = 424$  nm) as a function of the time for (a) *p*-nitrophenol/diisopropylamine/[bm<sub>2</sub>im][NTf<sub>2</sub>] and (b) *p*-nitrophenol/amine/[bm<sub>2</sub>im][NTf<sub>2</sub>].

be explained considering that different ion pairs are present in the mixtures. This hypothesis agrees perfectly with the formation of ion pairs in imidazolium-based ILs.<sup>13</sup>

All detected variations in the UV–vis and <sup>1</sup>H NMR spectra cannot be explained with the occurrence of an IL degradation reaction. Indeed, a *p*-nitrophenol/diisopropylamine/[bm<sub>2</sub>im][NTf<sub>2</sub>] solution was kept at 298 K for 2 h and then extracted with ethyl ether. The analysis of extract did not show any presence of demethylation reaction product of imidazolium cation, such as *p*-nitroanisole. Therefore, bearing in mind data previously reported,<sup>3d,g</sup> we hypothesized that the obtained variations could be ascribed to a reorganization process of IL molecules induced by the ion pair formation. On the other hand, recent literature reports outlined that several experimental methods indicate broadly distributed dynamics of the ions, which is not typical for molecular solvents of similar viscosity. So unusual features of ILs can be attributed to a microheterogeneous environment of the moving particles.<sup>14</sup>

Under this light, in order to have a better understanding of the peculiar features of our system, we thought that a useful support could derive from a resonance light scattering (RLS) investigation. This technique, introduced by Pasternack et al.,<sup>15</sup> allows information about the extent of aggregation of molecules containing chromophores to be obtained. Indeed, the intensity of the RLS band can be correlated to the size of the aggregates.<sup>16</sup> RLS has been largely used to study porphyrins and DNA assemblies,<sup>17</sup> and

more recently to investigate the structural properties of aqueous solutions of ILs<sup>18</sup> (RLS spectra recorded for [bm<sub>2</sub>im][NTf<sub>2</sub>] solutions of *p*-nitrophenol, butylamine, pyrrolidine, *N*-methylpyrrolidine, and corresponding ion pairs are reported in Figure 5 of the Supporting Information). Analysis of RLS data allows us to draw some interesting conclusions. Indeed, in some cases, the intensity of the RLS bands significantly decreases as a function of time, indicating that the thick network of interactions characterizing the [bm<sub>2</sub>im][NTf<sub>2</sub>] is affected by the presence of the ion pair constituents, which induce a decrease in the size of the aggregates. Among used amines, and according to UV–vis and <sup>1</sup>H NMR results, this influence seems to be more significant in the presence of the *p*-nitrophenol/pyrrolidine ion pair. Also in this case, the plot of RLS band intensity, at  $\lambda = 523$  nm, as a function of time showed a constant value after 40 min.

On the whole, collected data seem to confirm the hypothesis that variations in the UV–vis spectra have to be ascribed to slow solvent reorganization, induced by the ion pair. Obviously, this reorganization process seems to be a function of the cation–anion pair considered. In particular, in our case, among aromatic ILs tested, it can be detected only in [bm<sub>2</sub>im][NTf<sub>2</sub>], probably as a consequence of the less intense cation–anion interactions that make easier the movement of IL constituents induced by the ion pair formation. Furthermore, it is noteworthy that this process cannot be ascribed to the presence of water or impurities in the IL sample. As a matter of fact, spectra variations described above were detected in IL samples deriving from different preparations.

On the grounds of this information, we determined *K* values in [bm<sub>2</sub>im][NTf<sub>2</sub>] solution at different times of thermostation. In particular, data collected show that in some cases (diisopropylamine, pyrrolidine, and *N*-methylpyrrolidine), *K* values changed going from 20 to 60 min, but remained constant going from 60 to 80 min. Then, in order to have constant absorbance values in the time, we carried out spectrophotometric determinations in [bm<sub>2</sub>im][NTf<sub>2</sub>]

(13) (a) Avent, A. G.; Chaloner, A. P.; Day, M. P.; Seddon, K. R.; Welton, T. *J. Chem. Soc., Dalton Trans.* **1994**, 3405–3413. (b) Bonhôte, P.; Dias, A.-P.; Papageorgiou, N.; Kalyanasundaram, K.; Grätzel, M. *Inorg. Chem.* **1996**, 35, 1168–1178. (c) Tubbs, J. D.; Hoffmann, M. M. *J. Sol. Chem.* **2004**, 33, 381–394.

(14) Weingärtner, H. *Angew. Chem., Int. Ed.* **2008**, 47, 654–670.

(15) Pasternack, R. F.; Collings, P. J. *Science* **1995**, 269, 935–939.

(16) (a) Anglister, J.; Steinberg, I. Z. *J. Chem. Phys.* **1983**, 78, 5358–5368.

(b) Anglister, J.; Steinberg, I. Z. *J. Chem. Phys. Lett.* **1979**, 65, 50–64.

(17) See for example: (a) Pasternack, R. F.; Schaefer, K. F. *Inorg. Chem.* **1994**, 33, 2062–2065. (b) Liu, Y.; Fan, M.; Zhang, S.; Yao, J. *J. Phys. Org. Chem.* **2007**, 20, 874–877. (c) Cheng, D.-H.; Chen, X.-W.; Wang, J.-H.; Fang, Z.-L. *Chem.—Eur. J.* **2007**, 13, 4833–4839. (d) Wu, J.-J.; Li, N.; Li, K.-A.; Liu, F. *J. Phys. Chem. B* **2008**, 112, 8134–8138. (e) Wang, Y. T.; Jin, W. J. *Spectrochim. Acta, Part A* **2008**, 70, 871–877. (f) Li, J.; Li, M.; Li, X.; Tang, J.; Kang, J.; Zhang, H.; Zhang, Y. *Colloids Surf., B* **2008**, 67, 79–84. (g) Zheng, J.; Wu, X.; Wang, M.; Ran, D.; Xu, W.; Yang, J. *Talanta* **2008**, 74, 526–532.

(18) Zhang, H.; Li, K.; Liang, H.; Wang, J. *Colloids Surf., A* **2008**, 329, 75–81.

TABLE 1. Equilibrium Constant Values ( $K$ ) Relative to ArOH/Am Ion Pair Formation in Different Ionic Liquid Solutions at 298 K

amine	$pK_{BH^+}^a$	$K$ ( $M^{-1}$ )					
		benzene <sup>b</sup>	[bmim][BF <sub>4</sub> ] <sup>c,d</sup>	[bmim][PF <sub>6</sub> ] <sup>c,d</sup>	[bmim][NTf <sub>2</sub> ] <sup>c,d</sup>	[bm <sub>2</sub> im][NTf <sub>2</sub> ] <sup>c,d</sup>	[bmpyrr][NTf <sub>2</sub> ] <sup>c,d</sup>
butyl-	10.68	110	420	120	360	270	90
cyclohexyl-	10.66		510	120	270	280	90
diisopropyl-			1610	320	1230	630	360
dibutyl-	10.20	880	1120	405	940	790	290
<i>N</i> -butyl- <i>N</i> -methyl-			730	210	680	460	170
pyrrolidine	11.27	7630	1500	680	1390	740	315
piperidine	11.12	4490	1100	335	850	270	230
hexamethyleneimine	10.89	3050	1130	350	1030	440	235
heptamethyleneimine	10.78	2000	430	210	430	330	120
2,2,6,6-tetramethylpiperidine	11.07		1880	760	1420	940	430
methylpyrrolidine	10.46		330	300	470	410	180
triethyl-	10.75	2700	670	520	790	750	290

<sup>a</sup>See ref 19. <sup>b</sup>See ref 7. <sup>c</sup> $K$  values are the average between two independent measurements and are reproducible within  $\pm 5\%$ . <sup>d</sup>The highest standard deviation was equal to  $\pm 10\%$ .

solution thermostating measurement solutions at 298 K for 1 h. However, it must be outlined that  $K$  values determined by carrying out “fast” measurements (within 20 min) were only a little different from those obtained after 1 h.

The equilibrium constant values ( $K$ ), relevant to ion pair formation, in the presence of different amines in IL solution are reported in Table 1.

To have a comparison with the basic strength of used amines in molecular solvents,  $pK_{BH^+}$  values, previously collected in water solution, and stability constants of the ion pairs formed in the presence of 2,4-dinitrophenol, previously determined in benzene solution, are also reported in Table 1.

**Basic Strength in Conventional Organic Solvents and in IL Solution.** We are aware of the fact that, as a consequence of the different acid strength of two indicators used in benzene and in IL solution, the comparison between data collected in these solvent systems can be only qualitative. However, the analysis of data reported in Table 1 shows that  $K$  values in IL solutions are comparable with data collected in benzene solution. Bearing in mind the lower acidity of *p*-nitrophenol with respect to 2,4-dinitrophenol ( $pK_a = 7.24^{20a}$  and  $4.10^{20b}$  for *p*-nitrophenol and 2,4-dinitrophenol, respectively), a higher basicity of amines in IL solution could be hypothesized. This result agrees perfectly with data previously obtained by us on studying the amine-induced mononuclear rearrangement of the (*Z*)-phenylhydrazone of the 3-benzoyl-5-phenyl-1,2,4-oxadiazole in IL solution,<sup>3a</sup> but also with data previously reported by Mečiarová et al. about the higher efficiency of piperidine in catalyzing the Michael addition reaction in these media.<sup>6b,c</sup>

As far as basicity ratios are concerned, data reported in Table 1 show that these values are higher in benzene than in IL solution. Indeed, in the first case the highest basicity ratio is equal to 70, while in IL solution, the highest value is equal to 5.7. These results seem to indicate a different and lower ability of ILs, with respect to benzene, to discriminate amines on the grounds of their structure and their steric requirements. This effect could be a consequence of the smaller relevance of solvation effects in IL, which has been frequently claimed in the literature. Furthermore, it could

explain why in all solvent systems (ILs and benzene) the basicity changes along the same series (primary amines < tertiary amines < secondary amines), but with the degree of nitrogen substitution being equal, different variations in ILs and in benzene are detected.

As far as  $K$  values in IL solution are concerned, we tried to correlate stability constant values collected in different ILs. In all cases unsatisfactory correlations were obtained. A similar result has been recently obtained by us, studying in IL solution the protonation equilibrium of sodium *p*-nitrophenate in the presence of carboxylic acids.<sup>21</sup> Once more these data seem to evidence, according to what was previously reported by MacFarlane et al., the peculiarity of the effects that each IL can exert on a given reaction.<sup>22</sup>

As a consequence of this, in order to have a better understanding of results, data obtained by us will be discussed as a function of the solvent and amine nature.

**The Effect of IL Cation Structure.** Taking into account the [NTf<sub>2</sub>]<sup>-</sup>-based ILs, data reported in Table 1 show that for all amines studied in this work,  $K$  values decrease along the series [bmim][NTf<sub>2</sub>] > [bm<sub>2</sub>im][NTf<sub>2</sub>] > [bmpyrr][NTf<sub>2</sub>]. However, this order is completely different from the one determined on the basis of solvent parameter  $\alpha$  values ( $\alpha = 0.617, 0.427, \text{ and } 0.381$  for [bmim][NTf<sub>2</sub>], [bmpyrr][NTf<sub>2</sub>], and [bm<sub>2</sub>im][NTf<sub>2</sub>], respectively).<sup>10c</sup> On the other hand, also a classical solvent polarity parameter such as  $E_T^N$  does not explain the experimental trend ( $E_T^N = 0.544, 0.541, \text{ and } 0.644$  for [bmpyrr][NTf<sub>2</sub>], [bm<sub>2</sub>im][NTf<sub>2</sub>], and [bmim][NTf<sub>2</sub>], respectively).<sup>10c</sup> This result agrees with data previously reported about the carefulness that should be used in rationalizing data collected in IL solutions by means of the concept of liquid polarity.<sup>23</sup> For example, many attempts carried out in these solvent media have shown that their “polarity” is not an absolute parameter but is heavily dependent on the nature of the used probe.<sup>24</sup>

In general, for the same amine,  $K$  values decrease on going from aromatic to aliphatic IL. Bearing in mind the nature of

(19) Lide, D. R. *CRC Handbook of Chemistry and Physics*, 80th ed.; CRC: Boca Raton, FL, 1999–2000.

(20) (a) Judson, C. M.; Kilpatrick, M. *J. Am. Chem. Soc.* **1949**, *71*, 3110–3115. (b) Kortüm, G.; Vogel, W.; Andrussov, K. *Dissotziations Konstanten Organischen Säuren in wässriger Lösung*; Butterworths: London, UK, 1961.

(21) D'Anna, F.; La Marca, S.; Noto, R. *J. Org. Chem.* **2009**, *74*, 1952–1956.

(22) MacFarlane, D. R.; Forsyth, S. A. In *Ionic Liquids as Green Solvents. Progress and Prospects*; ACS Symp. Ser. No. 856; American Chemical Society: Washington, DC, 2003; pp 264–276.

(23) Tiwari, S.; Khypse, N.; Kumar, A. *J. Org. Chem.* **2008**, *73*, 9075–9083.

(24) (a) Armstrong, D. W.; He, L.; Liu, Y.-S. *Anal. Chem.* **1999**, *71*, 3873–3876. (b) Ding, J.; Desikan, V.; Han, X.; Xiao, T. L.; Ding, R.; Jenks, W. S.; Armstrong, D. W. *Org. Lett.* **2005**, *7*, 335–337.

the formed ion pairs, a further stabilization of these as a consequence of the occurrence of  $\pi$ - $\pi$  interactions with aromatic cations could be hypothesized. The importance of  $\pi$ - $\pi$  interactions in affecting the properties of aromatic substrates in ILs solution has been claimed on different occasions.<sup>3,25</sup> On this subject, we have recently ascertained that the aromaticity of the IL cation is able to positively affect the strength of aromatic carboxylic acids.<sup>21</sup> Of course, these interactions should be more effective when the IL shows a higher structural order degree. On this subject, a useful solvent parameter could be the cohesive pressure ( $\delta$ ), defined as a measurement of the strength of the interactions among solvent molecules. In some cases, this parameter has been invoked to rationalize ILs effects on the reaction course.<sup>21,26</sup> More significant interactions should induce a higher structural order degree. Then, in our case, on the grounds of  $\delta$  values ( $\delta = 25.2$  and  $24.2$  MPa for [bmim][NTf<sub>2</sub>] and [bm<sub>2</sub>im][NTf<sub>2</sub>], respectively),<sup>27</sup> [bmim][NTf<sub>2</sub>] gives  $\pi$ - $\pi$  interactions to a larger extension than [bm<sub>2</sub>im][NTf<sub>2</sub>].

**The Effect of IL Anion Structure.** In general, the analysis of data reported in Table 1 for bmim<sup>+</sup>-based ILs shows that  $K$  values decrease along the series [BF<sub>4</sub><sup>-</sup>] > [NTf<sub>2</sub><sup>-</sup>] > [PF<sub>6</sub><sup>-</sup>], according to  $\beta$  values ( $\beta = 0.376$ ,  $0.243$ , and  $0.207$  for [bmim][BF<sub>4</sub>], [bmim][NTf<sub>2</sub>], and [bmim][PF<sub>6</sub>], respectively).<sup>10c</sup>

The anionic part of IL should play an important role. Indeed, it is able to stabilize ammonium ions by coordination of their acid protons as previously reported.<sup>3c,g,28</sup> Similarly, in our case, the anion having the highest H-bond acceptor ability induces a more significant stabilization of the ion pair formed and consequently a more significant  $K$  value increase. Obviously, the extent of this effect decreases on going from primary to tertiary amines.

However, also the viscosity of the used IL seems to affect the stability of the ion pairs. Indeed, the small difference between  $\beta$  values corresponding to [bmim][NTf<sub>2</sub>] and [bmim][PF<sub>6</sub>] does not account for the significant differences detected among corresponding  $K$  values. This result can be rationalized by taking into account also the higher viscosity of [bmim][PF<sub>6</sub>] ( $\eta = 450$  and  $52$  cP for [bmim][PF<sub>6</sub>] and [bmim][NTf<sub>2</sub>], respectively)<sup>29</sup> that induces a decrease in the ion pair stability by slowing down the contact between reagents.

**The Effect of the Amine Structure.** As we mentioned above, irrespective of the IL considered, the basic strength changes along the series primary amines < tertiary amines < secondary amines. We tried also to determine the stability constant value for the morpholine/*p*-nitrophenol ion pair. However, in all cases, also the addition of concentrated solution of amine did not induce any changes in the UV-vis spectra. On the grounds of this result, aromatic amines, such as aniline, were not taken into account.

Among the amines used, the stability of the ion pairs formed in the presence of primary amines seems to be scarcely affected by the structure of the alkyl group.

A different behavior was observed for tertiary amines. Indeed, in this case, independently from the IL considered, the ion pair stability for triethylamine is higher than that for *N*-methylpyrrolidine.

As far as secondary amines are concerned, a distinction among data collected in the presence of cyclic and acyclic amines has to be made. In general, for all the ILs used in this work, the range of  $K$  values increases on going from acyclic to cyclic amines, pointing out a larger dependency of ion pair stability on amine structure when the basic center is involved in a cyclic structure ( $\Delta K = 880$  and  $1450$  M<sup>-1</sup> for acyclic and cyclic amines, respectively, in [bmim][BF<sub>4</sub>] solution).

On the whole, the behavior of secondary acyclic amines seems to indicate a smaller relevance, in IL solution, of steric effect. Indeed, with the only exception of data collected in [bmim][PF<sub>6</sub>] and in [bm<sub>2</sub>im][NTf<sub>2</sub>], diisopropylamine, having the highest steric hindrance, forms more stable ion pairs. Similarly, in all ILs used the ion pair stability decreases on going from dibutylamine to *N*-butyl-*N*-methylamine, according to the decrease of the steric hindrance around the amino nitrogen.

In general, the stability of the ion pairs formed in the presence of cyclic amines decreases on going from pyrrolidine to heptamethyleneimine, in line with the increase in ring size. Furthermore, this decrease is more significant in the presence of aromatic ILs than in the presence of aliphatic IL. For bmim<sup>+</sup>-based ILs, this variation decreases in the presence of anions having the lowest coordination ability. On the whole, these results could be ascribed to a higher flexibility of the amine ring. Indeed, the protonation of the amine nitrogen should induce conformational changes, which could affect the structural order degree of IL and consequently the ion pair stability (see above). The negative effect deriving from a higher conformational freedom of the amine ring agrees with our previous evidence on the elimination reactions carried out in IL solution.<sup>3d</sup>

Along the series of secondary cyclic amines, further interesting comparisons may be made among data collected in the presence of piperidine and in the presence of 2,2,6,6-tetramethylpiperidine. In general, the methylation of  $\alpha$ -carbon atoms to amino nitrogen induces a decrease of amine basicity in water solution that has usually been ascribed to the steric hindrance alkyl groups exert on the solvation process of the corresponding ammonium cation. In IL solution, irrespective of the nature of the constituting ions, the ion pair stability increases on going from piperidine to 2,2,6,6-tetramethylpiperidine. This result, together with the data previously discussed for secondary acyclic amines (see above), once more evidences the smaller relevance of steric effects in this solvent media. This could be a consequence of the different nature or of the smaller relevance of solvation effects in IL solutions than in conventional solvents, and it is consistent with previous reports.<sup>30</sup>

The last factor to be considered is the effect due to the methylation of amino nitrogen on going from pyrrolidine to

(25) (a) Atwood, J. L. *Liquid Clathrates, Inclusion Compounds*; Atwood, J. L., Davies, J. E. D., Mac Nicol, D. D., Eds.; Academic Press: London, UK, 1984; Vol. 1. (b) Hanke, C. G.; Johansson, A.; Harper, J. B.; Lynden-Bell, R. M. *Chem. Phys. Lett.* **2003**, *374*, 85–90. (c) Su, B.-M.; Zhang, S.; Zhang, Z. C. *J. Phys. Chem. B* **2004**, *108*, 19510–19517. (d) Iwata, K.; Kakita, M.; Hamaguchi, H. *J. Phys. Chem. B* **2007**, *111*, 4914–4919.

(26) Angelini, G.; Chiappe, C.; De Maria, P.; Fontana, A.; Gasparrini, F.; Pieraccini, D.; Pierini, M.; Siani, G. *J. Org. Chem.* **2005**, *70*, 8193–8196.

(27) Swiderski, K.; McLean, A.; Gordon, C. M.; Huw Vaughan, D. *Chem. Commun.* **2004**, 2178–2179.

(28) Crowhurst, L.; Lancaster, N. L.; Pérez Arlandis, J. M.; Welton, T. J. *Am. Chem. Soc.* **2004**, *126*, 11549–11555.

(29) (a) Huddleston, J. G.; Visser, A. E.; Reichert, W. M.; Willauer, H. D.; Broker, G. A.; Rogers, R. D. *Green Chem.* **2001**, *3*, 154–164.

(30) (a) Kobrak, M. N. *J. Chem. Phys.* **2006**, *125*, art. no. 064502. (b) Kobrak, M. N. *J. Chem. Phys.* **2007**, *127*, art. no. 184507. (c) Kobrak, M. N. *J. Phys. Chem. B* **2007**, *111*, 4755–4762. (d) Kobrak, M. N. *Green Chem.* **2008**, *10*, 80–86.

*N*-methylpyrrolidine. In all ILs, the ion pair stability decreases on going from the secondary to the tertiary amine and this decrease is more significant in the presence of IL anions having higher coordination ability, according to the stabilizing effect deriving from this IL component (see above). A similar trend was previously observed by us on measuring the rates of the Kemp elimination reaction in [bmim][BF<sub>4</sub>] solution. However, the effect of methylation seems to act more significantly on the thermodynamic process ( $K_{\text{Me-pyrr}}/K_{\text{pyrr}} \approx 4.5$ ) than on the kinetic one ( $k_{\text{Me-pyrr}}/k_{\text{pyrr}} \approx 1.7$ ).<sup>3g</sup>

## Conclusions

On the whole, data collected here show that the amine/*p*-nitrophenol ion pair formation is a suitable probe reaction in order to measure the basic strength in IL solutions. These solvent media are able to exalt the amine basicity, and this effect is a function of both IL nature and amine structure. The trend of experimental data evidence that the ion pair stability depends on its ability to get in the thick network of interactions that characterize these solvent media. On this issue, in some cases ([bm<sub>2</sub>im][NTf<sub>2</sub>]), the ion pair formation affects the structural organization of the solvent and the collected experimental evidences once again underline, according to previous reports by Kobrak<sup>31</sup> and recently reasserted by us,<sup>3g</sup> that the response of ILs to structural changes of solutes is slower than that in molecular solvents, as a consequence of the need to retain favorable interactions with neighboring ions.

## Experimental Section

**Materials.** [bmim][BF<sub>4</sub>], [bmim][PF<sub>6</sub>], 1,4-dioxane, and *p*-nitrophenol were used as purchased without further purification. [bmim][NTf<sub>2</sub>], [bm<sub>2</sub>im][NTf<sub>2</sub>], and [bmpyrr][NTf<sub>2</sub>] were prepared

according to literature procedures.<sup>32</sup> ILs were dried in vacuum at 60 °C for 2 h before use, then they were stored in a desiccator over calcium chloride. Amines were freshly distilled before use.

**Measurements and Calculations: UV–Vis Spectra.** The UV–vis spectra were recorded by using a spectrophotometer equipped with a temperature controller.

**Spectrophotometric Measurements.** Samples for a typical spectrophotometric measurement were prepared by injecting, by means of microsyringe, the appropriate volume of IL and 1,4-dioxane into a quartz cuvette (light path 0.2 cm). Amine and *p*-nitrophenol were added as concentrated solutions in 1,4-dioxane. The *p*-nitrophenol concentration was kept constant (0.0002 M). Each sample was thermostated at 298 K. The suitable wavelength was chosen by comparing the *p*-nitrophenol solution with the sample having the higher amine concentration. Experimental data were subjected to fitting regression analysis according to eq 4.

**<sup>1</sup>H NMR Measurements.** NMR spectra were collected on a 300 MHz spectrometer. In NMR measurements the opportune volumes of IL, amine, and/or *p*-nitrophenol solution were mixed in a 5 mm NMR tube. A steam coaxial capillary tube loaded with DMSO-*d*<sub>6</sub> was used for the external lock of the NMR magnetic field/frequency and its signal was used as the <sup>1</sup>H NMR external reference at 2.56 ppm.

**RLS Measurements.** Resonance light scattering measurements were carried out on a spectrofluorimeter employing a synchronous scanning mode in which the emission and excitation monochromators were preset to identical wavelengths. The resonance light scattering spectrum was recorded from 300 to 600 nm with both excitation and emission slit widths set at 1.5 nm.

**Acknowledgment.** We thank the University of Palermo for financial support (funds for selected research topics).

**Supporting Information Available:** UV–vis, <sup>1</sup>H NMR, and resonance light scattering (RLS) spectra. This material is available free of charge via the Internet at <http://pubs.acs.org>.

(31) Kobrak, M. N. *J. Chem. Phys.* **2006**, *125*, 064502–1–11.

(32) Schröder, U.; Wadhawan, J. D.; Compton, R. G.; Marken, F.; Suarez, P. A.; Consorti, C. S.; de Souza, R. F.; Dupont, J. *New J. Chem.* **2000**, *24*, 1009–1015. (b) Cammarata, L.; Kazarian, S. G.; Salter, P. A.; Welton, T. *Phys. Chem. Chem. Phys.* **2001**, *3*, 5192–5200.