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Effect of the Presence of MEA on the CO$_2$ Capture Ability of Superbase Ionic Liquids

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Supporting Information

ABSTRACT: The miscibility of monoethanolamine (MEA) in five superbase ionic liquids (ILs), namely the trihexyl-tetradecyolphosphonium benzo triazolide ([P$_{66614}$][Bentriz]), trihexyl-tetradecyolphosphonium benzimidazolide ([P$_{66614}$][Benzim]), trihexyl-tetradecyolphosphonium 1,2,3-triazolide ([P$_{66614}$][123Triz]), trihexyl-tetradecyolphosphonium 1,2,4-triazolide ([P$_{66614}$][124Triz]), and trihexyl-tetradecyolphosphonium imidazolide ([P$_{66614}$][Im]) was determined at 295.15 K using $^1$H NMR spectroscopy. The solubility of carbon dioxide (CO$_2$) in equimolar (IL + MEA) mixtures was then studied experimentally using a gravimetric technique at 295.15 K and 0.1 MPa. The effect of MEA on the CO$_2$ capture ability of these ILs was investigated together with the viscosity of these systems in the presence or absence of CO$_2$ to evaluate their practical application in CO$_2$ capture processes. The effect of the presence of MEA on the rate of CO$_2$ uptake was also studied. The study showed that the MEA can enhance CO$_2$ absorption over the ideal values in the case of [P$_{66614}$][123Triz] and [P$_{66614}$][Bentriz] while in the other systems the mixtures behave ideally. A comparison of the effect of MEA addition with the addition of water to these superbase ILs showed that similar trends were observed in each case for the individual ILs studied.

1. INTRODUCTION

Carbon dioxide is a greenhouse gas that is released into the atmosphere from flue gas streams. It is believed that CO$_2$ currently contributes to 60% of the greenhouse gases released into the atmosphere. With global CO$_2$ emissions set to increase there is a real and immediate need to develop a process that can effectively and efficiently capture CO$_2$ before being released into the atmosphere. The sources of CO$_2$ include mobile and stationary sources with the latter associated predominantly from power plants in the form of flue gas streams. As well as emissions of CO$_2$, it is also present in significant concentrations in biogas and natural gas. In these systems, the CO$_2$ limits the combustion of the methane and must be separated before the gas can be utilized as a fuel. In both emissions and methane containing streams, efficient separation is critical for both economic and environmental reasons. To enable these types of separations, industrial practice is currently dominated by aqueous alkanolamine solvents (e.g., 30 wt % aqueous monoethanolamine (MEA)) where a chemical reaction occurs between the CO$_2$ and amine solution. This process is attractive due to the low cost of solvent, high reactivity, and an absorption capacity of 1:2 mol of CO$_2$ per mole of solvent ratio. These advantages are offset by the volatile and corrosive nature of the aqueous solvent mixtures and the high energy consumption needed for regeneration with an enthalpy of regeneration of CO$_2$, $\approx$ 66.7–76.9 kJ·mol$^{-1}$, leading to high operational costs and environmental impact. Due to these disadvantages there has been a recent increase in interest in the search for a more suitable solvent for postcombustion CO$_2$ capture.

In this respect ionic liquids (ILs) have received an increasing amount of interest in recent years as an alternative solvent for CO$_2$ capture due to their potential as environmentally friendly or “green” solvents. ILs have highly attractive properties such as high thermal stability, negligible vapor pressure, and their ability to be task specific. Blanchard et al. were the first to report high CO$_2$ solubility in imidazolium-based ILs reaching a CO$_2$ mole fraction of 0.6 at 8 MPa and 298.15 K. It has also been shown that imidazolium ILs containing the bis-[(trifluoromethyl)sulfonyl]imide ([NTf$_2$]) anion have a high affinity for CO$_2$ capture, in comparison with hexafluorophosphate-based ILs, for example. However, many of these ILs only absorb the CO$_2$ physically and, unfortunately, their CO$_2$ capacity is too low to be industrially viable. However, by adding basic character to the IL it was observed that the CO$_2$ capacity could be increased. Bates et al. synthesized the first IL with an amine moiety, which resulted in an increase of the CO$_2$ capacity driven by chemical absorption of the CO$_2$. However, these ILs were still limited to a 1:2 mol of CO$_2$ per mole of IL ratio, denoted nCO$_2$/nIL. It was then reported that by adding an amine moiety to both the anion and the cation equimolar CO$_2$ capture could be reached. Gurkan et al. showed that, by selecting an amino acid anion-based IL such as the trihexyl-tetradecyolphosphonium proline, [P$_{66614}$][Pro], a 1:1 stoichiometry could be reached at room temperature and atmospheric pressure. Although these ILs showed high CO$_2$ capacity, their...
high viscosity after CO2 solvation made them unsuitable for a practical industrial application. This change in viscosity is mainly due to the formation of strong hydrogen-bonded networks resulting in an increase of the overall cohesive energy of the solution in the presence of CO2.14 Luo et al.15 showed that a high absorption capacity up to 1.6 mol of CO2 per mole of IL could be reached by using pyridine containing anion functionalized ILs. The highest value reported (1.6 mol of CO2 per mole IL) was found using a 4-azabenzimidazole-based IL [P66614]·[4-ABI] at 20 °C and 1 atm. The group also reported these ILs to be recyclable with no loss in absorption after 6 cycles using a desorption temperature of 80 °C.

Recently, a range of superbase ILs have been reported which showed efficient and reversible CO2 capture with an uptake greater than the 1:1 nCO2/nIL.16−18 Unlike the amino acid-based ILs, these ILs have a limited number of hydrogen atoms on the anion and, therefore, only a small increase in viscosity upon CO2 absorption is observed. In addition, as flue gas/methane streams contain water it is also important to understand how these ILs behave under wet conditions. The presence of water has been shown to affect the physical properties of the IL and also has the potential to compete with the CO2 for absorption.19,20 Using the trihexyl-tetradecylphosphonium cation, [P66614]+, paired with a range of superbasic anions, including 1,2,3-triazole ([123Triz]), 1,2,4-triazole ([124Triz])−, phenoxide ([PhO]−), benzo-triazolide ([Bentriz])−, and benzimidazolide ([Benizim])−, work within our laboratory has shown that the presence of water can have a positive or negative effect on the CO2 uptake depending on the anion structure.18 Where the effect was detrimental the change was small, for example in the case of [P66614][124Triz] a decrease from 0.92 nCO2/nIL under dry conditions to 0.89 nCO2/nIL under wet conditions was observed at 295.15 K and atmospheric pressure. However, it was also shown that the presence of water could enhance the CO2 absorption in some superbase ILs, with an increase from 0.26 to 0.63 or from 0.35 to 0.86 nCO2/nIL observed in the case of the [Bentriz]− or [123Triz]− based IL under dry and wet conditions at 295.15 K and atmospheric pressure, respectively. Furthermore, as expected, under wet conditions, the viscosity of these ILs before and after CO2 absorption is much lower than that under dry conditions, showing that these ILs are more promising solvents for industrial application. NMR studies carried out showed that the CO2 competes effectively with water for CO2 absorption.18,20 The studies also showed that the [124Triz]− based IL retained 97% of its CO2 capacity under wet conditions over six cycles showing that CO2 absorption is recyclable and suggesting that the water does not have an irreversible effect on the CO2 uptake.18

In recent years, solutions of alkanolamine and IL mixtures have also been investigated for their CO2 capture capabilities. Reports have shown these mixtures could help overcome drawbacks associated with the current technology.21−28 Camper et al.21 studied mixtures of pure MEA and [C3mim][NTf2]. Therein, the CO2 solubility was found to approach 0.5 mol of CO2 per mole of amine and suggests that these mixtures may reduce the energy of desorption in comparison with the aqueous alkanolamine solution. Other studies have reported that a mixture of MEA and the 1-butyl-3-methylimidazolium tetrafluoroborate, [C4mmim][BF4] could give a higher capacity reaching 0.638 mol CO2 per mole of MEA and better regeneration than that of a pure amine system.22 Another study reported by Yang et al.23 showed the advantage of adding an IL, such as the [C4mim][BF4] to an aqueous amine solution in order to enhance its CO2 solubility. During this study, these authors showed that the mixture based on 30 wt % MEA + 40 wt % [C4mim][BF4] + 30 wt % H2O had a 37.2% lower energy consumption than that of a standard aqueous amine solution. Zhang et al. also found that the addition of the methylthiadianolamine (MDEA) could increase the rate and the CO2 capacity in tetramethylammonium glycinate, [N11111][Gly].24−26 During this work, this group showed that the aqueous mixture based on 15 wt % [N11111][Gly] + 15 wt % MDEA reaching a capacity of 0.64 nCO2/namine in less than 100 min. Zhao et al.27 investigated the performance of CO2 in 16 different absorbents including (amine + IL + H2O), (IL + H2O), and (amine + H2O) systems as a function of the composition and temperature (from 303.15 to 343.15 K) and at 1.50 MPa. These solutions also present a viscosity lower than 17.00 mPa·s at room temperature, which is much lower than that of the pure ILs studied, which is an advantage for a practical application. Therein, it was also reported that the (monodiethanolamine + 2-[hydroxyethyl(methyl)amino]-ethanol chloride + water + piperase) solution had a large CO2 capacity close to 0.158 in CO2 weight fraction. A more recent study on the aqueous 1-hydroxyethyl-3-methylimidazolium glycinate, [C2OHmim][Gly], and MEA mixtures showed an optimum composition of 0.7 mol·dm−3 MEA and 0.3 mol·dm−3 of IL with a capacity of 0.534 nCO2/nabsorbent and suggested that the components mutually promote the CO2 solubility. Furthermore, this particular system showed good stability and CO2 uptake and release as a capacity of 0.480 nCO2/nabsorbent was still observed after the fourth cycle.28 These recent studies have shown that these mixtures can reduce the viscosity of the ILs and decrease absorption time with maximum absorption being reached in 60 min in some cases.26,27,29 It was also shown that they can reduce the loss of material and energy consumption associated with current amine technology.30−32 This solution presents itself as an attractive method to utilize the performance of amines while using the desirable properties of ILs.

In this study MEA was added to the previously studied superbase ILs to investigate the effect of a chemically absorbing cosolvent compared with our previous work with water as a nonreacting cosolvent.33 As noted, water can have a positive influence on the CO2 solubility in superbase ILs which is thought to be partly due to the formation of OH− on reaction with the superbase. Herein, the CO2 solubility in five superbase IL/MEA systems is reported at 295.15 K and atmospheric pressure to examine the influence of a molecule containing both an −OH group as well as the reactive −NH2 in order to explore whether a synergistic effect could be obtained. Few studies have examined similar reactive systems and this also has the potential to reduce the volatility of the MEA which is a major problem in the current aqueous MEA solutions utilized industrially. Furthermore, no CO2 solubility data in binary mixtures containing a superbase IL and the pure MEA has been reported, to date. Five superbase ILs, namely the trihexyl-tetradecylphosphonium ([P66614]+) cation paired with anions [123Triz]−, [124Triz]−, [Bentriz]−, [Benizim]−, and [Im]− were selected due to their high CO2 capture capacity even under wet conditions.18 Their CO2 capacity, viscosity data before and after CO2 uptake as a function of the MEA concentration as well as the recyclability of the system has been reported and compared with an analogous nonreactive ionic liquid based on the bis[(trifluoromethyl)sulfonyl]imide.
systems were described in a previous work. The measurement of the CO2 uptake in the IL/MEA mixtures followed the same gravimetric procedure. The sample (~0.5 g) was pretreated by bubbling with dry nitrogen at 353 K for 1 h (50 cm3·min−1) at 313.15 K and the weight of the sample was recorded on an electronic balance with an accuracy of ±0.1 mg. In addition, the loss of MEA during this pretreatment was determined via 1H NMR spectroscopy. The sample was then bubbled with CO2 (50 cm3·min−1) at room temperature (295.15 ± 0.5 K) for the required period of time, and the weight gain was used to calculate the amount of CO2 absorbed considering the mole fractions of each component (taking into account the mass of the headspace). The CO2 absorption was monitored at intervals of 5, 15, 30, and 60 min. For desorption of CO2, the sample was bubbled with N2 (50 cm3·min−1) at 353.15 ± 0.5 K for 120 min. A desorption experiment was also carried out at 393.15 ± 0.5 K for 60 min. The MEA loss during desorption was quantified by 1H NMR and the additional MEA needed to restore an equimolar system was calculated and added prior to the CO2 being reabsorbed, as described above.

2.4. Equilibrium Constant for IL/MEA System. The equilibrium constant for investigated [P66614][Bentriz]/MEA mixtures at varying molar ratios were calculated based on the following four reactions:

- **IL + MEA kIL/MEA complex**
- **IL + CO2 kIL/CO2**
- **MEA + CO2 kMEA/CO2**
- **IL/MEA complex + CO2 kIL/MEA/CO2**

where ki is equilibrium constant of the reaction i (with i = 1, 2, 3, or 4).

To solve this problem, several assumptions have been made, such as (i) the formation of the IL/MEA complex is driven by an 1:1 molar ratio (eq 1); (ii) the solubility of the CO2 in free IL was calculated on the basis of a 1.023 mol of IL per mole of CO2 ratio based on experimental data for the pure IL system (eq 2); (iii) the solubility of the CO2 in free MEA was determined on the basis of a 1.042 mol of MEA per mole of CO2 ratio based on experimental data for the pure MEA system; (iv) the solubility of the CO2 in the IL/MEA complex was assumed to follow a 1:1 mol of IL/MEA complex per mole of CO2 ratio.

Equation 5 was then used to calculate the equilibrium constant k1 (eq 1).

$$K = \frac{x}{(1-x)^2}$$

where k1 is the equilibrium constant of the eq 1 and x defines the number of moles of IL/MEA complex formed on the basis of an 1:1 molar ratio.

This equilibrium constant was then calculated using solver within Microsoft Excel using the overall mass balances for the CO2, IL, and MEA as boundary conditions and a constant value of k1 as the convergence criterion across all data. The accuracy was tested against experimental values of CO2 uptake in [P66614][Bentriz]/MEA molar ratios of 1:3, 1:1, and 3:1 mol of IL per mole of MEA resulting in a RAAD of 13% in mole fraction units which is within the error of the experimental data.
3. RESULTS AND DISCUSSION

3.1. MEA Miscibility in Selected Superbase ILs. Prior to the determination of the CO₂ solubility in the selected (IL + MEA) mixtures, the miscibility of MEA in each IL was determined at 295.15 K by using ¹H NMR spectroscopy. As shown in Table 1, the MEA is only partially miscible in each selected IL. The MEA solubility in the five ILs, expressed in MEA weight fraction unit, ranged from 0.26 to 0.41. MEA was the most soluble in [P66614][124Triz] with a wMEA = 0.41 and it then decreased in the following order: [P66614][124Triz] > [P66614][Bentriz] > [P66614][Im] > [P66614][Benzim] > [P66614][NTf2] with a wMEA = 0.20. The high solubility reflects a strong interaction between MEA and the anion through N···H···O hydrogen bonds. From data reported in Table 1, it is clear that the solubility of MEA is higher in triazolide-based ILs than the imidazolide-based ILs which may be associated with the potential for the −OH, for example, to hydrogen bond in a bidentate structure across two adjacent nitrogen(s) and, as a result, becomes more stable in the former ILs. The decrease in the MEA solubility for those anions with an additional aromatic system, that is, [P66614][Bentriz] and [P66614][Benzim], suggests that the anion–anion interactions through, for example, increased π···π stacking significantly affect the MEA solvation process in these superbase ILs. As reported in Table 2, the water solubility in the selected ILs decreases in the following order: [P66614][Im] = [P66614][123Triz] ≥ [P66614][124Triz] > [P66614][Bentriz] > [P66614][Benzim]. The water solubility in these ILs was found to range typically from 0.07 to 0.12 in water weight fraction units at 295.15 K (Table 2) showing that the MEA is more soluble in these superbase ILs than the water. Investigations were also performed on the three-component system (IL + MEA + water). In this case, even at low concentrations of water, a biphasic system was obtained with the aqueous phase containing more than 90% of the MEA. This was demonstrated by adding 0.6 g of H₂O into 0.5 g of a mixture containing an equimolar amount of [P66614][Bentriz] and MEA at 295.15 K. In this case, the water layer contained 0.00074 mol (i.e., 91.4%) of MEA while 0.00007 mol (i.e., 8.6%) of MEA was present in the IL layer. Due to this effect, only the binary mixtures were examined for CO₂ capture, hereafter.

3.2. CO₂ Solubility. The CO₂ solubility in the five superbase IL and MEA absorbents was determined using a gravimetric method at atmospheric pressure and room temperature (e.g., 295.15 ± 0.5 K). Initially the volatility of the MEA in the mixture was determined via ¹H NMR spectroscopy, that is, by assessing the quantity of MEA lost during the initial pretreatment of the mixture. This allows for a more accurate representation of the CO₂ capture ability. The mol % loss is shown in Table S2 of the Supporting Information.

The effect of the presence of MEA on the CO₂ solubility in the selected superbase ILs is shown in Table 2 and Figure 1.

### Table 1. MEA Solubility Expressed in MEA Weight Fraction, wMEA, in the Selected ILs at 295.15 K and Atmospheric Pressure along with the Water Content (wt%) in Each Equilibrium Mixture

<table>
<thead>
<tr>
<th>Ionic Liquid</th>
<th>Water Content (wt%)</th>
<th>MEA (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>[P66614][124Triz]</td>
<td>&lt; 0.8</td>
<td>0.379</td>
</tr>
<tr>
<td>[P66614][123Triz]</td>
<td>&lt; 0.6</td>
<td>0.409</td>
</tr>
<tr>
<td>[P66614][Bentriz]</td>
<td>&lt; 0.2</td>
<td>0.325</td>
</tr>
<tr>
<td>[P66614][Benzim]</td>
<td>&lt; 0.8</td>
<td>0.261</td>
</tr>
<tr>
<td>[P66614][Im]</td>
<td>&lt; 0.9</td>
<td>0.316</td>
</tr>
<tr>
<td>[P66614][NTf₂]</td>
<td>&lt; 0.04</td>
<td>0.203</td>
</tr>
</tbody>
</table>

\[ w(\text{water content}) = (0.01 \times \text{water content}) \text{ wt}\%; \quad u(w_{\text{MEA}}) = 0.001; \quad u(T) = 0.5 \text{ K}. \]

### Table 2. Comparison of the CO₂ Solubility in Pure ILs (Dry Conditions), Water Saturated ILs (Wet Conditions) and of the Investigated Equimolar (IL + MEA) Mixtures

<table>
<thead>
<tr>
<th>Ionic Liquid</th>
<th>CO₂ Absorption after 1 h—corresponding to the equilibrium</th>
<th>CO₂ Absorption under Wet Conditions</th>
<th>CO₂ Absorption in Equimolar (IL + MEA) Mixture</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Pure IL</td>
<td>Absorption under Wet Conditions</td>
<td>Absorption in Equimolar (IL + MEA) Mixture</td>
</tr>
<tr>
<td></td>
<td>Water Content before CO₂ absorption (wt%)</td>
<td>nCO₂/nIL</td>
<td>Water Content after Saturation with Water and CO₂ (wt%)</td>
</tr>
<tr>
<td>[P66614][124Triz]</td>
<td>0.046</td>
<td>0.92</td>
<td>11.27</td>
</tr>
<tr>
<td>[P66614][123Triz]</td>
<td>0.078</td>
<td>0.35</td>
<td>11.57</td>
</tr>
<tr>
<td>[P66614][Bentriz]</td>
<td>0.042</td>
<td>0.23</td>
<td>7.42</td>
</tr>
<tr>
<td>[P66614][Benzim]</td>
<td>0.075</td>
<td>1.06</td>
<td>8.42</td>
</tr>
<tr>
<td>[P66614][Im]</td>
<td>0.019</td>
<td>0.83</td>
<td>11.59</td>
</tr>
</tbody>
</table>

Herein, reported CO₂ absorption data, carried out at 295.15 K and 0.1 MPa, are considered as equilibrium properties. Where nCO₂/nAbsorbent data represents the moles of CO₂ absorbed per the total number of moles of absorbents: i.e., including superbase IL, water and/or MEA, while nCO₂/nIL data consider only the number of mole of IL in each solution. u(water content) = (0.01 × water content) wt%; u(nCO₂/nAbsorbent) = 0.01; u(T) = 0.5 K; u(P) = (0.01 × P) MPa.
These results show that, for a 1:1 mol ratio of IL/MEA, [P66614][124Triz] has the highest capacity with a value of 0.68 nCO2/nabsorbent at 295.15 K and 0.1 MPa. Absorption capacities then decrease for the mixtures in the following order: [P66614][124Triz] > [P66614][Benzim] = [P66614][Im] > [P66614][123Triz] > [P66614][Bentriz]. Comparing these CO2 capacities with those of the pure ILs (see Table 2) it is observed that the presence of MEA in solution can have a detrimental effect on the CO2 capture ability of these superbase ILs. For example, the CO2 capacity in [P66614][124Triz] decreased from 0.92 nCO2/nIL to 0.68 nCO2/nabsorbent in the presence of MEA (1:1 mol ratio). A similar trend was also followed for the [P66614][Benzim] and [P66614][Im]. Interestingly, the results show that, in the case of the [P66614][123Triz] and [P66614][Bentriz], the presence of MEA in solution enhanced the CO2 capacity. For [P66614][123Triz], the capacity increased from 0.35 nCO2/nIL for the pure IL to 0.58 nCO2/nabsorbent for the equimolar (IL + MEA) mixture. Similarly, for [P66614][Bentriz], its CO2 capacity increased from 0.23 nCO2/nIL to 0.56 nCO2/nabsorbent by adding an equimolar amount of MEA in solution. These values may be compared with previously reported studies using [C6OHmim][Gly] in an aqueous blend with MEA for CO2 capture. The highest value reported for this system by Lv et al. was 0.534 using a mixture of 0.7 mol-dm−3 MEA and 0.3 mol-dm−3 IL.28 Increased values for CO2 capture are found in this study in comparison.

The trends observed on the addition of MEA to the superbase ILs is similar to the previously reported results when these superbase ILs were tested for their CO2 capacity after saturation with water at 295.15 K (Table 2).38 Again, both [P66614][123Triz] and [P66614][Bentriz] exhibited an enhancement in the presence of water whereas a decrease was obtained for [P66614][124Triz], [P66614][Benzim] and [P66614][Im] ILs. The fact that for both a reactive and a nonreactive additive similar trends are observed may indicate that it is not solely due to the availability of the anion for CO2 interaction. Upon comparison it would appear that these solvents have a similar effect on the ILs tested whether this is an enhanced or undesirable effect on the CO2 uptake of the pure superbase ILs. As water acts as a nonreactive solvent and MEA is a solvent that chemically binds with the CO2 it is desirable for CO2 absorption as it initially could appear with the [P66614][123Triz] giving 0.19 nCO2/nabsorbent after saturation with water compared to 0.58 nCO2/nabsorbent in an equimolar ratio with MEA.

Table 3 shows the CO2 absorption in terms of the mass of CO2 absorbed in the pure IL, the water saturated ILs, and 1:1 mol ratio IL/MEA systems displayed as wt %. To evaluate the potential industrial application of these systems, it is useful to consider the CO2 uptake in terms of mass of absorbent: higher wt % CO2 will result in a lower overall cost of the system. As can be seen, the IL/MEA mixtures provide the highest absorption in terms of mass, with the highest found for the [P66614][124Triz] IL at 9.7 wt % of CO2. In general, water saturation also led to an increase in the wt % CO2 absorption capacity with ~7 wt % found for [P66614][124Triz], [P66614][Benzim], and [P66614][Im]; however, only ~2 wt % was observed for the analogous wet [123Triz]− and [Bentriz]− based ILs. For the pure ILs the CO2 capacity was in the range 4–6 wt %. Figure 2 shows the comparison of the pure IL and equimolar IL/MEA systems with the industrial benchmark for CO2 capture process. It shows that each of these systems gives a better CO2 capacity than the aqueous amine system currently implemented.

3.3. Viscosity. Previous studies27–29 have shown that the addition of MEA to IL reduces the viscosity issues associated with the pure ILs for CO2 uptake, for example. The viscosities of equimolar (superbase IL + MEA) mixtures were measured at 298.15 K before and after CO2 solvation as reported in Table 4.

<table>
<thead>
<tr>
<th>Table 3. CO2 Absorption in Superbase IL-Based Solutions Reported in wt %a</th>
</tr>
</thead>
<tbody>
<tr>
<td>ionic liquid</td>
</tr>
<tr>
<td>---------------</td>
</tr>
<tr>
<td>[P66614][124Triz]</td>
</tr>
<tr>
<td>[P66614][123Triz]</td>
</tr>
<tr>
<td>[P66614][Bentriz]</td>
</tr>
<tr>
<td>[P66614][Benzim]</td>
</tr>
<tr>
<td>[P66614][Im]</td>
</tr>
</tbody>
</table>

“Herein, reported CO2 absorption data, carried out at 295.15 K and 0.1 MPa, are considered as equilibrium properties in the case of the pure ILs (dry conditions), water saturated ILs (wet conditions) and of the investigated equimolar (IL + MEA) mixtures. u( wt % CO2) = 0.05; u(T) = 0.5 K; u( P) = (0.01 × P) MPa.

Figure 2. Comparison of CO2 absorption in pure ILs (■) and in equimolar (IL + MEA) mixtures (□) with the current industrial benchmark for CO2 capture (2:1 nCO2/amine) represented by the dashed line.

Table 4. Physical Properties of Investigated Equimolar (IL + MEA) Mixtures at 298.15 K and Atmospheric Pressure Including Water Content and Viscosity (η) before and after CO2 Absorptiona

<table>
<thead>
<tr>
<th>ionic liquid</th>
<th>water content (wt %)</th>
<th>viscosity (mPa s)</th>
<th>before CO2</th>
<th>after CO2</th>
</tr>
</thead>
<tbody>
<tr>
<td>[P66614][124Triz]</td>
<td>0.78</td>
<td>169</td>
<td>1135</td>
<td></td>
</tr>
<tr>
<td>[P66614][123Triz]</td>
<td>0.54</td>
<td>154</td>
<td>806</td>
<td></td>
</tr>
<tr>
<td>[P66614][Bentriz]</td>
<td>0.14</td>
<td>197</td>
<td>1174</td>
<td></td>
</tr>
<tr>
<td>[P66614][Benzim]</td>
<td>0.84</td>
<td>336</td>
<td>1286</td>
<td></td>
</tr>
<tr>
<td>[P66614][Im]</td>
<td>1.10</td>
<td>168</td>
<td>863</td>
<td></td>
</tr>
<tr>
<td>[P66614][NTf2]</td>
<td>0.04</td>
<td>90</td>
<td>229</td>
<td></td>
</tr>
</tbody>
</table>

“u( water content) = (0.01 × water content) wt %; u(η) = (0.01 × η) mPa-s; u(T) = 0.01 K.”
The results confirm that, prior to the treatment with CO2, the viscosity of all the ILs investigated are lower when MEA is present in solution compared with the pure IL, as expected. For example, the viscosity of the pure [P66614][124Triz] decreased from 569 mPa·s to 169 mPa·s by adding an equimolar amount of MEA. The viscosities are found to decrease in the order [P66614][Bentriz] > [P66614][124Triz] > [P66614][123Triz] > [P66614][Bentriz] > [P66614][NTf2] for the pure ILs (Table S3) and [P66614][Bentriz] > [P66614][124Triz] > [P66614][123Triz] > [P66614][124Triz] ≥ [P66614][Im] > [P66614][123Triz] > [P66614][NTf2] for the 1:1 IL/MEA mixtures. Interestingly, unlike the results found for the pure ILs, upon absorption of CO2 into the IL/MEA mixtures a large increase in viscosity is observed. The largest increase was observed for the equimolar [P66614][124Triz]/MEA system with a viscosity change from 167 mPa·s before CO2 absorption to 1315 mPa·s after CO2 absorption. This is unlike the pure IL in which CO2 absorption resulted in only a small increase in viscosity. Similar increases in viscosity upon CO2 sorption were also reported by Fu et al.31 wherein, for a mixture of 0.35 wt% MEA/0.15 wt% [Bmim][Gly] in water, the viscosity increased from 6.59 mPa·s to 10.35 mPa·s upon the addition of CO2. The addition of MEA had a similar effect on the viscosity to the addition of water as was previously reported (Table S3), that is, a decrease was observed in both cases. However, a much greater change in viscosity was observed upon CO2 absorption in the equimolar (IL + MEA) systems compared with the wet IL systems. In the latter, only a small increase in viscosity from 114 to 200 mPa·s was found on absorbing CO2 into the water saturated [P66614][124Triz], for example. This difference in viscosity between the MEA and water cosolvents after CO2 absorption is expected due to the reactive nature of the MEA in the presence of CO2 leading to an increase of ions in solution. This induces an increase of the Coulombic interactions in solution affecting the cohesive energy of the solution.

This increase in viscosity on CO2 absorption for the equimolar (IL + MEA) systems is important as it has been reported widely that the presence of an amine in ILs leads to enhanced kinetics of absorption over the pure IL, in the case of nonreactive ILs.1 Viscosity data for (reactive IL + MEA) systems have not been generally observed in the case of the selected superbase ILs in the absence and presence of water and this should be examined for other potential IL systems. Figure 3 compares the absorption kinetics for the pure IL, (water + IL) at the water saturation level and equimolar (IL + MEA) systems for the case of [P66614][124Triz] as represented as % of the saturation number of moles of CO2 absorbed with respect to time (Table S4). No significant differences were observed for the three systems studied suggesting that the rate of absorption of CO2 is not just dependent on the viscosity. This is further demonstrated in Figure 4 as no correlation was found by fitting the viscosity of the equimolar (IL + MEA) mixtures before the CO2 uptake as the function of their initial rate of CO2 uptake. This indicates that mass transfer of CO2 is not rate determining for the uptake and, therefore, it is likely due to the structure of the system and varying strengths of interaction between the ion pairs also playing a significant role.

3.4. Synergy between the IL and MEA Components.

The five equimolar (IL + MEA) mixtures were probed to investigate their synergistic properties. That is, are the components of the absorbent working together to increase the CO2 capture ability or are the two components working independently to absorb CO2. In Table 2, the calculated ideal results are displayed, using a capacity of 0.42 nCO2/nabsorbent for the MEA component and the experimental values obtained previously for the pure ILs.18 A comparison of the ideal and experimental results for the [124Triz]− based IL with MEA in an equimolar ratio shows an experimental value of 0.68 nCO2/nabsorbent versus a theoretical value of 0.69 nCO2/nabsorbent, suggesting that this acts as an ideal system, that is, the CO2 capacity is a weighted average of the two individual capacities for the IL and MEA. This is also found using the [P66614][Bentriz] and [P66614][123Triz] ILs.

In comparison, for the [P66614][123Triz] and [P66614][Bentriz] ILs an increase in the overall absorption capacity is observed compared with the ideal case, Figure 5 shows the CO2 uptake as a function of IL/MEA mole ratio for the [P66614][Bentriz]. The dashed line is representative of the ideal CO2 uptake and demonstrates that at all compositions of IL/MEA the expected uptake is achieved. MEA and MEA behave in a synergistic manner with enhanced CO2 solubility compared...
with the ideal case. Figure 5 also shows that the highest synergy is achieved at the 0.5:0.5 IL/MEA ratio. Using these solubility data and the mathematical model described in section 2.4, the equilibrium constant $k_1$ was then calculated. A very high value for the equilibrium constant $k_1$, close to 4190, was obtained upon modeling the data, which suggests that the $[\text{P66614}]$-[Bentriz] and MEA react forming a strongly bound complex. Therefore, the enhanced CO2 sorption in this binary mixture is associated with the complex formed plus the contribution from the remaining free IL or MEA, depending on which is in excess. A similar effect was recently reported by Lv et al.28 studying an aqueous solution containing the $[\text{C2OHmim}]$[Gly] and the MEA, suggesting that in this case both the IL and MEA react with the CO2 through the formation of a zwitterion, which leads to the IL and MEA competing for CO2 absorption and so an enhancement in capture. This synergistic effect discussed is an added benefit from the addition of MEA to these superbase ILs. The effect is not widely discussed in literature with many groups reporting the reduction in viscosity when an amine is present as a main advantage of these systems for CO2 capture.24–26

A comparison of these systems was also performed with a physically absorbing IL (e.g., $[\text{P66614}]$[NTf2]). Due to the weak hydrogen bonding capability of the $[\text{NTf2}]^-$ anion the MEA forms a much weaker interaction with the IL compared with the superbase ILs and, therefore, its volatility is very high. For example, during the gas absorption process, significantly increased levels of MEA loss of close to 50 mol % were found after 60 min in this system compared with only <8 mol % loss from ILs containing the superbasic anion. In addition, the $[\text{P66614}]$[NTf2] only has an absorption capacity of 0.04 leading to very large increases in absorption upon the addition of MEA. Such a large difference leads to significant errors in assessing any synergy between the components.

The enhanced CO2 capture observed suggests that the CO2 competes effectively with the $-\text{OH}$ group on the MEA to reprobe the anion and both the IL and MEA components chemically absorbing the CO2. This is similar to what was reported in the previous work in which the H2O is also seen to compete effectively with the CO2.18 It is likely that this is an entropy driven reaction and the presence of a cosolvent causes a weakening of the interaction between the anion and cation giving the CO2 greater opportunity for interaction with the anion of the IL, as noted for the system in the presence of water.18 It was reported previously that the H2O disrupts the anion—cation interaction in which an enhancement was observed allowing increased interaction of the anion with CO2. In general, this occurs for the anions which have low CO2 solubility in the pure IL system. It is therefore possible that the MEA acts in a similar manner although in this case a stable complex is formed which is unlikely to be the case in the H2O containing system.

### 3.5. Desorption and Regeneration

Regeneration tests were carried out using an equimolar $[\text{P66614}]$[Bentriz]/MEA mixture at desorption temperatures of 353.15 and 393.15 K. As shown in Table 5 it was found that at 353.15 K, the CO2 absorption test for the $([\text{P66614}]$[Bentriz] + MEA) mixture after CO2 desorption at 353.15 and 393.15 K.

#### Table 5. Comparison of Recyclability of Equimolar $([\text{P66614}]$[Bentriz] + MEA) Mixture after CO2 Desorption at 353.15 and 393.15 K

<table>
<thead>
<tr>
<th>CO2 absorption test</th>
<th>nCO2/nabsorbent</th>
<th>theoretical ideal</th>
<th>nCO2/nabsorbent</th>
<th>theoretical ideal</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.51</td>
<td>0.31</td>
<td>0.51</td>
<td>0.32</td>
</tr>
<tr>
<td>2</td>
<td>0.42</td>
<td>0.32</td>
<td>0.49</td>
<td>0.23</td>
</tr>
<tr>
<td>3</td>
<td>0.24</td>
<td>0.33</td>
<td>0.42</td>
<td>0.27</td>
</tr>
</tbody>
</table>

* $u(n\text{CO}_2/n\text{absorbent}) = 0.01$; $u(T) = 0.5$ K.

4. CONCLUSIONS

In this work, five trihexyl-tetradecyolphosphonium-based ionic liquids paired with superbasic anions, namely the, 1,2,4-triazolide, 1,2,3-triazolide, benzo triazolide, benzimidazolide, and imidazolide were mixed in equimolar amounts with MEA and the effect on the CO2 capture was investigated. The results showed that, depending on the anion structure, the presence of MEA could either increase or decrease the CO2 capture ability. Increases in CO2 capacity were only observed for $[\text{P66614}]$-[123Triz] and $[\text{P66614}]$-[Bentriz] ILs, and in these cases the two components were observed to work synergistically leading to an enhanced CO2 capacity due to the formation of a stable IL/MEA complex. In all the other systems examined ideal behavior was observed. Importantly, while the addition of the MEA cosolvent decreased the viscosity compared with the pure IL, a large increase was seen upon absorption of CO2 which was much higher than that found using water as the cosolvent. After careful investigation of these systems it is apparent that water has more potential as a cosolvent compared with MEA for CO2 absorption. This is due to the large increase in viscosity upon CO2 absorption in the presence of MEA making it unsuitable for industrial use.
ASSOCIATED CONTENT

Supporting Information
The Supporting Information is available free of charge on the ACS Publications website at DOI: 10.1021/acs.jced.5b00710. Supporting data are openly available on Queen’s University Research Portal http://pure.qub.ac.uk/portal/en/datasets.

H NMR spectra for synthesized ILs are available in Figure S1. Name, source, initial, and final purity of compounds used during this work are listed in Table S1. Data of the relative volatility of the MEA in each of the investigated mixtures, the viscosity of (water + IL) mixtures and the CO2 solubility in each of the investigated mixture as the function of time are tabulated in Tables S2–S4, respectively (PDF).

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Notes
The authors declare no competing financial interest.

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