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A Guide to the Selection of Switchable Functional Groups for CO_2-Switchable Compounds

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Many CO_2-responsive species, including many of the CO_2-switchable surfactants, solvents, solutes, gels, colloids, and surfaces, rely on the ability of CO_2 to lower the pH of water. Uncharged basic groups on the CO_2-responsive species are therefore converted from a neutral state to a protonated cationic state (a bicarbonate salt), which causes dramatic and useful changes to the properties of the species. However, this switching process only works correctly if a basic group of appropriate basicity has been selected. This article presents a comprehensive guide to the selection of basic groups for CO_2-switchable species for use in water. The appropriate basicity, as measured by the pK_{aH} (the pK_a of the protonated compound), is a function of the concentration of the switchable species, the temperature, the pressure of CO_2, the presence or absence of an organic liquid phase, and the solubility of the neutral form of the compound.

1. Introduction

Switchable or stimuli-responsive materials offer great flexibility for a wide range of applications, because they resolve situations involving time-separated conflicting requirements. For example, if one needs a material to have one property for part of a process and then have a completely opposite property later in the process, then this is a time-separated conflicting requirement. Switchable materials can change their properties reversibly whenever a stimulus or trigger is applied or removed. Suitable stimuli include light, voltage, acids, bases, oxidants, reductants, and CO_2. Carbon dioxide is particularly appealing because it is a recycled, non-hazardous and easily removed waste material, it does not accumulate in the system, and it requires neither transparency or conductivity. CO_2-switchable materials so far reported include solvents,[1, 2] surfactants,[3, 4] gels,[5] adhesives,[6] surfaces,[7-10] and catalysts.[11, 12] The field was reviewed in 2012.[3] The interest in this field is increasing exponentially; there were only 2 papers published in 2006 but 92 in 2014. When CO_2-switchable materials are being designed, a suitable CO_2-responsive functional group must be incorporated into the structure. The choice of switchable functional group (SFG) depends greatly on the conditions and the concentration needed, but some general design principles exist. This paper describes the factors that influence the choice of SFG for switchable materials intended for use in water.

2. Discussion

Many factors influence the choice of SFG, including the intended temperature, pressure, and concentration of switchable material in water, plus any constraints that limit the range of acceptable pH values for the application. For example, if enzymes were to be included in the system, then extreme pH values could not be used to induce switching due to the fact that most enzymes are pH sensitive.

2.1 Reactions of CO_2 with SFG

There are four chemical reactions that have been used as the basis for CO_2-switching in water. The most common reaction, and that which is the basis for this paper, is the reaction of a neutral organic base with CO_2 and water to form a bicarbonate salt (equation 1). The basic functional group is typically an amine, an amidine, or a guanidine, but could in theory be any other basic functional group. The reaction with CO_2 converts the compound from its more hydrophobic neutral form into the hydrophilic form, the bicarbonate salt. At high pH (above 10), significant quantities of carbonate anion will form in addition to bicarbonate anion, but in most cases the pH of the aqueous solution after CO_2 addition is too far below 10 for carbonate ion formation to be significant. Reaction 1 is readily reversed if CO_2 is flushed from the system. The compound can switch back and forth between the neutral and charged states as many times as needed.

$$B + CO_2 + H_2O \underset{-CO_2}{\overset{+CO_2}{\rightleftharpoons}} [BH^+][HCO_3^-] \qquad (1)$$

If the organic base is anionic rather than neutral in its unprotonated state, then equation 2 applies. Anionic organic

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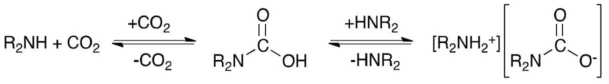
Electronic Supplementary Information (ESI) available: Derivations of the equations, experimental methods, additional graphs and tables concerning the recommended pK_{aH} ranges for the Scenarios. See DOI: 10.1039/x0xx00000x bases so far used in this manner include carboxylates and phenolates.[4] This chemistry is, in a sense, the reverse of that shown in equation 1 because here the compound is converted by CO_2 from the more hydrophilic form into the less hydrophilic compound. The physical chemistry governing the switching of such anionic compounds has been described at least briefly before,[4] and will not be covered in this paper.

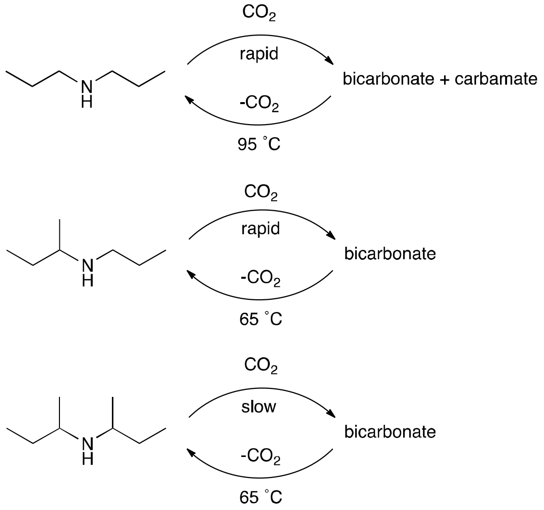
$$A^- + CO_2 + H_2O \underset{-CO_2}{\overset{+CO_2}{\rightleftharpoons}} HA + HCO_3^- \qquad (2)$$

Neutral organic bases that contain N-H bonds, including secondary and primary amines, amidines, and guanidines, could form carbamic acid and carbamate salts when exposed to CO_2 (equation 3). Much of the earliest work on CO_2-switchable compounds, primarily by the groups of Dmitry Rudkevich[13-15] and Richard Weiss,[16, 17] used this chemistry rather than bicarbonate formation. If bases having N-H bonds are reacted with CO_2 in water, the carbamic acid or carbamate salts are formed in addition to bicarbonate salts (Scheme 1, top). Carbamate formation is harder to reverse than bicarbonate formation,[18] and therefore is not preferred for applications at moderate temperatures. However, carbamates may be useful for applications in which the hydrophilic ionic form needs to be stable at higher temperatures, such as 70-90 ˚C, where a bicarbonate salt would probably be too unstable. Carbamate formation is typically faster than bicarbonate salt formation, which is advantageous. Disadvantages of carbamate formation inclue the higher temperatures and greater time required to convert back to the neutral form when CO_2 is removed, and the fact that carbamate salt formation creates only one ion per amine, whereas bicarbonate salt formation creates two ions per amine. Thus bicarbonate salt formation is preferred for applications in which the ionic form is intended to elevate the ionic strength or osmotic pressure of the solution. Similarly, carbamate salt formation from a polyamine would create zwitterionic species containing a mixture of cationic and anionic centres,[19] while bicarbonate salt formation from a polyamine creates a polycationic species containing only cationic charges. Carbamate salt formation can be avoided by choosing an organic base having a) no N-H bonds, b) particularly weak basicity, such as an aniline,[20] or c) at least one secondary or tertiary carbon attached to the N-H nitrogen atom.[21-24]



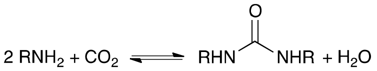
$$R_2NH + CO_2 \underset{-CO_2}{\overset{+CO_2}{\rightleftharpoons}} R_2N{-}\overset{O}{\overset{\|}{C}}{-}OH \underset{-HNR_2}{\overset{+HNR_2}{\rightleftharpoons}} [R_2NH_2^+]\left[R_2N{-}\overset{O}{\overset{\|}{C}}{-}O^-\right] \qquad (3)$$

If bicarbonate salt formation is preferred, but the rapid switching observed during carbamate salt formation is appealing, then it is possible to have the best of both worlds.[22] A secondary amine containing one secondary or tertiary alkyl group and one primary alkyl chain, such as N-propylbutan-2-amine, has the rapid reaction with CO_2 that one would expect with carbamate formation but exclusively gives bicarbonate salt as the product (Scheme 1, middle). Putting too much bulk around the nitrogen makes a secondary amine behave very much like a tertiary amine: slow reaction with CO_2 but reversion to the neutral form is reasonably facile (Scheme 1, bottom).



Scheme 1. The effects of steric hindrance on the rate of reaction with CO_2 in water and the typical temperature required to reverse the process at a reasonable rate.

The fourth reaction that can form the basis of a CO_2-switchable system is urea formation (equation 4). The conversion of a primary amine to a dialkylurea by an uncatalyzed thermal reaction with CO_2 takes place at a reasonable rate at temperatures above ~180 ˚C.[25-27] This reaction has been proposed as a switching mechanism for use in superheated water underground[28] but is not applicable to applications below the normal boiling point of water.



$$2\,RNH_2 + CO_2 \rightleftharpoons RHN{-}\overset{O}{\overset{\|}{C}}{-}NHR + H_2O \qquad (4)$$

The remainder of this paper will be concerned with the equilibria which govern bicarbonate salt formation (equation 1) and how knowledge of these equilibria can guide the choice of an SFG.

2.2 The behaviour of CO_2 in water

The solubility of CO_2 in water is low and highly dependent on temperature and pressure. Fortunately, in the region of interest (0 to 80 ˚C and 0.1 to 1.0 MPa), the solubility has a fairly linear dependence on pressure.[29] Carroll *et al.* reported an equation for calculating the Henry's Law constant (in MPa per mol fraction) of CO_2 in water between 0 and 160 ˚C (equation 5). Conversion to units of M/MPa, taking into account the temperature-dependent density of water,[30] gives the Henry's Law constants shown in Table 1.

$$\ln H = -6.8346 + 1.2817\times10^4/T - 3.7668\times10^6/T^2 + 2.997\times10^8/T^3 \qquad (5)$$

The dissolution of CO_2 in water acidifies the solution due to the dissociation of both carbonic acid and hydrated dissolved CO_2 (Figure 1). The combined dissociation constant K^*_{a1}, defined as shown in equation 6, is more useful than the formal dissociation constant K_{a1} because the former takes into account both sources of labile protons. K^*_{a1} and K_{a2} are given in Table 1 for a variety of temperatures. For temperatures other than those shown, the constants can be calculated by Cai's equations for CO_2 in freshwater (equations 7 and 8, valid from 0.2 to 35 ˚C).[31] The pH of water in equilibrium with 0.1 MPa or 1.0 MPa of CO_2 at 22 ˚C is 3.94 or 3.53, respectively.[32]

$$K^*_{a1} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3] + [CO_{2(aq)}]} \qquad (6)$$

$$pK^*_{a1} = -14.8435 + 3404.71/T + 0.032786\times T \qquad (7)$$

$$pK_{a2} = -6.4980 + 2902.39/T + 0.02379\times T \qquad (8)$$

Table 1. The Henry's Law constant and dissociation constants[33] for CO_2 in water[a] and the dissociation constant for water in pure water.[34, 35]

T, ˚C	K_H, M/MPa	pK^*_{a1}	pK_{a2}	pK_w
5	0.626	*6.52*	*10.55*	14.73
10	0.530	6.49	10.51	14.53
15	0.454	*6.42*	*10.43*	14.35
20	0.392	*6.38*	*10.38*	14.17
25	0.342	6.36	10.35	14.00
30	0.301	*6.33*	*10.29*	13.83
35	0.267	*6.31*	*10.25*	13.68
40	0.238	6.29	10.23	13.53
45	0.215			13.40
50	0.195	6.30	10.18	13.26
55	0.179			13.14
60	0.165	6.31	10.15	13.02
65	0.153			
70	0.142			
75	0.133			12.70
80	0.126	6.33	10.10	
85	0.119			
90	0.114			

[a] Values in italics were calculated using equations 7 and 8. Values in normal type are from the references.

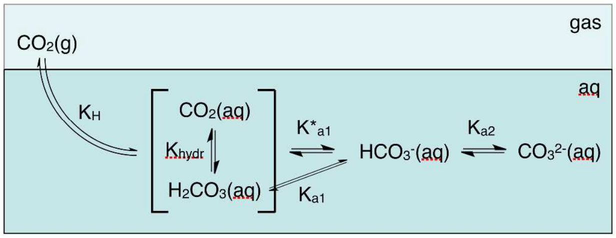


Figure 1. The equilibria involved in the dissolution of CO_2 in water.

2.3 The importance of pK_{aH} and ΔH_{prot} in SFG selection

The switching of CO_2-switchable compounds using equation 1 requires that the pH of the aqueous phase in the absence of CO_2 is above the system midpoint and the pH in the presence of CO_2 is below the system midpoint. The system midpoint is defined as the pH at which the number of moles of unprotonated base *in the system* is equal to the number of moles of protonated base in the system. The system includes the aqueous phase, any precipitated amine, and any other liquid, gaseous or supercritical phases present. Contrast this to the definition of the aqueous phase midpoint, which is defined as the pH at which the number of moles of unprotonated base *in the aqueous phase* is equal to the number of moles of protonated base in the aqueous phase. In order for a compound to be "switched" adequately by CO_2 addition, so that its properties are significantly changed, it must be converted from a largely unprotonated state to a largely protonated state. Therefore, one should choose an SFG that will ensure that the pH without CO_2 and the pH with CO_2 are on opposite sides of the system midpoint.

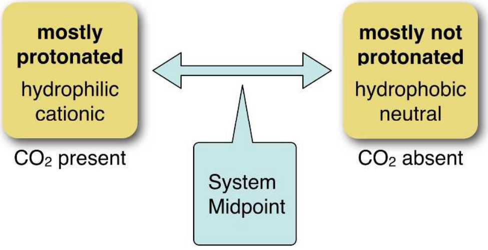


Figure 2. The principle of CO_2-switchable chemistry in water.

The system midpoint can be calculated if one knows the temperature, the concentration, the partitioning equilibria (if multiple phases are present), and the pK_{aH} of the base. The base is chosen largely for its having a pK_{aH} that is appropriate for the conditions. A pK_{aH} is defined as the pK_a of the protonated form or conjugate acid of the base. A list of representative bases and their literature pK_{aH} values is shown in Table 2. This can be used as a guide to the selection of appropriate SFG. However, because pK_{aH} values are temperature (Figure 3) and concentration[36] dependent, it is necessary to measure them at the temperature and concentration planned for an application.

Table 2. The pK_{aH} of common bases at 25 ˚C and their enthalpies of protonation.

Base	pK_{aH}	ΔH_{prot}, kJ/mol (at ˚C)
$MeN{=}C(NMe_2)_2$	13.8[37]	
$PhN{=}C(NMe_2)_2$	12.18[38]	
quinuclidine	11.4[39]	-47.2 (25)[40]
DBU	12.2[a]	
NEt_3	10.68[41]	-46.4 (25)[40]
N-methylpiperidine	10.13[41]	
NMe_3	10[42]	
4-DMAP	9.87[43]	
$Me_2NCH_2CH_2OH$	9.2[44]	
$PhCH_2NMe_2$	9.03[41]	
DABCO	8.72[41]	
N-methylmorpholine	7.8[42]	-25.7 (40)[45]
triethanolamine	7.76[46]	
2,4,6-trimethylpyridine	7.39[47]	
N-methylimidazole	7.25[48]	
pyridine	5.37[43]	-21.4 (40)[45]
$PhNMe_2$	5.15[49]	

[a] At 30 ˚C. This work.

Bases with a more negative enthalpy of protonation are able to maintain the bicarbonate form at higher temperatures, but require more energy input and more time when reversion to the neutral form is wanted. Enthalpies of protonation are shown for selected bases in Table 2.

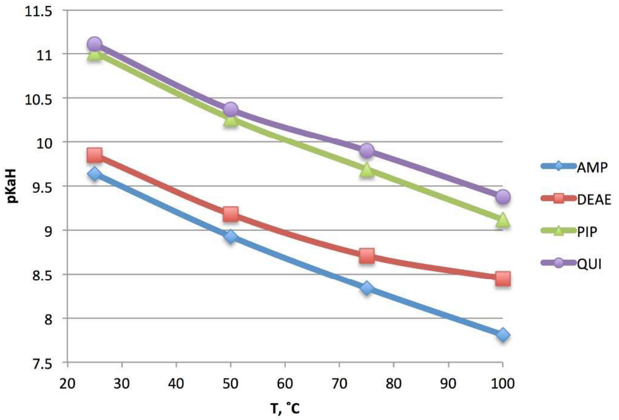


Figure 3. The pK_{aH} constants for four amines in water as a function of temperature. The amines are 2-amino-2-methyl-1-propanol (AMP), 2-diethylaminoethanol (DEAE), piperidine (PIP), and quinuclidine (QUI).[50]

When used as CO_2-switchable compounds in water, dibasic or polybasic compounds such as diamines have a great advantage over monobasic compounds such as monoamines. Given that dissolved CO_2 is not a particularly strong acid, it is normal for a switchable base to be incompletely protonated when in carbonated water. For a *monobasic* compound such as a monoamine, 90% conversion to the bicarbonate form means that 10% of the compound remains in the neutral form. This result can be catastrophic in some applications, such as the extraction of a CO_2-switchable species from an organic product into a carbonated water phase. If 10% of the amine remains neutral, then that amount of amine will likely remain in the organic product, giving a highly contaminated product and significant amine losses. In contrast, 90% conversion to the bicarbonate form of a *diamine* would result in most of the compound being diprotonated, some being monoprotonated, and very little remaining uncharged. For example, a diamine having pK_{aH1} = 10 and pK_{aH2} = 9.4 would have 90% of its N atoms protonated at a pH of 8.75, and only 1% of the diamine would still be in its neutral form. Thus for applications in which conversion of nearly all of the neutral form to charged is crucial, a diamine should greatly outperform a monoamine.

Polymeric amines have significant advantages as switchable compounds due to their non-volatility and non-toxicity, so their pK_{aH} values are also relevant to this discussion. There is not as much literature on the pK_{aH} values of polymeric amines, but the data available shows that the values are significantly lower than the pK_{aH} values of the corresponding monomers. For more discussion of this issue, see a recent tutorial paper.[51]

Simple equilibrium calculations can be used to predict how basic an SFG needs to be in order to switch from largely unprotonated to largely protonated when CO_2 is added to the aqueous solution. This is the subject of the remainder of this paper. The equations required to make such predictions are presented in three sections, in which three typical scenarios are presented. In these discussions, we assume that there are no other constraints on the system and there are no other compounds, other than the base and the CO_2, that will affect the pH. We also assume that the equilibrium constants K_H and K^*_{a1} are unaffected by the presence of the base in the solution, because the values of these constants in amine/water mixtures are unknown for the majority of amines and conditions to be discussed. Due to this assumption and because the equations are simplified, in that they use molarities rather than activities, the resulting predictions are not likely to be highly accurate but should be sufficiently accurate to guide the selection of SFG. The derivations for the equations are given in the Supplementary Information. Future work will include modelling the equilibria to account for deviations in the activity coefficients.

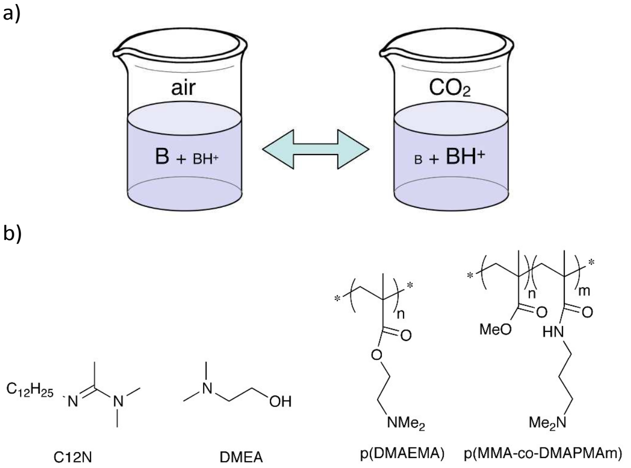
If one chooses a switchable compound having the right basicity, for the planned temperature and concentration, then simply adding that compound to water at that concentration will give a pH at which the % protonation is low, and then adding an atmosphere of CO_2 will give a pH at which the % protonation is high. Finding the right basicity or pK_{aH} to guarantee this ideal behaviour can be achieved by trial and error, but we aim to show that simple equlibrium considerations can tell us what pK_{aH} values will be appropriate. This information removes the guesswork associated with designing switchable compounds. The readers of this paper can use the equations provided or, if they prefer, can simply look at the summary graphs provided in each section (Figures

8, 12, 16, and S1 through S5) to find out what pK_{aH} their switchable compound should have.

The equations that will be presented were developed specifically for monobasic compounds having a nominal concentration of $[B]_0$, meaning the total moles of base in the system, regardless of location and degree of protonation, divided by the volume of the aqueous phase. Thus undissolved base or base dissolved in a separate organic liquid phase are included in the calculation of the total moles of base. In many cases, the predictions also apply to dibasic or polybasic compounds such as diamines and polyamines. For dibasic and polybasic compounds, $[B]_0$ is the total moles of protonatable sites in the system divided by the volume of water; thus for a diamine that is fully dissolved in water, $[B]_0$ is twice the concentration of the diamine.

3. Scenario 1: Neutral and ionic forms are fully soluble in water

In this scenario, the base and its bicarbonate salt are both fully soluble in water at all concentrations of interest (Scheme 2). As a consequence, the system midpoint is equal to the aqueous-phase midpoint and the pK_{aH} of the base.



Scheme 2. a) In scenario 1, both the neutral form and the bicarbonate salt are fully soluble in water at the concentrations used. b) Example compounds that would behave in this manner include switchable surfactants such as N-dodecyl-N',N'-dimethylacetamidine[52] and "switchable water" ionogens such as dimethylethanolamine, poly(dimethylaminoethylmethacrylate) and poly(methylmethacrylate-co-dimethylaminopropylmethacrylamide),[53] although those with more than one protonation site are mathematically more complex.

3.1 Scenario 1 in the absence of CO_2

When a neutral base is fully dissolved in pure water at a concentration $[B]_0$, under air, the resulting pH is in the basic region (Fig. 4, equation 9). The base is partly protonated due to the production of the hydroxide salt $[BH^+][OH^-]$. From $[H_3O^+]$, the % protonation can be calculated (equation 10, Figure 5). For an ideal switchable compound, the % protonation would be very low. We arbitrarily selected 5% as the upper limit for the % protonation of an ideal SFG under air. As is evident in the figure, for each base there is a minimum usable concentration, below which the % protonation exceeds 5%. For example, a base of $pK_{aH} = 11$ can only meet the <5% protonation requirement at concentrations above 380 mM. In general, stronger bases are only appropriate as CO_2-switchable groups if they are to be used at very high concentrations. In contrast, the minimum concentration for a weaker base is much lower. For example, a base of $pK_{aH} = 8$ meets the <5% protonation requirement at any concentration above 0.38 mM.

$$0 = [H_3O^+]^3 + (K_{aH}+[B]_0)[H_3O^+]^2 - K_w[H_3O^+] - K_wK_{aH} \qquad (9)$$

$$\%\ \text{protonation} = \frac{[H_3O^+]}{[H_3O^+] + K_{aH}} \times 100\% \qquad (10)$$

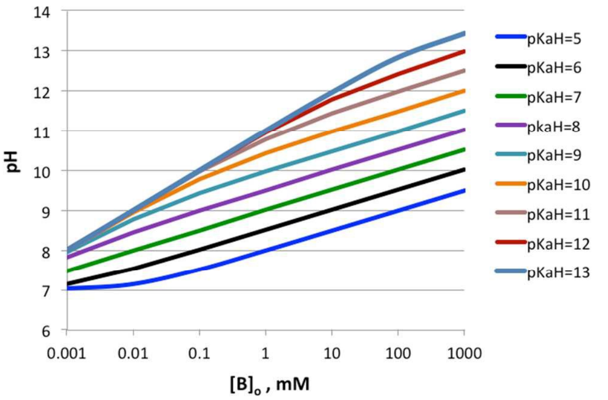


Figure 4. The pH of an aqueous solution of a base at 25 ˚C under air, as a function of the concentration and pK_{aH} of the base.

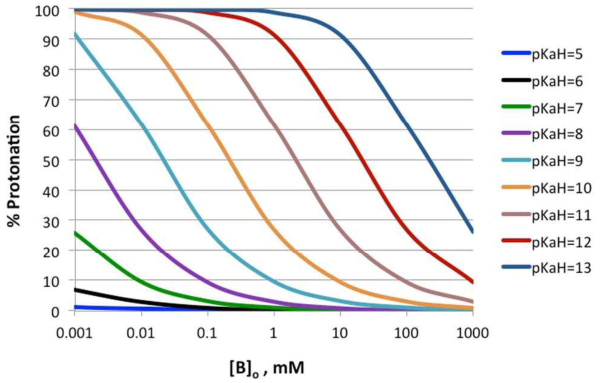


Figure 5. The % protonation of an organic base in water at 25 ˚C under air, as a function of the concentration and pK_{aH} of the base.

3.2 Scenario 1 in the presence of CO_2

Once CO_2 is introduced to the solution at a pressure P_{CO2}, the pH dramatically decreases and the base is partly converted to the bicarbonate salt. The resulting pH can be calculated using equation 11. Figure 6 shows the predicted pH for bases in water at 25 ˚C under 0.1 MPa of CO_2. Below 1 mM, all of the bases (pK_{aH} from 5 to 13) give the same pH because they are all nearly completely protonated. However, at increasing concentrations above 1 mM, there begins to be some differentiation between the weaker and stronger bases.

$$0 = [H_3O^+]^3 + (K_{aH}+[B]_0)[H_3O^+]^2 - (K^*_{a1}K_HP_{CO2}+K_w)[H_3O^+] - (K^*_{a1}K_HP_{CO2}+K_w)K_{aH} \qquad (11)$$

The % protonation of the base under CO_2 is calculated theoretically from equation 10, using $[H_3O^+]$ concentrations calculated using equation 11. Figure 7 shows the predicted % protonation for bases in water at 25 ˚C under 0.1 MPa of CO_2. An ideal switchable base would have a high % protonation under CO_2; let us assume that ≥95% would be quite acceptable. At concentrations below 1 mM, all of the bases (pK_{aH} from 6 to 13) have acceptably high % protonation, but at higher concentrations the weaker bases have unacceptably low % protonation values. The maximum usable concentration for a weaker base of pK_{aH} 6 is 1 mM but a stronger base of pK_{aH} 12 could theoretically be used up to a concentration of about 9 M. We do not show the predictions at concentrations that high because predictions above about 1 M are not likely to be accurate.

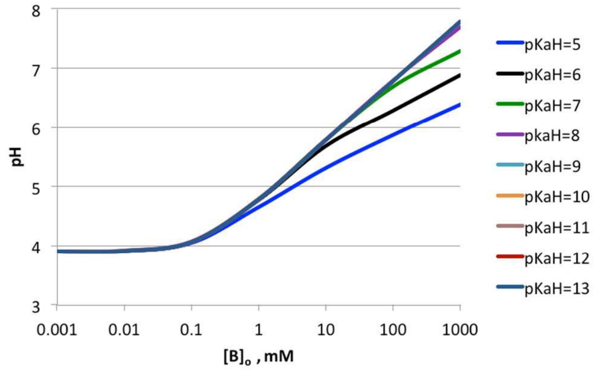


Figure 6. The pH of an aqueous solution of a base at 25 ˚C under 0.1 MPa of CO_2, as a function of the concentration and pK_{aH} of the base.

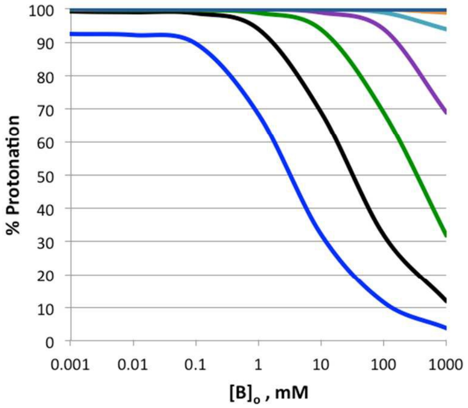


Figure 7. The % protonation of an organic base in water at 25 ˚C under 0.1 MPa of CO_2, as a function of the concentration and pK_{aH} of the base.

3.3 Scenario 1 interpretation

The ideal switchable SGF would be only slightly protonated under air, while in the presence of CO_2 would be mostly converted to the bicarbonate salt. The restrictions on the choice of pK_{aH}, as discussed above, can be summarized in a single figure. Figure 8a shows the % protonation of bases having different pK_{aH} values as a function of $[B]_0$ in the absence and presence of CO_2 at 25 ˚C and 0.1 MPa CO_2. The dashed lines represent the % protonation under air. For example, the red dashed line represents 5 % protonation under air. Any base having a pK_{aH} below that red dashed line would have an acceptably-low protonation under air. Note that this limiting line is concentration dependent. However, the base must also have a high % protonation (preferably >95%) when CO_2 is present. The solid lines represent the % protonation under 0.1 MPa of CO_2. The solid blue line represents 95% protonation; any base having a pK_{aH} above that solid blue line would have an acceptably-high protonation under CO_2. Thus the ideal switchable group would have a pK_{aH} between the red dashed line and the solid blue line, although those having pK_{aH} values slightly outside this range would probably suffice for many applications. For example, if a switchable species is needed in water at a concentration of 1 mM at 25 ˚C, then a base with a pK_{aH} between 6.1 and 8.4 would be ideal, and a pK_{aH} slightly outside that range would be acceptable but not ideal. Graphs such as Figure 8a can guide researchers in the choice of a suitable base or SFG.

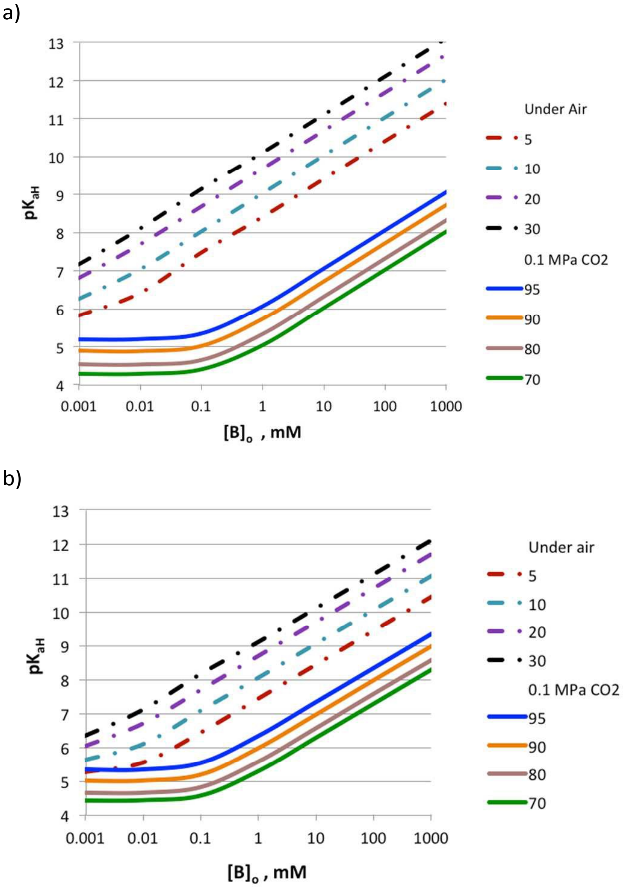


Figure 8. The pK_{aH} required for a base to have a specified % protonation when mixed with water at a) 25 ˚C and b) 60 ˚C. The dashed lines show the required pK_{aH} to obtain the specified % protonation in the absence of CO_2. The solid lines show the pK_{aH} required to obtain the specified % protonation values in the presence of 0.1 MPa of CO_2.

We also present similar graphs for other temperatures or CO_2 pressures (Figures 8b and S1 through S5). Figure 8b gives the same information but for 60 ˚C and 0.1 MPa; the lines have moved because of the temperature dependence of the equilibrium constants K_{aH}, K^*_{a1}, K_H and K_w. Comparing Figures 8a (25 ˚C) and 8b (60 ˚C), we find that the range of acceptable pK_{aH} values has narrowed considerably. The line for 5% protonation under air has moved downwards by one pK_{aH} unit, because K_w increases tenfold (pK_w drops by 1) when the temperature rises from 25 ˚C to 60 ˚C (Table 1). Because $pK_{aH} = pK_w - pK_b$, a base must have a pK_{aH} of 1 unit lower at 60 ˚C in order to have the same basicity that it had at 25 ˚C. In the presence of CO_2, raising the temperature will lead to changes in K_w and K_H but not K^*_{a1}. As discussed above, the increase in K_w lowers the required pK_{aH} but the lowering of the K_H has the opposite effect; driving equilibrium 1 to the left. Thus switchable groups must be slightly stronger bases (higher pK_{aH}) at 60 ˚C in order to have sufficient % protonation under CO_2. The narrower range of acceptable pK_{aH} values at higher temperatures suggests that careful choice of switchable functional group is more important at those temperatures.

Figures S2 through S5 show that the use of higher pressures of CO_2 (>0.1 MPa) make it possible to obtain at least 95% protonation even when weaker bases are used. Thus the acceptable range of pK_{aH} values is widened when CO_2 is used at higher pressures. This may not be important at 25 ˚C, where the range is already quite wide, but would be far more important at temperatures above 60 ˚C, where the range of acceptable pK_{aH} values would otherwise be very narrow (compare Figures 8b and S5).

Both temperature and pressure affect the recommended range of pK_{aH} values. Table 3 shows the recommended range for several temperature and pressure combinations. Similarly, raising the pressure of CO_2 widens the range of acceptable pK_{aH} values because it causes 95% protonation to be achievable with weaker bases.

The diagonal lines in Figure 8 level out, at lower concentrations, to a constant pK_{aH} because at those very low concentrations the amount of base is insufficient to modify the pH of the solution to a significant extent. In that situation, $[H_3O^+]$ is $K_w^{0.5}$ under air and $(K_w + K^*_{a1}K_HP_{CO2})^{0.5}$ under CO_2. The required pK_{aH} to obtain the desired % protonation can be calculated using equation 10. Thus for 25 ˚C, at very low $[B]_0$, the red dashed line representing 5% protonation under air (Figure 8a) becomes horizontal at a pK_{aH} of 5.7 according to equation 10 (see SI section 1.3). At concentrations below 0.1 mM, the blue solid line representing 95% protonation under CO_2 (Figure 8a) levels out at a pK_{aH} of 5.2. This value can be predicted (see SI) using equation 10 and a pH of 3.9, which is the pH of carbonated water at 0.1 MPa of CO_2. Thus at very low concentrations, the ideal switchable species at 25 ˚C would have a pK_{aH} between 5.2 and 5.7, much less basic than normally used at higher concentrations.

Table 3. The recommended range of pK_{aH} for CO_2-switchable functional groups at high and low concentrations to allow a change of % protonation from ≤5 % under air to ≥95 % under the specified pressure of CO_2.

T/ ^{o} C	P_{CO2}/MPa	pK_{aH} range	
		$[B]_o$= 1 mM	$[B]_o$= 1000 mM
25	0.1	6.1 – 8.4	9.1 – 11.4
30	0.1	6.1 – 8.3	9.1 – 11.3
60	0.1	6.3 – 7.5	9.3 – 10.5
25	1	5.1 – 8.4	8.1 – 11.4
30	1	5.2 – 8.3	8.1 – 11.3
60	1	5.4 – 7.5	8.3 – 10.5

As an example of the utility of these diagrams, we cite the use of water-soluble tertiary amines as switchable solutes for forward osmosis.[54-58] In this application, a high concentration of the trialkylammonium bicarbonate salt is used as a draw solution to pull water out of wastewater or seawater, across a membrane, and into the draw solution. CO_2 is then removed to switch the amine to its neutral form, which can then be removed by a variety of methods. This application requires a high concentration and a high % protonation. If we assume that the amine is at a concentration of 5 M, then the recommended pK_{aH} range at 25 ˚C is approximately 9.5 to 11.7. Examples of the additives that have been proposed for this purpose are trimethylamine (pK_{aH} = 10)[42] and poly(2-(dimethylamino)ethylmethacrylate) p(DMAEMA) (pK_{aH} = 7.5[59]), although the latter is significantly lower than the recommended range of pK_{aH} values and therefore will probably have a low % protonation under 0.1 MPa of CO_2.

A second example application is the use of polymeric amines as their bicarbonate salts for settling suspensions of fine clay particles, which is important in the mining of phosphate minerals and oil sands.[53] These applications use 10 to 100 ppm by weight of polymers such as p(MMA-co-DMAPMAm). At such a low loading, which roughly corresponds to an amine group nominal concentration of 0.0006 M (at 100 ppm), the recommended pK_{aH} range would be approximately 5.5 to about 8.0. p(DMAPMAm) has a pK_{aH} of about 8.8, slightly higher than the recommended range. As a consequence of the pK_{aH} being somewhat too high, the % protonation in the absence of CO_2 would be roughly 10%, so that at best only a ten-fold change in % protonation could be achieved by the addition of CO_2.

3.4 Comparing predictions to experimental data

In order to check the accuracy of the predictions, the pH values of aqueous solutions of 1,8-diazabicyclo[5.4.0]undec-7-ene (DBU, pK_{aH}=12.2), 3-(dimethylamino)-propanol (DMAPO, pK_{aH}=9.39[60]), 2-imidazol-1-yl-ethanol (2-IME, pK_{aH}=6.71), and 3-(pyridin-4-yl)propan-1-ol (PYPO, pK_{aH}=5.8) were measured at various concentrations at 30 ˚C under air and under 0.1 MPa of CO_2 (see SI for details). The pK_{aH} values for the four bases at 30 ˚C were found in the literature (DMAPO) or measured experimentally (DBU, 2-IME and PYPO). K_H, K^*_{a1} and K_w values at 30 ˚C were those in Table 1.

In the absence of CO_2, the experimental pH values (Figure 9) for aqueous solution of bases were within 0.5 of the theoretical pH values (calculated from equation 9) at concentrations of 10 and 100 mM but deviated more at 1000 mM. This suggests, not surprisingly, that the assumptions mentioned in section 2.3 lead to greater error at higher concentrations. Fortunately, the ideal ranges for pK_{aH} values for SFG are up to 3 units wide (see Figures 8, 12, and 16), so the inaccuracies caused by these assumptions are not believed to be problematic at concentrations below 1 M. Better modelling will be required for predictions at 1 M and higher. The largest deviations were observed with the weakest of the four bases, PYPO.

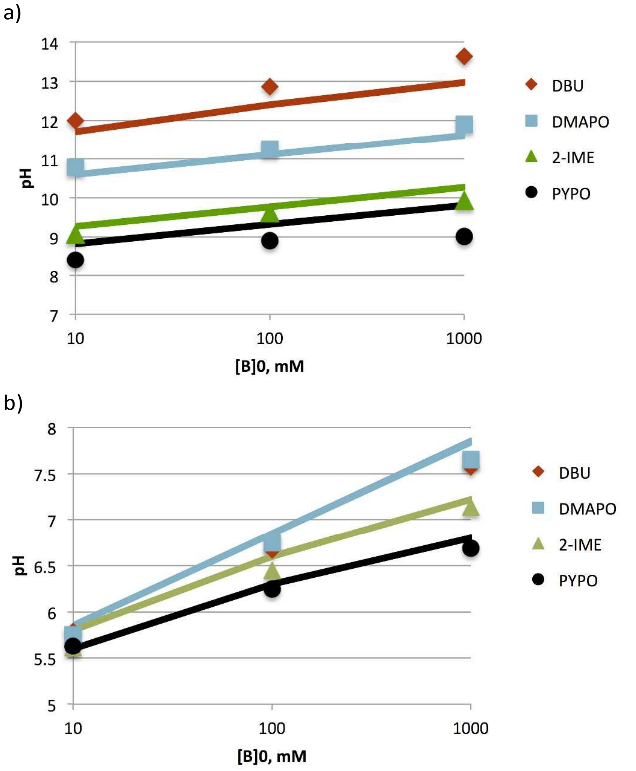


Figure 9. The experimental (symbols) and theoretical (lines) pH values of aqueous solutions of 1,8-diazabicyclo[5.4.0]undec-7-ene (DBU, pK_{aH}=12.2), 3-(dimethylamino)-propanol (DMAPO, pK_{aH}=9.39), 2-imidazol-1-yl-ethanol (2-IME, pK_{aH}=6.71), and 3-(pyridin-4-yl)propan-1-ol (PYPO, pK_{aH}=5.8) at 30 ^{o}C under a) air and b) 0.1 MPa of CO_2. In Fig. 9b, the blue line and symbols are largely covered by the red line and symbols.

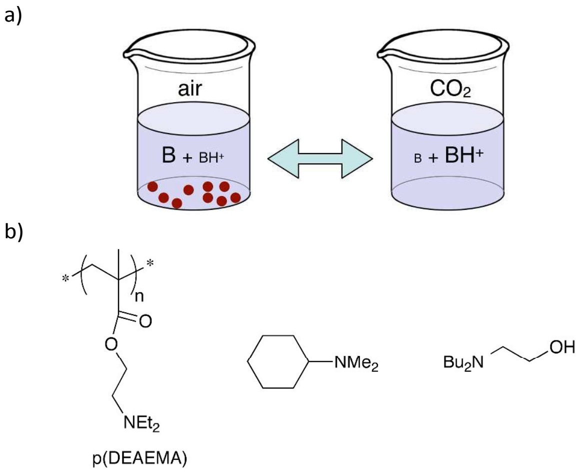
In the presence of 0.1 MPa CO_2, the differences between experimental and predicted pH values were <0.3 even at $[B]_0$ = 1 M.

4. Scenario 2: The neutral form is poorly soluble

4.1 Introduction to Scenario 2

For many organic bases, the neutral form has much less solubility in water than the bicarbonate form. In scenario 2 (Scheme 3), the neutral form has a limited solubility S_0 (in moles per litre) while the protonated form has much higher solubility. The undissolved neutral base could be solid or, in the case of SHS, liquid; the two situations are conceptually similar but the case of a liquid base would involve some loss of water to the liquid base phase. Therefore to keep the math simple, we will assume that the undissolved material is solid. For the purposes of our discussion, the bicarbonate salt will be assumed to be infinitely soluble. The moles of base will be divided amongst unprotonated molecules in the solid state, unprotonated molecules in solution, and protonated molecules in solution. The overall solubility S of the base in an aqueous solution, including protonated and unprotonated forms, is a function of the pH and the pK_{aH} (equation 12); this is the Henderson-Hasselbalch equation for solubility of bases.[61]

$$\log S = \log S_0 + \log(1 + 10^{pK_{aH}-pH}) \qquad (12)$$



Scheme 3. a) In scenario 2, the bicarbonate form is fully soluble in water but the neutral form of the base has limited solubility. The red dots indicate precipitated neutral base that may be present. There may also be precipitated neutral base under CO_2, but that would only occur at very high $[B]_0$. b) Switchable compounds having neutral forms of limited solubility in water, include solids such as PDEAEMA[54] and liquids such as SHS (e.g. N,N-dimethylcyclohexylamine and dibutylethanolamine).[62-64]

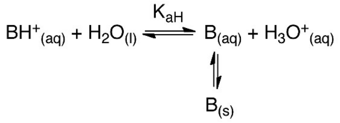
The value of S_0, the solubility of the neutral form B, can be obtained in many ways. S_0 can be measured by either a) plotting the solubility of an amine as a function of pH and taking the solubility limit that is approached as pH exceeds pK_{aH} by at least 2 units or b) by fitting equation 12 to the observed data. Alternatively, S_0 can be calculated less accurately from the observed water solubility of the amine when added to pure water (equation 13, $K_b = K_w/K_{aH}$). In a situation where S_0 needs to be predicted rather than measured, such as for computer-based screening of many possible compounds in the absence of solubility data, it can be approximately predicted from $\log K_{ow}$, the octanol/water partition coefficient. Equation 14 predicts S_0 while taking into account the molecular weight of the amine (f_i = 1.008 for aliphatic amines, 1.300 for pyridines).[65] Equation 15, which does not take into account molecular weight, is based upon observations by Box et al.[66] of neutral or basic drugs having melting points below 140 ˚C ($\log K_{ow}$ range of 2 to 5). The $\log K_{ow}$ values of amines can be fairly accurately predicted by many QSAR equations. Meylan et al.[65] have reviewed the prediction of solubility from $\log K_{ow}$ values.

$$S = S_0\left(1 + \frac{K_b}{\sqrt{K_w + S_0/K_b}}\right) \qquad (13)$$

$$\log S_0 = 0.796 - 0.854 \log K_{ow} - 0.00728\, M_W + \sum f_i \qquad (14)$$

$$\log S_0 = 0.48 - \log K_{ow} \qquad (15)$$

The partial precipitation of the neutral form has the effect of decreasing the total % protonation of the base in the system (Scheme 4). As long as the solution is saturated with neutral B, equation 16 can be used to calculate the % protonation of the base in the system at a specified pH. If [B] falls below S_0 (i.e. the solution is not saturated in B), then equations 9 and 10 should be used instead. The shift to a lower overall % protonation causes a shift of the system midpoint to a pH lower than the pK_{aH} of the base (equation 17 and Figure 10). For example, if the nominal concentration $[B]_0$ (the total moles of base in the system divided by the volume of the aqueous phase) is 100 times greater than S_0, the system midpoint will be 1.70 pH units lower than the pK_{aH} (equation 17). In Figure 10, the light blue curve indicates the behaviour when B is fully dissolved; the midpoint is then at pH = pK_{aH}. The other curves show the behaviour when B is incompletely dissolved. Each of the other curves joins the blue curve when the solution becomes sufficiently acidic that the concentration of neutral form is equal to or less than S_0.



$$\begin{tikzcd} BH^+_{(aq)} + H_2O_{(l)} \arrow[r, shift left, "K_{aH}"] & B_{(aq)} + H_3O^+_{(aq)} \arrow[l, shift left] \arrow[d, shift left] \\ & B_{(s)} \arrow[u, shift left] \end{tikzcd}$$

Scheme 4. The precipitation of neutral B shifts the dissociation equilibrium towards a lower % protonation.

$$\%\ \text{protonation} = \frac{S_0[H_3O^+]}{[B]_0 K_{aH}} \times 100\% \qquad (16)$$

$$pH_{midpoint} = pK_{aH} - \log\big([B]_0/2S_0\big) \qquad (17)$$

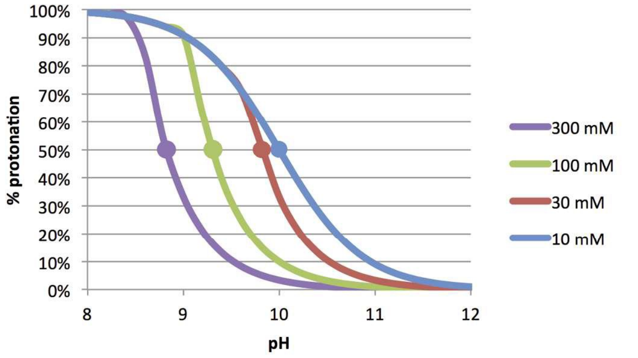


Figure 10. The overall % protonation (including both dissolved and undissolved material) versus pH for a base having a pK_{aH} of 10 and an S_0 of 10 mM, at different nominal concentrations $[B]_0$ at 25 ˚C. Circles represent the system midpoint. The light blue curve indicates the behaviour when no precipitation of the neutral form takes place (i.e. $[B]_0 \le S_0$).

4.2 Scenario 2 in the absence of CO_2

When a neutral base is mixed with pure water in amounts above the solubility limit of the base, some of base will remain undissolved and the expected pH is governed by equation 18. However, at lower amounts of base, all of the base will dissolve and equation 9 should be used instead. Figure 11 shows the pH of solutions of bases of varying pK_{aH} as a function of $[B]_0$ (the moles of base added to the system per litre of aqueous solution). At lower loadings of base, increasing the amount of base increases the pH. The inflection point found somewhat above $[B]_0 = S_0$ is the point at which the neutral form B reaches saturation. The inflection point happens at higher loadings for compounds of greater basicity. After the inflection point, further increases in the amount of base added have no effect on the solution pH.

$$[H_3O^+] = K_w^{0.5}/(1 + S_0/K_{aH})^{0.5} \qquad (18)$$

In order to build a graph that can guide the selection of bases as switchable species in Scenario 2, we need to be able to predict the % protonation at the pH that would be generated naturally by a base/water mixture. The equation for that can be obtained by combining equations 16 and 18, giving equation 19. In Fig. 12, the dashed line for 5% protonation under air deviates upwards as the concentration passes S_0, which is 0.01 M for the example shown. The upwards deviation indicates that particularly strong bases can still meet the ≤5% protonation requirement if much of the base is precipitated in its neutral form.

$$\%\ \text{protonation} = \frac{S_0 K_w^{0.5}}{[B]_0 K_{aH}\sqrt{1 + S_0/K_{aH}}} \times 100\% \qquad (19)$$

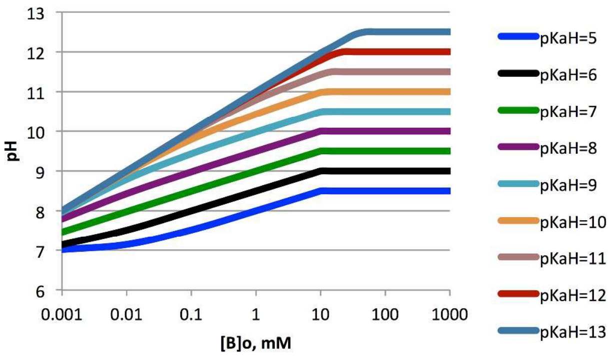


Figure 11. The pH of the aqueous phase of a base/water mixture in the absence of CO_2 as a function of the pK_{aH} of the base and $[B]_0$, the number of moles of base added to the system per litre of solution, at 25 ˚C. For each base, S_0 is assumed to be 10 mM.

4.3 Scenario 2 in the presence of CO_2

Once a base of limited solubility has been added to water, the addition of CO_2 lowers the pH to a value governed by equation 20 if some base remains undissolved, and equation 11 if all of the base has dissolved. The % protonation of the base will be that given by equations 21 (some undissolved) and 10 (all dissolved).

$$[H_3O^+] = \sqrt{\frac{K_w + K^*_{a1} K_H P_{CO2}}{1 + S_0/K_{aH}}} \qquad (20)$$

$$\%\ \text{protonation} = \frac{S_0\sqrt{\dfrac{K_w + K^*_{a1} K_H P_{CO2}}{1 + S_0/K_{aH}}}}{[B]_0 K_{aH}} \times 100\% \qquad (21)$$

These equations can be used to determine the appropriate pK_{aH} for a switchable species at a given nominal concentration $[B]_0$. If we assume that an ideal switchable species will be at least 95% protonated in the presence of CO_2 and less than 5% protonated in the absence of CO_2, then the lines shown in Figure 12 can guide the selection of an appropriate base to use as a CO_2-switchable species for Scenario 2. Any base having a pK_{aH} below the dashed line and above the solid line will meet our requirements for an ideal SFG. Bases just outside this range would still switch, but the degree of change in % protonation and in properties would decrease the further the pK_{aH} was from the ideal range.

For example, the switchable-hydrophilicity solvent (SHS) $CyNMe_2$ has an S_0 of about 110 mM (see Supplementary Information) and, when mixed 1:1 v/v with water, a $[B]_0$ of 6.6 M. If we want to predict what pK_{aH} would be most appropriate for such a switchable species, we would need to extending the lines of Fig. 12 to concentrations above 1 M, despite the fact that greater inaccuracy is expected at such concentrations, as discussed in Section 3.4. By doing so, we can anticipate that such an SHS should have a pK_{aH} in the order of 10 to 13. The pK_{aH} of $CyNMe_2$ is 10.48.[67] Speaking more generally, we can conclude from Fig. 12 that SHS with lower S_0 values should have higher pK_{aH} values. That is consistent with previous observations that more hydrophobic SHS must be more basic.[68]

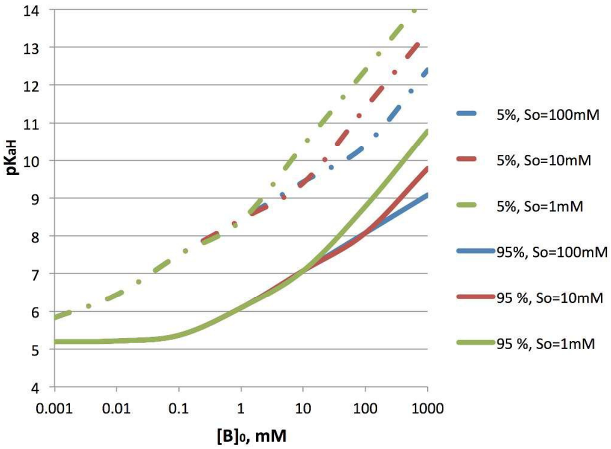
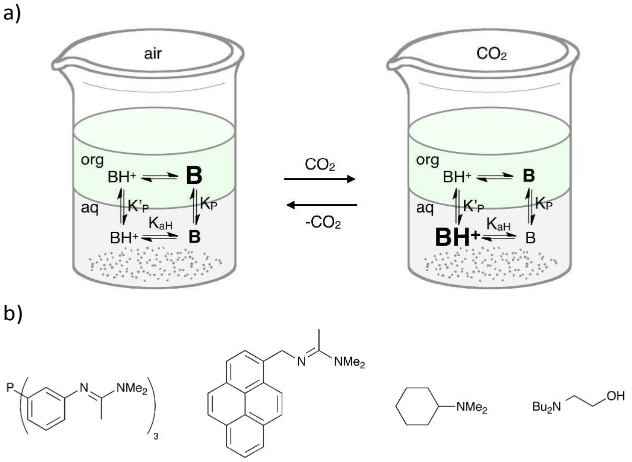


Figure 12. The pK_{aH} required for a base to have 5% protonation when mixed with water under air (dashed lines) or a 95% protonation in water under 0.1 MPa CO_2 (solid lines), for incompletely soluble bases having an S_0 value of 100 mM, (blue), 10 mM (red) or 1 mM (green), at 25 ˚C. The data at $[B]_0 < S_0$ are the same as those in Figure 8. For example, a switchable agent having a S_0 of 0.01 M would ideally have a pK_{aH} falling between the blue dashed and blue solid lines.

5. Scenario 3: Both forms partition between an aqueous and an organic phase

5.1 Scenario 3 in the absence of CO_2

When an organic phase is present in addition to the aqueous phase, then solutes such as switchable bases partition between the two phases (Scheme 5). The partition coefficient for the neutral form of the base is K_P, defined as shown in equation 22. The partition coefficient K'_P for the protonated form is as shown in equation 23. The distribution coefficient, D, governs the partitioning of both the neutral and protonated forms of the base (equation 24). The distribution coefficient is approximately equal to K_P at high pH, K'_P at low pH, and strongly pH dependent at intermediate pH (Figure 13). Equation 25 describes the pH dependence of D.[69]



Scheme 5. a) In scenario 3, the neutral and protonated forms are fully soluble but are partitioning between an aqueous phase and an organic phase. The addition or removal of CO_2 changes the concentrations of each species in each phase. b) Switchable compounds that behave in this manner, when placed in a biphasic mixture of water and an organic liquid, include ligands such as *N*',*N*'',*N*'''-(phosphinetriyltris(benzene-3,1-diyl))tris(*N*,*N*-dimethylacetimidamide),[12] dyes such as N,N-dimethyl-N'-(pyren-1-ylmethyl)acetimidamide,[70] and SHS such as N,N-dimethylcyclohexylamine and N,N-dibutylethanolamine.[63, 64]

$$K_P = \frac{[B]_{org}}{[B]_{aq}} \qquad (22)$$

$$K'_P = \frac{[BH^+]_{org}}{[BH^+]_{aq}} \qquad (23)$$

$$D = \frac{[B]_{org} + [BH^+]_{org}}{[B]_{aq} + [BH^+]_{aq}} \qquad (24)$$

$$D = \frac{K_P K_{aH} + K'_P[H_3O^+]_{aq}}{K_{aH} + [H_3O^+]_{aq}} \qquad (25)$$

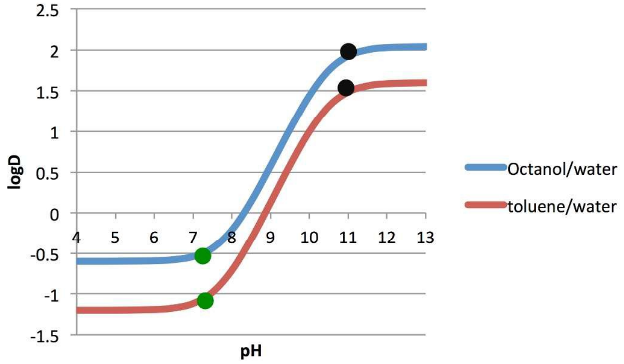


Figure 13. The pH dependence of the octanol/water (upper curve) and toluene/water (lower curve) distribution coefficients of cyclohexyldimethylamine. The black and green dots represent the distribution coefficients obtained in the absence and presence of 1 bar of CO_2, respectively.[68]

The % protonation of a base in an organic/aqueous biphasic mixture under air is described by equation 26, where V_{rat} is the ratio of the volumes of the organic and aqueous phases (V_{org}/V_{aq}).[68] A graph of the % protonation (Figure 14) for a hypothetical base shows that the presence of an organic phase lowers the overall % protonation. This change is expected because the low polarity of the organic phase would stabilize B relative to its hydroxide salt $[BH^+][OH^-]$. For example, at pH 10, this particular base would have a % protonation of 50% if dissolved in just water. In contrast, if dissolved in a organic/water biphasic solvent mixture (V_{rat} = 0.5), this same base would have a % protonation of 50% in the aqueous phase and a very low % protonation in the organic phase, so that the % protonation in the entire system would be only 7%. The majority of the base has partitioned into the organic phase, so that the overall % protonation drops greatly. As a result of the lowered % protonation when an organic phase is added, the system midpoint moves to a lower pH, sometimes by 2 or more units. The difference between pK_{aH} and the system midpoint pH is a function of K_P, K'_P, and V_{rat} (equation 27).

$$\%\ \text{protonation} = \frac{[H_3O^+]_{aq}(K'_PV_{rat} + 1)}{K_{aH}(K_PV_{rat} + 1) + [H_3O^+]_{aq}(K'_PV_{rat} + 1)} \times 100\% \qquad (26)$$

$$pH_{midpoint} = pK_{aH} - \log_{10}(K_PV_{rat} + 1) + \log_{10}(K'_PV_{rat} + 1) \qquad (27)$$

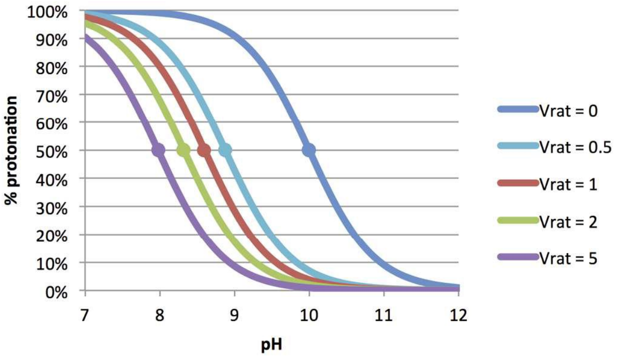


Figure 14. The dependence of the % protonation of a hypothetical amine on the pH in an organic/aqueous biphasic mixture in the absence of CO_2. The amine is assumed to have pK_{aH} = 10, $\log K_P$ = 1.4 and $\log K'_P$ = -1.4. The system midpoint (indicated by a filled circle) moves to much lower pH values as V_{rat} is increased from 0 to 5.

When a neutral base is added to an organic/aqueous biphasic mixture in the absence of CO_2, the aqueous phase pH rises to a value given by equation 28. Here, $[B]_0$ is defined as the number of moles of base added to the system divided by V_{aq}. If V_{rat} = 0, then this equation simplifies to equation 9 that we saw in the first scenario. Figure 15 shows the dependence of the aqueous-phase pH on $[B]_0$, for a hypothetical base having a $\log K_P$ of 1.4 and a $\log K'_P$ of -1.4. Note that all of the lines are lowered by about 0.5 pH units relative to the positions of the same lines in Figure 4 (scenario 1) because the organic phase withdraws some of the base from the aqueous phase.

$$0 = \mathbf{[H_3O^+]^3_{aq}}(K'_PV_{rat} + 1) + \mathbf{[H_3O^+]^2_{aq}}\{K_{aH}(K_PV_{rat} + 1) + [B]_0\} - \mathbf{[H_3O^+]_{aq}}K_w(K'_PV_{rat} + 1) - K_{aH}K_w(K_PV_{rat} + 1) \qquad (28)$$

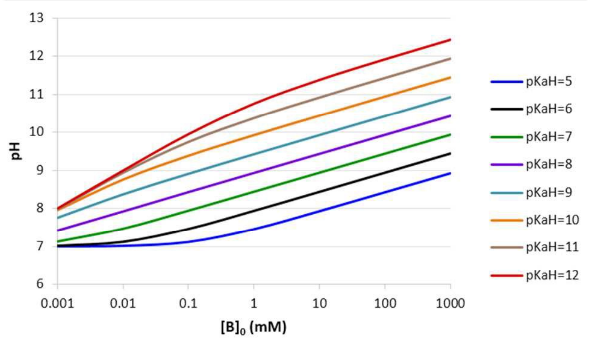


Figure 15. The aqueous-phase pH obtained when an amine or similar base is added to an organic/aqueous biphasic system in the absence of CO_2. $[B]_0$ = moles of base added/V_{aq}. The bases are assumed to have $\log K_P$ = 1.4, $\log K'_P$ = -1.4, and V_{rat} is assumed to be 0.5.

5.2 Scenario 3 in the presence of CO_2

When a base is added to a biphasic mixture containing an aqueous phase and an organic phase, and then CO_2 is introduced to the system, the pH of the aqueous phase is governed by equation 29. The % protonation and the system midpoint pH are still governed by equations 26 and 27, respectively.

$$0 = \mathbf{[H_3O^+]^3_{aq}}(K'_PV_{rat} + 1) + \mathbf{[H_3O^+]^2_{aq}}\{K_{aH}(K_PV_{rat} + 1) + [B]_0\} - \mathbf{[H_3O^+]_{aq}}\{K_w(K'_PV_{rat} + 1) + K_{aH}K^*_{a1}K_HP_{CO2}(K'_PV_{rat} + 1)\} - K_{aH}K_w(K_PV_{rat} + 1) - K_{aH}K^*_{a1}K_HP_{CO2}(K_PV_{rat} + 1) \qquad (29)$$

In cases where an organic phase exists in addition to an aqueous phase, equations 26-29 can be used to determine the appropriate pK_{aH} for a switchable species at a given $[B]_0$. While for scenario 2 $[S]_0$ affected the % protonation of a base, in this scenario the % protonation is affected by K_P. In describing an ideal switchable compound that will be less than 5% protonated under air and more than 95% protonated under 1 atm of CO_2, Figure 16 can be used to select an appropriate switchable base for scenario 3. The figure was created assuming that V_{rat} = 0.5 and that $\Delta\log K_P$, the difference between $\log K_P$ and $\log K'_P$, is 2.8, a value experimentally determined for $CyNMe_2$.[68] A base with a pK_{aH} below the dashed line will be less than 5% protonated under air while a base with a pK_{aH} above the solid line will be more than 95% protonated under 1 atm of CO_2. The ideal switchable base would have a pK_{aH} between these lines. For example, if you anticipate using a base having a $\log K_P$ of 3, then you should choose one having a pK_{aH} between the solid and dashed yellow lines.

The effect of the organic phase is an increase in the pK_{aH} required to have an ideal switchable base. When an organic phase is present, some of the base will partition into that phase, where the base is less likely to be protonated. As a result, the pK_{aH} of a base must be larger in order to have the same % protonation at a given $[B]_0$ if an organic phase is present than if only an aqueous phase is present. The magnitude of this effect is related to the hydrophobicity of the base, represented by $\log K_P$. More hydrophobic bases with larger $\log K_P$ values will have increased pK_{aH} requirements when compared to compounds with smaller $\log K_P$ values.

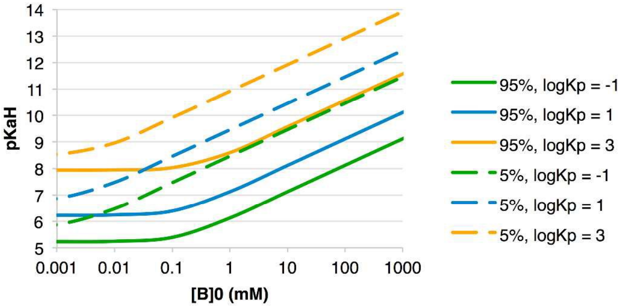


Figure 16. The pK_{aH} required for a base to have 5% protonation when placed in a water:organic biphasic mixture under air (dashed lines) or a 95% protonation in the biphasic mixture under 0.1 MPa CO_2 (solid lines), for bases having a $\log K_P$ value of -1, (green), 1 (blue) or 3 (orange), at 25 ˚C. A base having a pK_{aH} between the dashed and solid lines, at the concentration desired, should be selected. A $\Delta \log K_P$ of 2.8 was assumed.

Equation 29 and Figure 16 are not valid in all situations. First, K_P values are determined for compounds at very low concentrations. Depending on the value chosen for $[B]_0$, the observed partitioning may deviate from the partitioning predicted by K_P and K'_P. Second, the base must be sufficiently soluble in both the aqueous and organic phases. Compounds with high $\log K_P$ values are likely poorly soluble in water. The equations used in scenario 2 are likely more appropriate for modelling poorly soluble compounds. Compounds with low $\log K_P$ could likewise be modelled using the equations discussed for scenario 1, though the equations 26, 28, and 29 give similar answers to equations 9-11 when $\log K_P$ is sufficiently low.

6. Conclusions

CO_2-triggered switching requires that the organic base switch from mostly not protonated, in the absence of CO_2, to mostly protonated in the presence of CO_2. This requires that the addition and later removal of CO_2 moves the pH back and forth across the system midpoint. The system midpoint is defined as the aqueous-phase pH at which half of the base in the system is protonated. Careful selection of the switchable functional group, and particularly in terms of its pK_{aH}, ensures that the above requirements will be met at the conditions and concentrations needed for an application.

We have proposed that an ideal CO_2-switchable compound in water would have <5% protonation under air and >95% protonation under CO_2. The basicity and therefore the kind of SFG that can meet this requirement depends strongly on the concentration, temperature, CO_2 pressure, solubility of the neutral form, and the presence or absence of an organic liquid phase. Higher concentrations necessitate the use of a stronger base. Higher CO_2 pressures have the effect of widening the range of basicities that can meet the ideal requirements. In contrast, higher temperatures have the unfortunate effect of narrowing the range of basicities that can meet the ideal requirements, so that above 60 ˚C one must either accept a lesser degree of switching or use a CO_2 pressure greater than 1 bar. The use of a CO_2-switchable compound that has poor aqueous solubility in its neutral form makes it necessary to use a stronger base to achieve ideal switching. The presence of an organic liquid phase also necessitates the use of a stronger base.

The equations and graphs presented in this paper should help researchers select the most appropriate pK_{aH} for their switchable compound, and from that information the most appropriate kind of SFG.

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8. Notes and references

1. P. G. Jessop, D. J. Heldebrant, L. Xiaowang, C. A. Eckert and C. L. Liotta, *Nature*, 2005, **436**, 1102.
2. P. G. Jessop, *Aldrichimica Acta*, 2015, **48(1)**, 18-21.
3. P. G. Jessop, S. M. Mercer and D. J. Heldebrant, *Energy Env. Sci.*, 2012, **5**, 7240-7253.
4. E. Ceschia, J. R. Harjani, C. Liang, Z. Ghoshouni, T. Andrea, R. S. Brown and P. G. Jessop, *RSC Advances*, 2014, **4**, 4638-4645.
5. E. Carretti, L. Deia and R. G. Weiss, *Soft Matter*, 2005, **1**, 17-22.
6. T. Yu, K. Wakuda, D. L. Blair and R. G. Weiss, *J. Phys. Chem. C*, 2009, **113**, 11546-11553.
7. F. S. Mohammed, S. Wuttigul and C. L. Kitchens, *Ind. Eng. Chem. Res.*, 2011, **50**, 8034-8041.
8. S. Kumar, X. Tong, Y. L. Dory, M. Lepage and Y. Zhao, *Chem. Commun.*, 2013, **49**, 90-92.
9. K. J. Boniface, R. R. Dykeman, A. Cormier, H.-B. Wang, S. M. Mercer, G. Liu, M. F. Cunningham and P. G. Jessop, *Green Chem.*, 2016, **18**, 208-213.
10. Q. Zhang, G. Q. Yu, W. J. Wang, H. M. Yuan, B. G. Li and S. P. Zhu, *Langmuir*, 2012, **28**, 5940-5946.

11. A. Andreetta, G. Barberis and G. Gregorio, *Chimica e l'Industria*, 1978, **60**, 887-891.
12. S. L. Desset and D. J. Cole-Hamilton, *Angew. Chem. Int. Ed.*, 2009, **48**, 1472-1474.
13. H. Xu and D. M. Rudkevich, *Chem. Eur. J.*, 2004, **10**, 5432-5442.
14. E. M. Hampe and D. M. Rudkevich, *Tetrahedron*, 2003, **59**, 9619-9625.
15. E. M. Hampe and D. M. Rudkevich, *Chem. Commun.*, 2002, 1450-1451.
16. M. George and R. G. Weiss, *Langmuir*, 2003, **19**, 1017-1025.
17. M. George and R. G. Weiss, *J. Am. Chem. Soc.*, 2001, **123**, 10393-10394.
18. F. Barzagli, F. Mani and M. Peruzzini, *Energy Environ. Sci.*, 2009, **2**, 322-330.
19. D. J. Heldebrant, P. K. Koech, T. Ang, C. Liang, J. E. Rainbolt, C. R. Yonker and P. G. Jessop, *Green Chem.*, 2010, **12**, 713.
20. S. M. Fields and K. Grolimund, *J. High Resol. Chrom., Chrom. Commun.*, 1988, **11**, 727-729.
21. G. Sartorl and D. W. Savage, *Ind. Eng. Chem. Fundam.*, 1983, **22**, 239-249.
22. F. Bougie and M. C. Iliuta, *J. Chem. End. Data*, 2012, **57**, 635-669.
23. J. T. Yeh and H. W. Pennline, *Energy & Fuels*, 2001, **15**, 274-278.
24. E. F. da Silva and H. F. Svendsen, *Ind. Eng. Chem. Res.*, 2006, **45**, 2497-2504.
25. J. F. Mulvaney and R. L. Evans, *Ind. Eng. Chem.*, 1948, **40**, 393-397.
26. T. Hayashi and M. Kuyama, *Nat. Sci. Rept. Ochanomizu Univ. (Tokyo)*, 1951, **2**, 79-86.
27. T. V. Charlu and M. R. A. Rao, *Ind. J. Technol.*, 1969, **7**, 111-114.
28. H. B. Jung, K. C. Carroll, S. Kabilan, D. J. Heldebrant, D. Hoyt, L. Zhong, T. Varga, S. Stephens, L. Adams, A. Bonneville, A. Kuprata and C. A. Fernandez, *Green Chem.*, 2015, **17**, 2799-2812.
29. J. J. Carroll, J. D. Slupsky and A. E. Mather, *J. Phys. Chem. Ref. Data*, 1991, **20**, 1201.
30. NIST Chemistry WebBook, NIST Standard Reference Database Number 69, http://webbook.nist.gov.chemistry/, (accessed May 2015).
31. W.-J. Cai and Y. Wang, *Limnol. Oceanogr.*, 1998, **43**, 657-668.
32. C. Roosen, M. Ansorge-Schumacher, T. Mang, W. Leitner and L. Greiner, *Green Chem.*, 2007, 455-458.
33. A. Stefánsson, P. Bénézeth and J. Schott, *Geochimica et Cosmochimica Acta*, 2013, **120**, 600–611.
34. H. S. Harned and R. A. Robinson, *Trans. Faraday Soc.*, 1940, **36**, 973-978.
35. A. V. Bandura and S. N. Lvov, *J. Phys. Chem. Ref. Data, Vol. 35, No. 1, 2006*, 2006, **35**, 15-30.
36. J. Armstrong and R. B. Barlow, *Br. J. Pharmac.*, 1976, **57**, 501-516.
37. S. J. Angyal and W. K. Warburton, *J. Chem. Soc.*, 1951, 2492-2494.
38. P. Pruszynski, *Can. J. Chem.*, 1987, **65**, 626-629.
39. E. A. Castro, P. R. Campodonico, Contreras, Fuentealba, J. G. Santos, J. R. Leis, L. Garcia-Rio, J. A. Saez and L. R. Domingo, *Tetrahedron*, 2006, **62**, 2555-2562.
40. V. B. Naidoo and M. Sankar, *J. Solution Chem.*, 2011, **30**, 825-836.
41. V. Frenna, N. Vivona, G. Consiglio and D. Spinelli, *J. Chemi. Soc., Perkin Trans. 2*, 1985, 1865-1868.
42. I. T. Ibrahim and A. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1982, 1459-1466.
43. E. A. Castro, M. Cubillos and J. G. Santos, *J. Org. Chem.*, 2004, **69**, 4802-4807.
44. M. Ahmad, A. Masohan and S. S. Sawhney, *Asian J. Chem.*, 2014, **26**, 975-980.
45. K. Ballerat-Busserolles, M. R. Simond, Y. Coulier and J.-Y. Coxam, *Pure Appl. Chem.*, 2014, **86**, 233–243.
46. D. Xiong, Z. Li, H. Wang and J. Wang, *Green Chem.*, 2013, **15**, 1941-1948.
47. Y. Kitamura and T. Itoh, *J. Solution Chem.*, 1987, **16**, 715-726.
48. J. Catalan, R. M. Claramunt, J. Elguero, J. Laynez, M. Menendez, F. Anvia, J. H. Quian, M. Taagepera and R. W. Taft, *J. Am. Chem. Soc.*, 1988, **110**, 4105-4111.
49. I. Kaljurand, R. Lilleorg, A. Murumaa, M. Mishima, P. Burk, I. Koppel, I. A. Koppel and I. Leito, *Journal of Physical Organic Chemistry*, 2013, **26**, 171-181.
50. P. V. Balakrishnan, *J. Solution Chem.*, 1988, **17**, 825-840.
51. M. F. Cunningham and P. G. Jessop, *Eur. Polymer J.*, 2016, **76**, 208–215.
52. Y. Liu, P. G. Jessop, M. Cunningham, C. A. Eckert and C. L. Liotta, *Science*, 2006, **313**, 958-960.
53. T. Robert, S. M. Mercer, T. J. Clark, B. E. Mariampillai, P. Champagne, M. F. Cunningham and P. G. Jessop, *Green Chem.*, 2012, **14**, 3053-3062.
54. P. G. Jessop and S. Mercer, *International Patent application,* PCT/CA2011/050777, 2011.
55. P. G. Jessop and S. Mercer, *International Patent application Pat.,* PCT/CA2011/050075, 2011.
56. M. Ikeda, K. Miyamoto, *European Patent Application 11829222.6,* 2011.
57. Y. Cai, W. Shen, R. Wang, W. B. Krantz, A. Fane and X. Hu, *Chem. Commun.*, 2013, **49**, 8377-8379.
58. C. Boo, Y. Khalil and M. Elimelech, *J. Membrane Sci.*, 2015, **473**, 302–309.
59. P. van de Wetering, N. J. Zuidam, M. J. van Steenbergen, O. A. G. J. van der Houwen, W. J. M. Underberg and W. E. Hennink, *Macromolecules*, 1998, **31**, 8063-8068.
60. M. R. Simond, K. Ballerat-Busserolles, Y. Coulier, L. Rodier and J.-Y. Coxam, *J. Solution Chem.*, 2012, **41**, 130-142.
61. G. Volgyi, E. Baka, K. J. Box, J. E. A. Comer and K. Takács-Novák, *Anal. Chim. Acta*, 2010, **673**, 40-46.
62. P. G. Jessop, L. Phan, A. Carrier, S. Robinson, C. J. Dürr and J. R. Harjani, *Green Chem.*, 2010, **12**, 809-814.
63. P. G. Jessop, L. Kozycz, Z. G. Rahami, D. Schoenmakers, A. R. Boyd, D. Wechsler and A. M. Holland, *Green Chem.*, 2011, **13**, 619-623.
64. J. R. Vanderveen, J. Durelle and P. G. Jessop, *Green Chem.*, 2014, **16**, 1187-1197.
65. W. M. Meylan, P. H. Howard and R. S. Boethling, *Env. Toxicol. Chem.*, 1996, **15**, 100-106.
66. K. Box, J. E. Comer, T. Gravestock and M. Stuart, *Chemistry & Biodiversity*, 2009, **6**, 1767-1788.

67. J. M. Devereux, K. R. Payne and E. R. A. Peeling, *J. Chem. Soc.*, 1957, 2845-2851.
68. J. Durelle, J. R. Vanderveen and P. G. Jessop, *PhysChemChemPhys*, 2014, **16**, 5270-5275.
69. S. Carda-Broch and A. Berthod, *J. Chromatography A*, 2003, **995**, 55-66.
70. L. Phan and P. G. Jessop, *Green Chemistry*, 2009, 307-308.

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