THE JOURNAL OF PHYSICAL CHEMISTRY A LETTERS

2009, *113,* 8152–8156
Published on Web 07/01/2009

Acid Dissociation versus Molecular Association of Perfluoroalkyl Oxoacids: Environmental Implications

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Received: June 1, 2009; Revised Manuscript Received: June 16, 2009

Perfluorooctanoate (PFO) and perfluorooctanesulfonate (PFOS) surfactant anions, once released, may rapidly reach remote regions. This phenomenon is puzzling because the water-bound anions of strong F-alkyl acids should be largely transported by slow oceanic currents. Herein, we investigate whether these hydrophobic F-alkyl oxoanions would behave anomalously under environmental conditions, as suggested elsewhere. Negative electrospray ionization mass spectra of micromolar aqueous PFO or PFOS solutions from pH 1.0 to 6.0 show (1) $m/z = 499$ (PFOS) signals that are independent of pH and (2) $m/z = 413$ (PFO) and 369 (PFO − CO_2) signals, plus $m/z = 213$ ($C_3F_7CO_2^-$) and 169 ($C_3F_7^-$) signals at higher collision energies, and, below pH ∼ 4, $m/z = 827$ signals from a remarkably stable $(PFO)_2H^-$ cluster that increase with decreasing pH. Since the sum of the $m/z = 369$, 413, and 827 signal intensities is independent of pH, that is, effectively encompasses all major species, we infer that $pK_a(PFOSA) < 1.0$ and $pK_a(PFOA) < 1.0$. We also derive $K_2 \leq 4 \times 10^7$ M^{-2} for the clustering equilibrium $2PFO + H^+ \rightleftharpoons (PFO)_2H$. Thus, although $(PFO)_2H$ is held together by an exceptionally strong homonuclear covalent hydrogen bond, neither PFOS nor PFO will associate or protonate significantly at environmentally relevant subnanomolar concentrations above pH ∼ 1.

Introduction

Perfluoroalkyl (F-alkyl) chemicals (PFCs) began to be produced and commercialized about 50 years ago.[1−3] Exceptional chemical inertness confers on these materials valuable properties but also ensures unwanted environmental persistence.[4,5] As a result, they have spread and bioaccumulated globally with unforeseeable consequences.[5−13] The most conspicuous congeners perfluorooctanoate (PFO) and perfluorooctanesulfonate (PFOS) have been detected in surface waters and precipitation,[14−16] sediments,[17] and biota worldwide.[18−22] F-alkyl oxoanions apparently perturb peptide chains and DNA strand conformations via noncovalent, entropy-driven interactions.[9,11,23,24]

The rapid decline of PFOS levels in Canadian Arctic seals following its phaseout in 2000 strongly suggests an atmospheric transport mechanism[25] and defies the notions that oceans are the ultimate sink and that slow ocean currents are the long-range conduits for these weakly basic F-alkyl oxoanions.[26−28] The issue of whether marine aerosols enriched in these anionic surfactants[29,30] or their gas-phase conjugated acids mediate atmospheric transport[31] clearly hinges on the extent of F-alkyl oxoacids dissociation under environmental conditions.[32] Their long-range transport can also be indirectly effected, in part, by degradable gas-phase precursors. Although the powerful electron-withdrawing F-alkyl chains demonstrably stabilize these anions, namely, $pK_a(CF_3COOH) = 0.3$ versus $pK_a(CH_3COOH) = 4.8$,[33] and more than ∼8 CH_2 links are required to insulate functional groups from F-alkyl segments,[34] the acidity of PFOA remains elusive. Titrations in water/alcohol solvents yielded $pK_a(PFOA) = 2.8$ and 3.8,[35,36] whereas SPARC/COSMO models[37] and semiempirical PM6 computations[38] predict $pK_a(PFOA) \leq 0.7$. The significantly larger than predicted experimental $pK_a(PFOA)$ values have been tentatively ascribed to the aggregation of hydrophobic PFOA (note that PFO aggregation should have the opposite effect) in aqueous solvents at amenable laboratory mM concentrations.[39−41] Herein, we address these basic issues[36,42−46] and report experiments on the speciation of the PFOA (perfluorooctanoic acid) and PFOSA (perfluorooctanesulfonic acid, not to be confused with perfluorooctane sulfonamide) in micromolar aqueous solutions as a function of pH via pneumatically assisted electrospray ionization mass spectrometry (ESI-MS).

Experimental Section

$PFONH_4$ and PFOSK (3M), $NaClO_4$ (EM Science, >99%), Na-hexanoate and Na-octanoate (Sigma Aldrich, >99%), and 3 M NaOH and 6 M HCl solutions (VWR, reagent grade) were used as received. Aqueous solutions were prepared with purified water from a Millipore Milli-Q system (18.2 MΩ cm resistivity). Aqueous 1−10 μM PFO or PFOS solutions also contained ClO_4^- [$pK_a(HClO_4) < -7$] at 5-fold larger concentrations as the internal standard. HCl or NaOH was used to adjust the pH in the range of 1.0−6.0 at constant ionic strength, unless otherwise specified. Solutions were directly infused into a HP 1100 MSD ESI-MS operated in the negative ion mode.[47−49] The initial search for anion signals in the $50 \leq m/z \leq 2000$ range was performed in the scan mode. Signal intensities of $m/z = 499$ (PFOS), 413 and 369 (PFO, PFO − CO_2), 99 and 101 ($^{35}ClO_4^-$, $^{37}ClO_4^-$), and 827 [$(PFO)_2H$] peaks were quantified from mass spectra acquired in the SIM mode under the following

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10.1021/jp9051352 CCC: $40.75

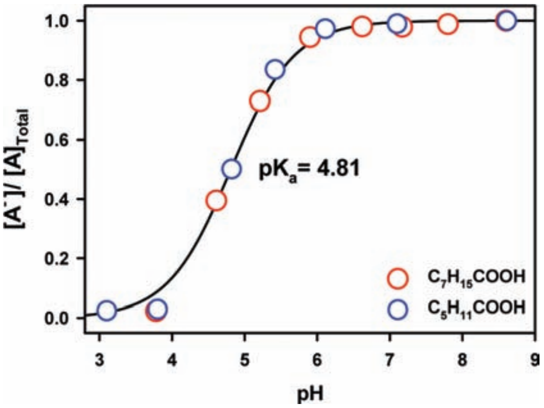


Figure 1. Titration curves of *n*-hexanoic and *n*-octanoic acids. ESI-MS signal intensities of *n*-hexanoate ($m/z = 115$, blue circles) and *n*-octanoate ($m/z = 143$, red circles) relative to ClO_4^- ($m/z = 99$, 101) as functions of solution pH. Solutions are 100 μM in $NaClO_4$ and *n*-hexanoic or *n*-octanoic acids. HCl or NaOH solutions were used to adjust the pH while keeping the total chloride concentration at 1.0 mM by NaCl addition.

conditions: drying gas flow rate = 10 L min^{-1}, drying gas temperature = 250 °C, capillary voltage = 3500 V, and fragmentor (cone) voltage FV varied from 30 to 150 V.

Results and Discussion

Given the ongoing debate about whether proton activity at the air/water interface, from which the ions detected by ESI-MS arise, is larger or smaller than that in bulk solution,[50−54] we deemed it essential to validate our procedures by reproducing the titration curves of *n*-hexanoic and *n*-octanoic acids in this setup (Figure 1). Nonlinear regressions ($R^2 = 0.995$) through the experimental data based on the universal titration function, eq 1

$$\frac{[A^-]}{[A]_T} = \frac{1}{1 + 10^{pK_a - pH}} \tag{1}$$

led to pK_a(*n*-hexanoic acid) = 4.81 ± 0.05 and pK_a(*n*-octanoic acid) = 4.81 ± 0.06 values in excellent agreement with their pK_a values in bulk solution.[55] This agreement cannot be regarded fortuitous or accidental and has important implications. Since eq 1 can be construed as a function of the difference (pK_a − pH) rather than of pH alone, the same data would have been obtained had pK_a and pH shifted equally at the interface relative to their bulk values.[56] Such coincidental shifts, however, are deemed unlikely because we cannot envision a physical reason that it should be so. More importantly, the observed agreement further implies that the output signal sets generated by our ESI mass spectrometer are linear transfer functions of the ionic composition of the interfacial layers of infused solutions. This is not a trivial observation because the detected ions are field-ejected from nanodroplets produced after extensive solvent evaporation from nascent microdroplets.[57,58] Thus, nascent microdroplets emanating from the aerial interface faithfully reflect its composition, which, as Figure 1 shows, is evidently preserved during successive solvent evaporation, microdroplet fragmentation, and ion ejection events. Since charge imbalances must persist in noninteracting microdroplets carrying anions in excess over cations, anion neutralization is prevented even in concentrated nanodroplets. Elsewhere, we have provided conclusive evidence that (1) the anion composition of the air/water interface may be quite different from that of the bulk[47,48] and (2) surfactant anion signals are linearly proportional to bulk anion submillimolar concentrations.[29] We infer that the pH of the interfacial layers sampled by our instrument is, on average, identical to that in bulk solution.

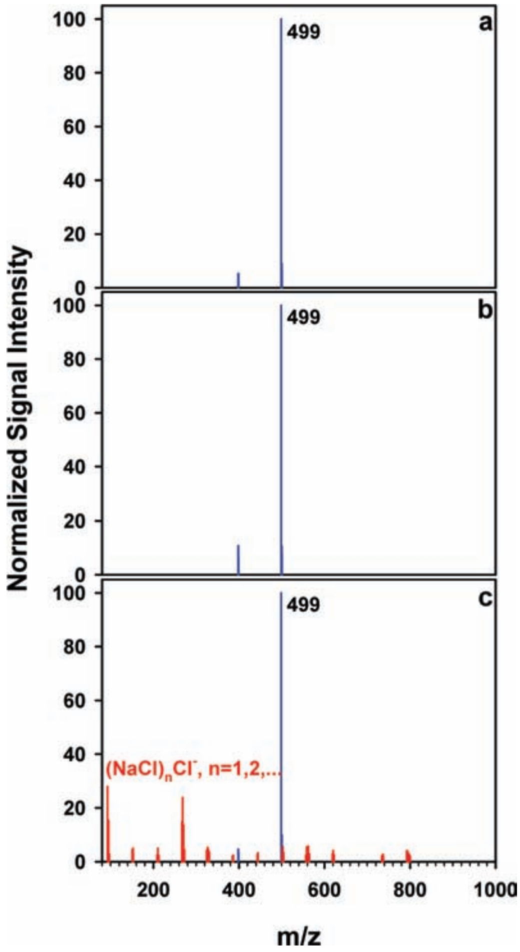


Figure 2. ESI-MS spectra of 10 μM PFOS in (a) Milli-Q water at pH 6.5, (b) 10 mM HCl at pH 2.0, and (c) 10 mM NaCl at pH 7.0. Spectra were acquired in the scan mode at a fragmentor voltage of 70 V. Maximum signal intensities ≡ 100.

Figures 2a−c show ESI-MS ($50 \le m/z \le 1000$) of 10 μM PFOS solutions in water at pH 6.5, in 10 mM HCl at pH 2.0, and in 10 mM NaCl at pH 7.0 obtained at FV = 70 V. PFOS only produces the molecular anion at $m/z = 499$ (PFOS) without evidence of a $(PFOS)_2H$ cluster at $m/z = 999$. The small signal at $m/z = 399$ is a perfluorohexanoate impurity. In contrast, ESI-MS of PFO solutions reveals the presence of a major $(PFO)_2H$ cluster anion at $m/z = 827$, in addition to the anticipated signals at $m/z = 413$ (PFO) and 369 (PFO − CO_2) (Figures 3a−c).[59] The relative intensity of the $m/z = 828$ $[^{13}C_1-(PFO)_2H]$ satellite peak confirms that $m/z = 827$ corresponds to a singly charged C_{16} species. The more extensive collisionally induced secondary dissociation of PFO at FV = 150 V (Figure 3d) leads to new signals at $m/z = 213$ ($C_3F_7CO_2^-$) and 169 ($C_3F_7^-$). Note that the $C_3F_7^-$ carbanion is a secondary species produced from $C_7F_{15}^-$ (PFO − CO_2) via a neutral C_4F_8 loss,[59] whereas $C_3F_7CO_2^-$ is a primary species ensuing from PFO by splitting C_4F_8, presumably through a higher-energy fragmentation channel. Remarkably, since we can still detect $m/z = 827$ ion signals under 150 V acceleration potentials, the $(PFO)_2H$ cluster is apparently held together by a very strong $[O{-}H\cdots O^- \leftrightarrow O^-\cdots H{-}O]$ homonuclear, three-center, four-electron covalent

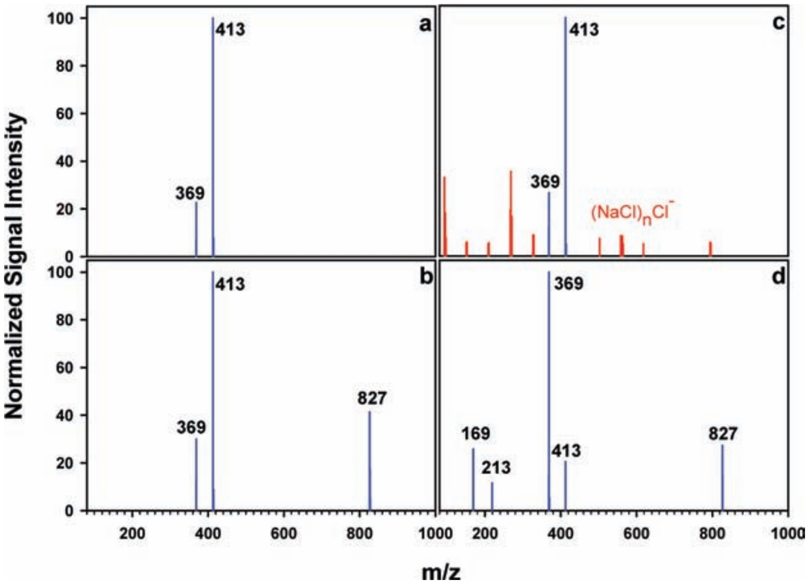
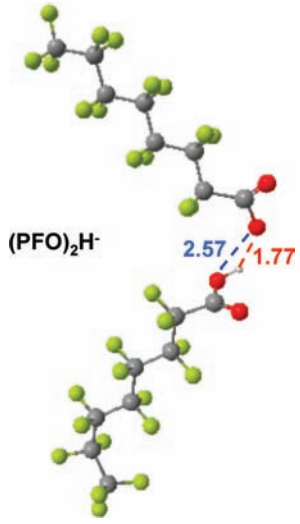


Figure 3. ESI-MS spectra of 10 μM PFOA in (a) Milli-Q water at pH 6.0, (b) 10 mM HCl at pH 2.0, (c) 10 mM NaCl at pH 7.0, and (d) 10 mM HCl at pH 2.0. Spectra were acquired in the scan mode at a fragmentor voltage set at 70 V for (a−c) and at 150 V for (d). Maximum signal intensities ≡ 100.

SCHEME 1: The MM2 Structure of the $(PFO)_2H^-$ Cluster



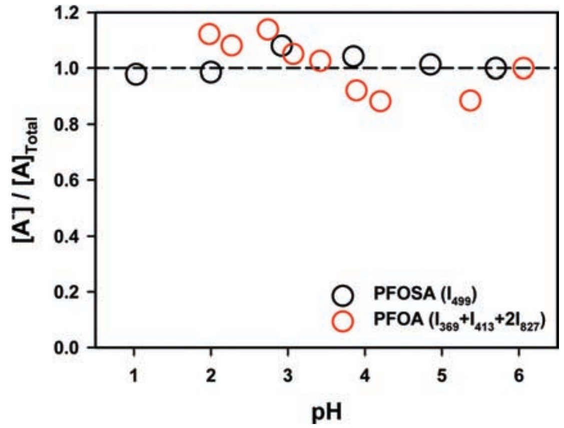


Figure 4. ESI-MS titration curves for PFOSA and PFOA. ESI-MS signal intensities from PFOS (m/z = 499, black circles) and PFOA ($I_{369} + I_{413} + 2I_{827}$, red circles) relative to ClO_4^- (m/z = 99, 101) as functions of solution pH. Solutions are 10 μM in $NaClO_4$ and PFOSA or PFOA. HCl (10 mM) and varying concentrations of NaOH were added to adjust the pH while keeping the total chloride concentration constant at 10 mM, with the exception of the solution at pH 1.

hydrogen bond whose resonant forms are rigorously equivalent (Scheme 1).[60,61] This bond is a much stronger version of those observed among most carboxylate−carboxylic acid dimers.[62,63] The detection of $(PFO)_2H$ signals in HCl, but not in NaCl solutions of identical ionic strength, and the absence of a $(PFOS)_2H$ cluster in PFOS solutions of similar concentrations suggest that clustering is not an analytical artifact under present experimental conditions.[64−66] There is no evidence for the formation of PFOS or PFOA trimeric/tetrameric aggregates under the present conditions.

Figure 4 shows that normalized PFOS (m/z = 499) signal intensities are independent of pH down to pH 1.0, confirming that PFOSA is a strong acid, that is, pK_a(PFOSA) < 1. The sum of the signal intensities of the anions derived from PFO (at FV = 70 V), $(2I_{827} + I_{413} + I_{369}) \propto [PFO]_T$, is also independent of pH, implying negligible concentrations of other species such as the undissociated PFOA acid at pH ≥ 1. Therefore, pK_a(PFOA) < 1. Figure 5 shows how the molar fraction $(2[(PFO)_2H^-]/[PFO]_T)$ varies with pH. This dependence is consistent with the clustering equilibrium, eq 2

$$2PFO + H^+ \rightleftharpoons (PFO)_2H \qquad (2)$$

$$K_2 = \frac{[PFO]_T - [PFO]}{2[H^+][PFO]^2}$$

$$[PFO] = \frac{1}{2} - \frac{1}{2}\left[\frac{-K_2^{-1} + \sqrt{K_2^{-2} + 8K_2^{-1}[PFO]_T 10^{-pH}}}{4[PFO]_T 10^{-pH}}\right] \qquad (3)$$

Nonlinear regressions to the experimental data of Figure 5 based on eq 3 and bulk concentration values yield $K_2 \sim (3.9 \pm 0.3) \times 10^7$ M^{-2}. Although many studies have shown that the

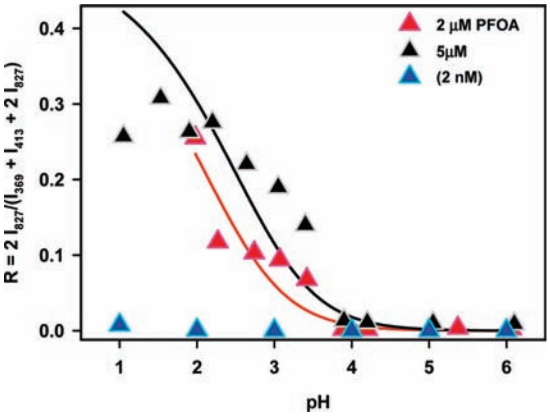


Figure 5. The ratio of the ESI-MS intensities of $(PFOA)_2H^-$ (m/z = 827) to the sum of the intensities of all PFOA species, $R = 2I_{827}/[I_{369} + I_{413} + 2I_{827}]$, as a function of pH for 2 (red triangles) and 5 μM (black triangles) PFOA solutions.

noncovalent complexes observed by electrospray mass spectrometry are not artifactual because their abundances respond to subtle molecular effects,[64−66] interfacial PFO concentrations are demonstrably larger than those in the bulk,[31] and the derived K_2 value should be strictly considered an upper limit. Thus, the calculated $2[(PFO)_2H^-]/[PFO]_T$ values (blue triangles in Figure 5) using $K_2 \sim 3.9 \times 10^7$ M^{-2} for $[PFO]_T = 2$ nM (a hard upper bound to PFOA concentrations in environmental aqueous media)[14,67,68] show that neither PFOS nor PFO will appreciably self-associate or protonate under realistic environmental conditions. [PFOSA]/[PFOS] and [PFOA]/[PFO] ratios should remain well below 10^{-7} in ocean waters at pH ∼ 8.1 but may significantly increase in marine aerosols that become acidified over polluted regions. Further work is underway.

Acknowledgment. This project was financially supported by the National Science Foundation (ATM-0714329). E.P. is grateful to the Fullbright Foundation for financial support.

Supporting Information Available: Additional data. This material is available free of charge via the Internet at http://pubs.acs.org.

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