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Quest for a Symmetric [C−F−C]$^+$ Fluoronium Ion in Solution: A Winding Path to Ultimate Success

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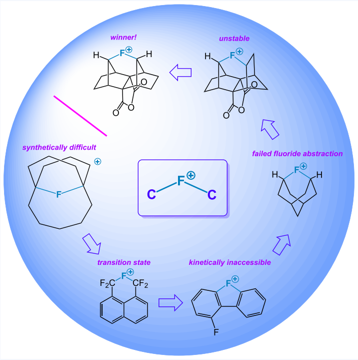


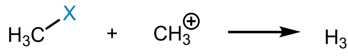
CONSPECTUS: In this Account, we chronicle our tortuous but ultimately fruitful quest to synthesize a [C−F−C]$^+$ fluoronium ion in solution, thus providing the last piece of the organic halonium ion puzzle. Inspiration for the project can be traced all the way back to the graduate career of the corresponding author, wherein the analogy between a [C−H−C]+ "hydrido" bridge and a hypothetical [C−F−C]$^+$ bridge was first noted. The earliest attempt to construct a bicyclo[5.3.3]tridecane-based fluoronium ion (based on the analogous hydrido bridged cation) proved to be synthetically difficult. A subsequent attempt involving a 1,8-substituted naphthalene ring was theoretically naive in retrospect, and it resulted in a classical benzylic carbocation instead. A biphenyl-based substrate, although computationally sound, proved to be kinetically untenable. At last, after some tweaking (including a dead-end detour into a fluoroadamantane skeleton), we finally achieved success with a highly rigid, semicage precursor based on the decahydro-1,4:5,8-dimethanonaphthalene system. This strained substrate possessed a triflate leaving group to enhance its solvolytic reactivity. Detailed isotopic labeling and kinetic studies supported the generation of a symmetrical [C−F−C]+ bridge; interesting solution behavior allowed the manipulation of the rate-determining step for solvolysis depending on solvent nucleophilicity. After initial generation as a transient intermediate, the fluoronium ion was later produced as a stable species in solution and was fully characterized by ^{19}F, ^{1}H, and ^{13}C NMR, with the resultant species displaying evident $C_s$symmetry through coordination of a molecule of SbF_5. This remarkable ion proved stable to −30 °C. We also address a disagreement surrounding the nomenclature of fluoronium ions in particular and its potential impact upon the naming of onium ions in general. We strove to highlight the dangers of confusing the arbitrary concept of calculated partial charge with IUPAC nomenclature. Finally, we discuss future directions, for example, the synthesis of a fluoronium ion in which fluorine resides within an aromatic ring.

INTRODUCTION

The placement of the last piece of a jigsaw puzzle is self-evident. On the other hand, the last piece of a science puzzle may present a much greater challenge. The story of the organic halonium ions presents such an analogy. Organic halonium ions are generally divalent species, with a formal positive charge on the central halogen, more or less equally bonded to two carbon atoms. For decades, halonium systems were well-defined for [R—I—R]$^+$, [R—Br—R]$^+$, and [R—Cl—R]$^+$, but [R—F—R]$^+$ presented an eternal challenge. It was the last "piece" of the halonium puzzle, seemingly due to the electronegativity of fluorine and its reluctance to share lone pairs of electrons with carbocationic centers. One need only look at the calculated enthalpies (*ω*b97xd/6-311++G**, Scheme 1) for the reaction of methyl halides with the methyl cation to see the trend; while still exothermic, it is quite a bit less favorable for CH_3F than for CH_3Cl and CH_3Br.

Significant work has been done regarding characterization of other halonium ions; stable iodonium[1] and bromonium[2] ions have long been known, and more recently chloronium ions have been of timely interest.[3] Additionally, stable species involving oxonium ions[4] and divalent, bridging hydrogen atoms have been synthesized.[5] Although we drew inspiration from these established systems when planning our work, fluorine's extreme electronegative nature poised it to be the most difficult atom to make to achieve bridging between two carbon atoms and a formal positive charge.

Scheme 1. Calculated Enthalpies of Halonium Ion Formation



X = F (ΔE = -50.1 kcal)
X = Cl (ΔE = -66.8 kcal)
X = Br (ΔE = -72.1 kcal)

Nevertheless, some seminal works chipped away at the problem. Morton's evidence for a fluoronium ion **1** in a gas phase mass spectrometry experiment is notable,[6] as is Dopfer's F-protonated fluorobenzene **2**.[7] Gabbai et al. synthesized a rapidly equilibrating species **3** in which fluorine was polarized by the adjacent cationic center.[8] Additionally, Müller and co-workers reported a silicon analogue (**4**) of a fluoronium (Scheme 2).[9] Over the decades, these exceptional reports seemed to put the observation of a true, symmetrical [C−F−C]$^+$ fluoronium ion in solution within reach.

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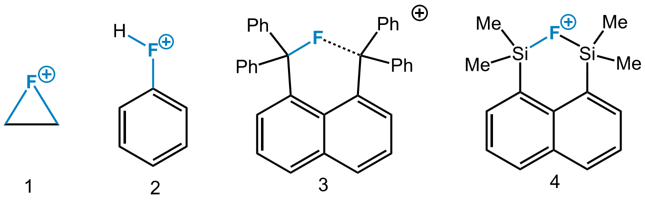
Scheme 2. Previous research on fluoronium ions



To put the corresponding author's personal journey in perspective, consider the concluding statement/figure of an *Accounts of Chemical Resarch* paper by McMurry and Lectka on the topic of [C−H−C] three-center, two-electron bonds published 27 years ago in 1992:[10] "What lies in the future for work in this field? Certainly it is possible to imagine other kinds of 3c−2e bonds that might be stabilized... Among the possibilities, stable C−F−C bonds... are well worth reaching for. We have already identified and carried out preliminary calculations on systems that look attainable... All that is needed now is to reduce the idea to practice." In fact, a C−F−C fluoronium ion (actually a 3c−4e system!) was also proposed at that time (Figure 1). Though this particular ion proved synthetically unfeasible at the start of the corresponding author's independent career, the concluding, light-hearted, and simple "all that is needed now" served as ironic foreshadowing for the 20+ years of trials to come before ultimate success in obtaining either indirect (2013) or direct (2018) evidence for a $[C-F-C]^+$ fluoronium ion in solution.

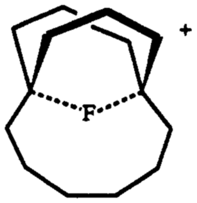
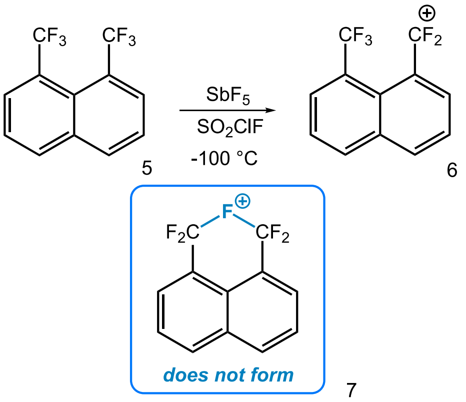


Figure 1. Fluoronium ion structure proposed 27 years prior during the corresponding author's graduate studies, reproduced as it appeared in a contemporaneous *Accounts of Chemical Research* paper. Reprinted with permission from ref 10. Copyright 1992 American Chemical Society.

The earliest attempt at making a $[C-F-C]^+$ fluoronium ion in our laboratory proved to be a naïve exercise (Scheme 3). We imagined that treatment of 1,8-bis(trifluoromethyl)-naphthalene **5** with a strong Lewis acid (such as SbF_5) could abstract a fluoride ion to form the desired symmetrical structure. Instead, the *asymmetrical* structure **6** was formed instead of symmetrical **7**. Rapid equilibration was not observed, as decomposition occurred at temperatures above −80 °C. In retrospect, we should have known better, as ex post facto DFT and MP2 calculations confirmed the asymmetric structure as a minimum, whereas the fluoronium form computed as a transition state.

Scheme 3. First Attempt at Fluoronium Formation



Yet another unsuccessful attempt followed a few years later.[11] In this case, we strove to generate the fluoronium as a reactive intermediate from the corresponding salt **11** through dediazotization, in a similar fashion to the classic halonium ions generated by Sandin and co-workers (Scheme 4).[2b] If successful, a mixture of isomeric products should arise after trapping. We also imagined that the resulting fluoronium would be stabilized to an extent by a potentially aromatic benzene-fused fluoracyclopentadiene. Unfortunately, under all circumstances, only products formed from trapping at the site of dediazotization were observed. Part of the problem was the choice of solvent: effective, nonradical dediazotization only occurred in fairly polar solvents, which in turn proved to be fairly nucleophilic and thus avid trapping agents.

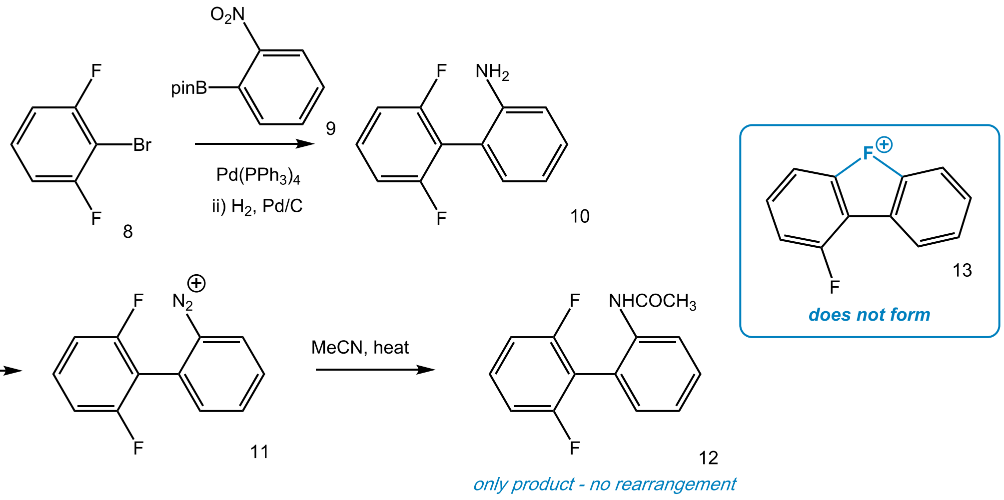
We finally realized that, for a successful attempt at a fluoronium generation, several criteria had to be met. Of course, the target species should be a calculated local minimum at a high level of theory. The second criterion mandated the use of a "cage" structure to keep the fluorine atom locked in, preventing undesired decomposition pathways (Scheme 5). There have been elegant examples in the literature of cages containing neutral fluorine as part of a C−F bond[12] and fluoride ion,[13] but a "fluoronium cage" would present significant challenges. Several cages are shown below, of which we first chose **17** as being the most synthetically reasonable; we would also later attempt to form **16**.

INDIRECT GENERATION OF A FLUORONIUM

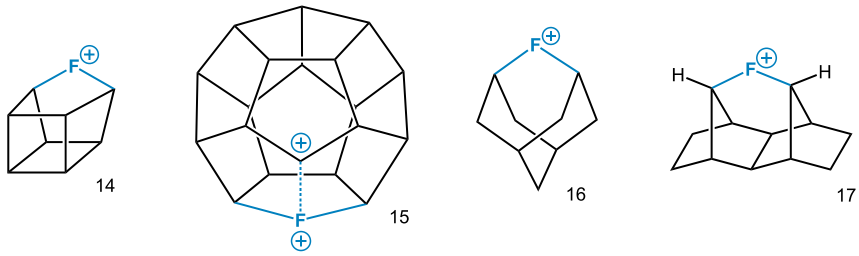
For synthetic efficacy, we imagined a precursor such as **18** (Scheme 6). Although the anhydride group would appear to be nonoptimal for Lewis acid ionization, its presence would allow the precursor to be assembled through a series of Diels−Alder reactions using a dienophile such as **20**. Our first attempt at making semicage system **18** involved the reaction of dienophile **20** with cyclopentadiene at 180 °C in a sealed tube (Scheme 7). Attempts to functionalize diastereomer **21** selectively failed, so we turned our attention to **22**, which was formed in comparable quantities. We reckoned that protonation of **22** with Magic acid[14] would result in the slightly asymmetrical fluoronium ion **23**. Unfortunately, only decomposition of **22** with attendant gas evolution (CO, CO_2) resulted.

Thus, we decided to try different methods of making a substituted version of **21**, this time using dienophile **28** and a silacyclopentadiene[15] (the silane would serve as a hydroxy group equivalent). Computations strongly indicated that such cycloadditions would afford the correct diastereomer, along with an equal amount of an undesired isomer.[16] At B3LYP/6-311++G** (Figure 2), the two most favorable diastereomeric transition states lie close together in energy and the desired isomer **24** is only 0.76 kcal higher in energy than **25**. The other two possible isomers, **26** and **27**, are unlikely to form, as the calculated transition states are considerably higher in energy (4.3 and 4.7 kcal). At the very least, this calculation led us to believe that the Diels−Alder reaction would have an excellent

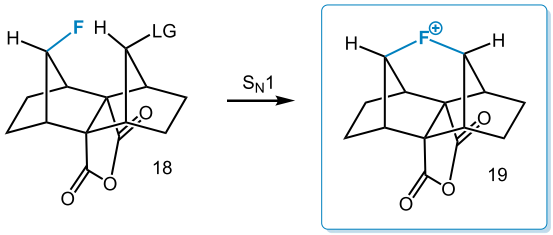
Scheme 4. Second Attempt at Forming a Fluoronium Ion through Dediazotization



Scheme 5. Possible Fluoronium Cage Structures



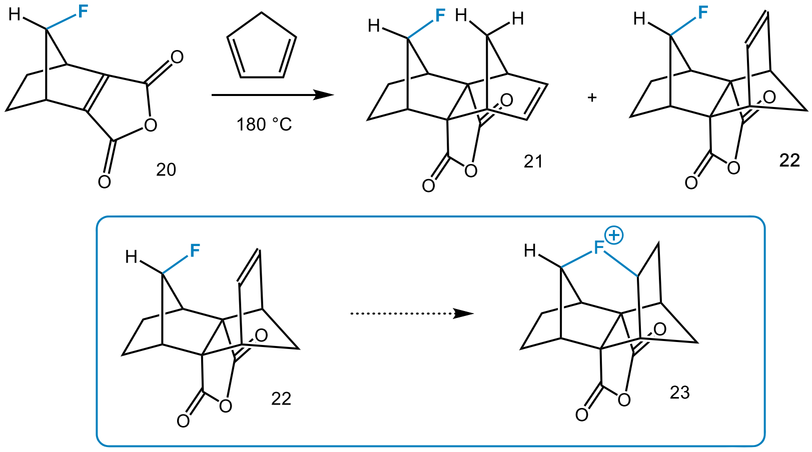
Scheme 6. Settled Strategy to the Fluoronium Ion



chance of producing some of the desired diastereomer. Alarmingly, all attempts at a thermal or Lewis acid catalyzed Diels–Alder reaction using diene **29** and dienophile **28** utterly failed. The diene, it appeared, was very susceptible to rearrangement[17] and decomposition, a finding not surprising in light of the migratory aptitude of silicon groups.

At this point, we tried a reaction under high pressure; after a few days at 12 kbar (several times the pressure at the bottom of the Mariana Trench!),[18] the crude ^{19}F NMR spectrum showed the presence of two major diastereomers whose structures we presumed to be those favored by calculation (Scheme 8). Based on the prediction of ^{19}F chemical shifts, the desired diastereomer **30** was produced in a slightly lower yield than **31**, as expected. The crude mixture was subjected to Fleming–Tamao oxidation; separation of the alcohols by column chromatography on silica gel was followed by hydrogenation

Scheme 7. Synthesizing the Sesquinorbonyl Cage



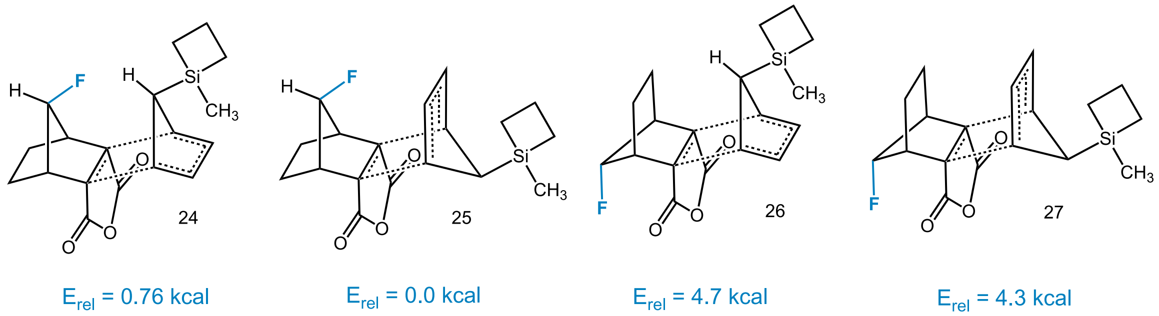
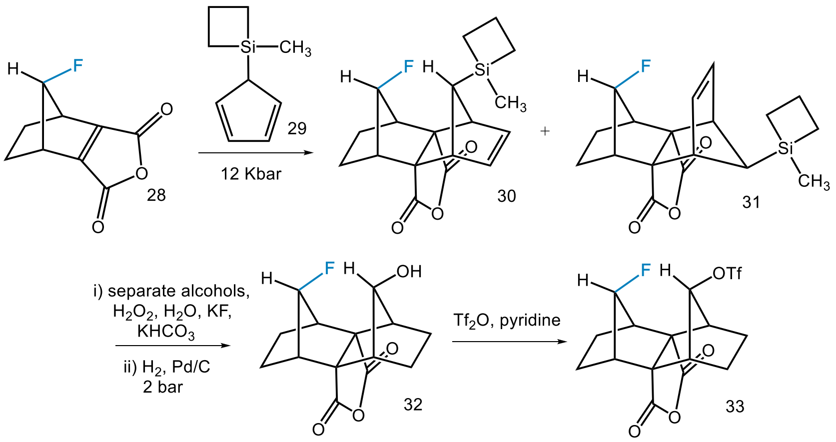
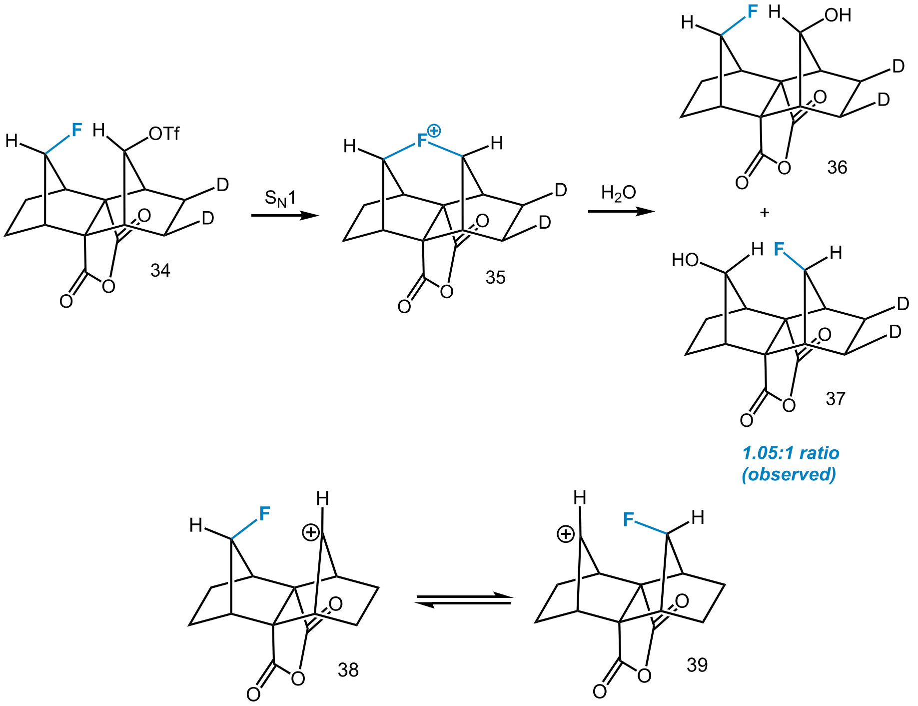


Figure 2. Calculated relative energies of Diels−Alder transition states for possible diastereomers.

Scheme 8. Synthesis of Triflate 33

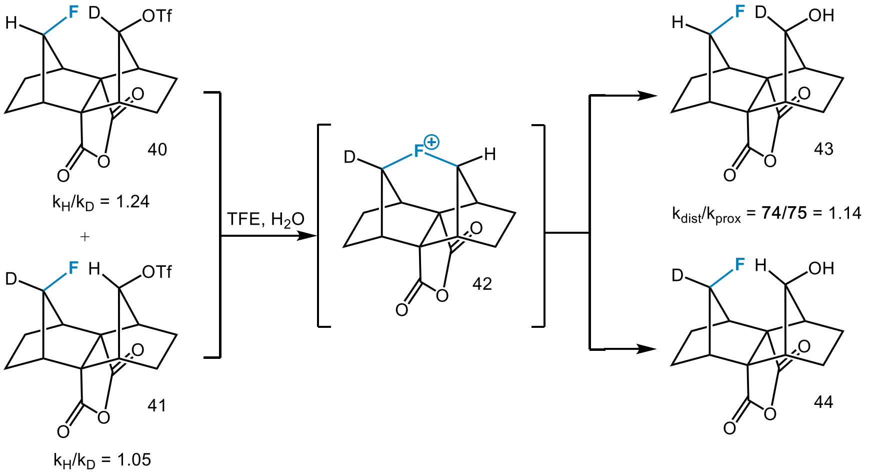


Scheme 9. Isotopic Labeling Experiments



and triflation. We selected a triflate as the leaving group, as it is good for the generation of reactive cationic intermediates.[19] Triflate **33** is an interesting species in its own right; the fluorine atom and the *in*-H engage in a "jousting" interaction,[20] in which the effects of bond compression are modulated by hydrogen bonding. For example, the through-space $^1H-^{19}F$ coupling is 25 Hz, implying a large interaction. In addition, the C-*in*-H stretch is seen in the IR spectrum at 3147 cm^{-1}, blue-shifted from a control by 39 cm^{-1}. The degree of s-character of the C-*in*-H bond is calculated to be a remarkable 29%.[21] We

Scheme 10. Isotopic Labeling Experiments Demonstrating α-Secondary and Remote Effects



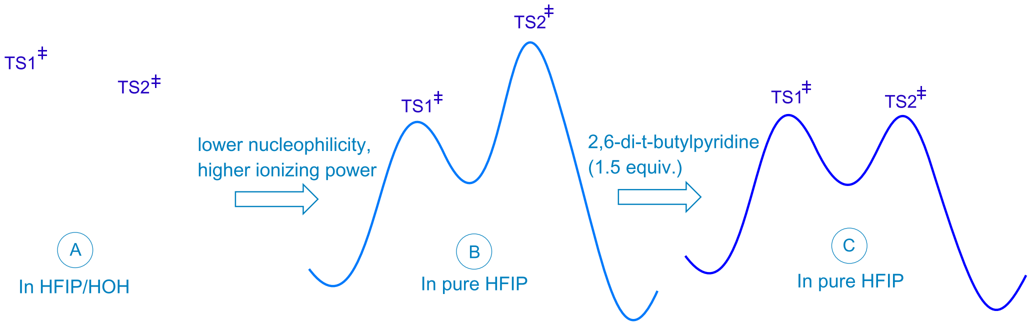


Figure 3. Different solvents and additives can change the rate-determining step of solvolysis.

obtained a crystal structure of triflate **33**, revealing the close interaction of F and the *in*-H at its estimated position.[22] With that in mind, we surmised that hydrolysis of **33** would release a fair amount of strain energy.

We chose to first generate the fluoronium as a reactive intermediate for several reasons. First of all, it was a simpler approach—as the old adage goes, you have to walk before you can run. Second, we felt that direct observation would be difficult due to the potential instability of the anhydride functional group, which is necessary for the synthesis of the molecule and could not be removed. Nevertheless, the hydrolysis of **33** was challenging in its own right. We found that an optimal system consisted of 2,2,2-trifluoroethanol (TFE)[23] and water. In 70/30 v/v TFE/water, **33** hydrolyzed to afford **32** in almost quantitative conversion (98%).[24] This indicated to us that we had a remarkably stable system with no apparent propensity to rearrange its fused norbornane skeleton. However, the identity of the product provided little useful information about the mechanism of the substitution reaction. We turned to an isotopic labeling experiment that permitted differentiation between various mechanistic alternatives; scrambling between positions is a classic indication of a cloaked symmetrization process. When labeled substrate **34** is subjected to hydrolysis in TFE/water 70/30 v/v, we obtain a (1.05/1) mixture of isotopomeric products **36** and **37**, and the situation is now greatly clarified (Scheme 9). Were the reaction proceeding through an S_N2 or S_Ni mechanism, a 1:1 ratio would not be expected. This ratio is the likely result of comparable trapping on both electrophilic sites of a symmetrical intermediate, taking into account a steric isotope effect. Although highly suggestive, the case is arguably circumstantial.

We also examined isotope effects more proximate to the putative fluoronium interaction.[25] Two significant kinetic isotope effects were measured. The first, representing the rate of consumption of labeled isomer **40** (geminal D) versus unlabeled triflate, was found to be $k_H/k_D = 1.24$, a significant α-secondary effect (Scheme 10). On the other hand, $k_H/k_D =$ 1.05 for the position α- to the fluorine atom (removed through space from the position of the triflate group). The fact that a measurable *remote* kinetic isotope effect (KIE)[26] was observed implies that fluorine is participating in the ionization. On the other hand, the remote effect may be at least partly attributable to the relief of steric strain upon ionization, although various computational models indicate participation of fluorine.

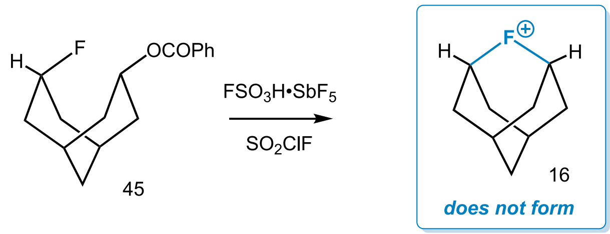
Predicted isotope effects were derived from analysis of transition states and computed to be 1.32 and 1.07, in line with what we found. Finally, KIEs for the trapping of water to form **43** and **44** were found once again to be 1.14 ± 0.03. These results are very important from a confirmatory standpoint: given stereochemically unique substrates (both *out*- and *in*-triflate isomers), we observe the exact same outcome, thus implying the exact same intermediate. Regardless of the starting ratio, the presence of a fluoronium intermediate demands an identical ratio of label isomers, which is what we observe.

DIRECT OBSERVATION OF A [C–F–C]$^+$ FLUORONIUM

Once we had established to our satisfaction that the [C–F–C]$^+$ fluoronium could be generated as a transient intermediate, our thoughts turned to synthesis of the fluoronium as an observable species in solution. Of course, direct observation provides the most convincing evidence as well. Still, we were at first doubtful that we could generate a fluoronium ion as a stable intermediate in superacid media. However, several propitious observations during our study of the solvolysis of **33** in different media began to change our minds. Most importantly, we established that the rate-determining step of the solvolysis of triflate **33** could be switched through the use of a sterically bulky polyfluorinated alcohol solvent. Pure hexafluoroisopropanol (HFIP) possesses increased ionizing power over that of aqueous mixtures, yet decreased nucleophilicity. Thus, in pure solvent media, the lifetime of the ion increases (Figure 3). Addition of a hindered base (2,6-di-*t*-butylpyridine) enhances the nucleophilicity of HFIP and consequently reduces the lifetime of **19**. These results suggested that in superacid media, a fluoronium could be observable.

The corresponding author was skeptical that it could be done with a precursor such as **33**, given the (presumed) susceptibility of the anhydride group to degradation by strong Lewis acids. At this point, we sought to make a reasonable precursor unencumbered by the anhydride group. The bicyclic desmethylene adamantane **45** seemed to be a viable precursor to the nifty fluoraadamantane skeleton (Scheme 11).

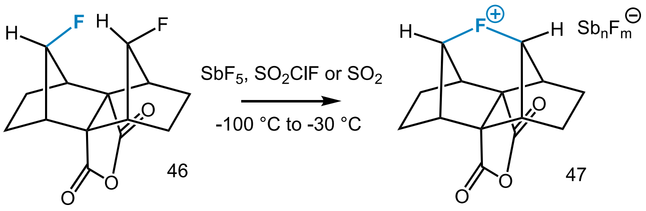
Scheme 11. Precursor Based on a Different Cage



Unfortunately, the synthesis of **45** proved to be problematic (though possible to synthesize, it was even more difficult to produce appreciable amounts of it than **33**), and its behavior in superacid media was less than ideal. Though it is unclear exactly what the fate of the bicyclononyl compound was when treated with Magic acid, a key observation was the disappearance of the ^{19}F NMR signal. We attributed this to defluorination by the very fluorophilic SbF$_5$ component of Magic acid. This led us to conclude that the bicyclononyl system, and possibly the project itself, was likely at a dead end.

However, the observation of defluorination by SbF$_5$ ultimately led us in the correct direction. Coauthors Holl and Pitts, of this publication, fortunately insisted upon revisiting precursor **33**, albeit in a slightly different form (despite the corresponding author's skepticism). Considering defluorination of organic fluorides by SbF$_5$ occurs readily, if we were to use a compound that contains two fluorine atoms, with one facing inward and another on the opposite bridge facing outward, then defluorination by SbF$_5$ of the *out*-F would give the fluoronium ion. Though two fluorine atoms are present, the *out* fluoride is more sterically accessible (and engages in a weaker bond to C), so it should be preferentially abstracted over the *in* fluoride, whereas a second defluorination of the resulting fluoronium is extremely unlikely. Additionally, since the cause of the issue with the sesquinorbonyl system was believed to be decomposition of the anhydride by the Brønsted component of Magic acid (FSO$_3$H), using the Lewis acid SbF$_5$ alone would be both sufficient to remove the fluorine and likely more compatible with the anhydride group. We also anticipated that synthesis of an equivalent bicyclononyl difluoride would be prohibitively difficult (indeed, later attempts to do so were not successful) and its capability of inversion to boatlike conformers would decrease the overall steric hindrance of the *in*-F, leaving the *wrong* fluorine atom vulnerable to abstraction. Thus, we optimized the synthesis of **46** and attempted stable fluoronium ion formation therefrom (Scheme 12). Upon treatment of **46** with SbF$_5$ in SO$_2$ClF, we observed the formation of a new species in the ^{19}F NMR that displayed a single fluorine resonance with what appeared to be an overall triplet splitting, as we predicted from the equivalent geminal hydrogen atoms. While this result was extremely promising, the spectrometers at Johns Hopkins University were not suitable for further detailed studies, so we availed ourselves of the help of Prof. Antonio Togni and the NMR facilities at ETH Zürich. Such quality NMR data were obtained of the new species that a reviewer later assumed that the spectrum was simulated! This new species was thermally stable up to ca. −30 °C. The ^{19}F NMR signal appears as a triplet of pentets, consistent with strong geminal coupling and remote "W-coupling." The ^{1}H NMR spectrum shows a highly diagnostic downfield-shifted doublet (2H, 8.11 ppm) representing the expected reciprocal geminal coupling to ^{19}F.[27]

Scheme 12. Fluoride Abstraction Generates a Stable Fluoronium Ion



On the other hand, we were alarmed upon seeing the ^{13}C NMR spectrum. Instead of the expected C_{2v} symmetry of the desired ion, the observed species revealed extra signals consistent with overall C_s symmetry. A clue was provided in the chemical shifts of the carbonyl groups, one of which rests considerably downfield from the other (159.5 vs 186.7 ppm). We concluded this shift to be a result of SbF$_5$ coordination to one of the anhydride carbonyls in the cationic structure (Figure 4). Calculation of the ^{13}C NMR spectrum of complex **48** (ωb97xd/6-311++G**, C_s symmetry) predicted chemical shifts consistent with experiment.

Fortunately, coordination still permits a *fully symmetric* C–F–C interaction, as while the C_{2v} symmetry is broken by SbF$_5$, the remaining mirror plane bisects the fluorine atom and leaves the bound carbon atoms as symmetry partners. One curiosity still intrigued us: why was the ^{19}F spectrum consistent with C_{2v} symmetry? Calculation of the W-coupling provides an answer: it is predicted to be 13 Hz for all four hydrogens. Thus, an apparent triplet of pentets is expected and observed (Figure 4).

We still needed to answer definitively the most important question: is the fluoronium structure *truly* symmetrical, with

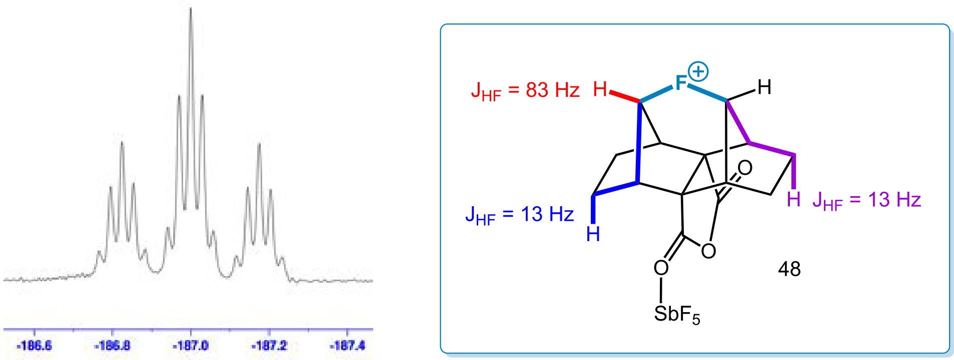
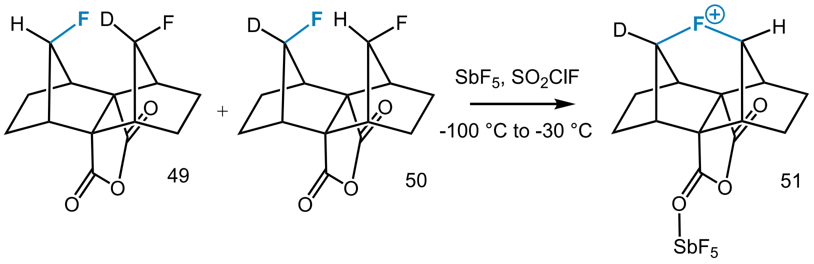


Figure 4. ^{19}F NMR spectrum of fluoronium ion **48**.

Scheme 13. Generation of the Fluoronium Ion with Labeled Precursors



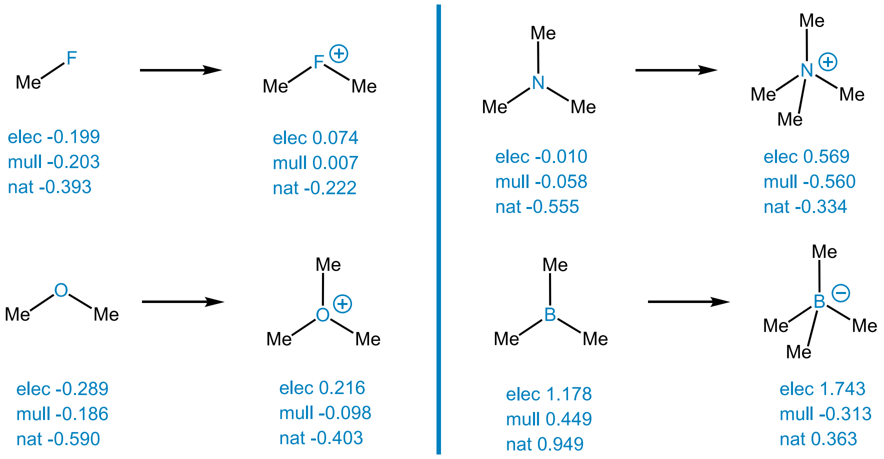


Figure 5. DFT calculated charges of select onium and -ate ions (ωb97xd/6-311++G**).

fluorine bound equivalently to both carbon atoms? The NMR data alone are fairly indicative, but substantial claims require substantial evidence. One notable approach to the classical versus nonclassical ion question lies in the Schleyer–Lenoir–Prakash–Olah chemical shift additivity test.[28] The authors write "the total ^{13}C chemical shift difference between a carbocation and the corresponding neutral hydrocarbon also provides a rough, but useful, structural index. Classical carbocations show large chemical-shift differences of 350 ppm or more, whereas nonclassical cations display differences often hundreds of ppm less." We measured the chemical shift additivity of system **48** as 111.8 ppm, which is well within the nonclassical realm (Scheme 13). Factoring in the effect of SbF_5 coordination, this value presents an upper limit on the difference that would be expected in the uncoordinated cation.

We next turned to the venerable Saunders isotopic perturbation test for the final word.[29] For an isotopically labeled symmetrical structure, the difference in chemical shift of the bridging carbons should be on the order of 1 ppm or less. On the other hand, a labeled equilibrating structure should display a much larger chemical shift difference (>10 ppm). Upon ionization with SbF_5, a mixture of two deuterium labeled precursors leads to the identical cation (Scheme 13). The difference in chemical shifts between the two carbons is 0.08 ppm, definitive of a nonclassical cation structure.

■ NOMENCLATURE

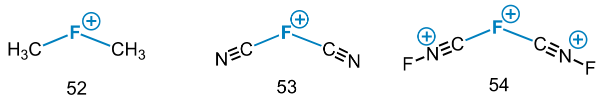
When we published our work in the area, we were unaware that a few members of the inorganic community objected to the use of the "fluoronium" nomenclature on the basis that fluorine itself could never really be positive in such species.[30] These authors suggested the term "fluorine atoms with diminished negative charge relative to the free fluoride ion" to describe the $[C-F-C]^+$ fluoronium, which of course would be true for pretty much all fluoroorganics, or virtually any fluorine compound for that matter. At the time of our first paper, the organic fluoronium nomenclature had already been well-established in the literature with neither apparent comment nor controversy.[6−8]

The critics' arguments were based on the calculation of partial negative charge at fluorine using the natural, electrostatic, and Mulliken protocols. We had in fact pointed out well before these authors that calculated partial charges in fluoronium **19** were negative at fluorine using these counting procedures, but deferred of course to precedent and standard IUPAC nomenclature on halonium ions,[31] in which calculated charges at the central atoms are irrelevant.

Upon reflection, it is easy to conclude that the edifice of IUPAC onium ion naming and the concept of formal charge in valence bond theory would crumble if such advice were followed. For example, in the expanded valence species of Figure 5, the central atom generally calculates as negative using natural charges (ωb97xd/6-311++G**),[32] whereas valence bond theory assigns a formal charge that is positive. Likewise, the "ate" nomenclature should topple as well, as the central atoms may calculate as positive, not negative. In fact, one critic of the term "fluoronium" has himself used the onium nomenclature in a coauthored study, thus adding to the inconsistencies.[33] Additionally, the calculated atomic charge values are in a sense subjective and depend on the method used to calculate them, so basing nomenclature thereupon would often afford erratic and contradictory results.

In the case of onium or ate ions, it is clearer to look at *trends* rather than absolute values. In each of the examples of Scheme 14, the calculated natural charges move in a positive (onium) or negative (ate) direction upon ion formation. On the other hand, the absolute calculated charges are often a confusing assemblage of positive and negative values. One must also keep in mind that the reactivity of the fluoronium ion is fully consistent with its valence bond description.

Scheme 14[a]



52 53 54

[a] Left: bridging fluorine atom in **52** calculates as negative (natural); right: bridging fluorine atom in **54** calculates as virtually neutral.

By manipulating a few substituents on the fluoronium, one can make the central F move in the positive direction (**52**−**54**, charges calculated at ωb97xd/6-311++G**). The actual C−F−C bond remains fairly unchanged. In **54**, it is the terminal fluorine atoms that calculate as partially positive! Granted, fluoroniums **52**−**54** are hypothetical "theorists' dreams," but we present this as a thought exercise to illustrate why it is hazardous to have calculated charges dictate nomenclature.

Lastly, these same authors also took exception to the use of the term "hypervalent" for fluoronium species, which was a minor statement we made that was subsequently overemphasized in the secondary literature.[34] Divalent fluoronium ions seemingly fit the original definition of "hypervalent molecules and ions" by Musher,[35] with fluorine being a group 17 element having a valence other than its lowest stable valence of 1, but this definition is somewhat archaic, and the critics have a point. From our perspective, there exists no clear and consistent definition of "hypervalent" in contemporary literature; it has been used in various contexts over the years, from the vague to the precise and narrow.[36] Under some definitions, involvement of d-orbitals is required for a molecule to be considered hypervalent, but in others, "the availability... of d orbitals has nothing to do with this state of affairs."[37] Additionally, studies have challenged the use of the term for species long considered to be so, such as the sulfate ion,[38] and other definitions propose using the term for species that do not fit traditional definitions, such as the nitrate ion.[39] On the other hand, IUPAC has weighed in on the issue to limit the term to species possessing expanded valences beyond the Lewis octet rule.[40] Under this definition, halonium ions are generally not hypervalent species. However, the significance of the octet rule to hypervalency, and the meaningfulness of the term itself, has been questioned.[41] In any case, this minor issue arises perhaps from differences in perception between chemists.

■ ELECTRONIC PROPERTIES OF THE $[C-F-C]^+$ FLUORONIUM

The C−F−C array is a three-center bond, and its orbital properties are similar to those of C−Cl−C, C−Br−C, and the isoelectronic ethers. The lowest energy orbital is the array shown in Figure 6, which is similar to those calculated for the aforementioned isoelectronic structures. Thus, there is nothing exotic about the C−F−C interaction to distinguish itself from other onium ions. The C−F−C interaction is calculated to be bent, which is consistent with an approximate sp^3 hybridization model at fluorine. Various levels of theory (DFT and MP2) are in basic agreement on the structure of the cation. The average C−F bond length is 1.58 Å among the methods chosen; the C−F−C bond angle averages at about 115°. An atoms-in-molecules[42] analysis (ωb97xd/6-311++G**) shows that each C−F interaction constitutes a little less than half of a normal C−F covalent bond (electron density at bond critical point = 0.14), and the Wiberg bond order for each of the C−F bonds was calculated to be ca. 0.5.

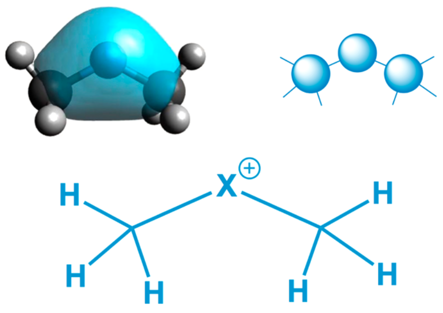


Figure 6. Three-centered bonding array of the fluoronium ion.

■ FUTURE DIRECTIONS

We have already applied the general reactivity concept of the fluoronium to some basic reaction chemistry, such as arene activation.[43] We are also intrigued by the possibility of an "acyl

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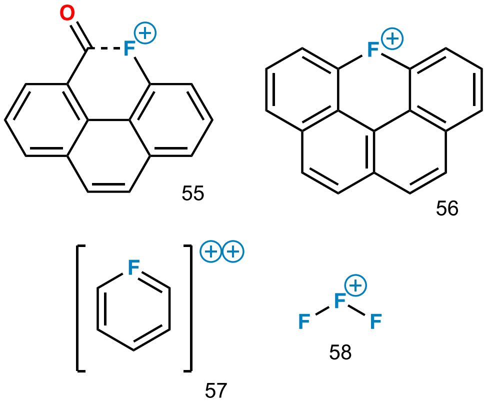
fluoronium" (**55**) that could be a key intermediate in amide hydrolysis, for example (Scheme 15). It is also possible that anchoring a C−F−C array around a benzo[*c*]phenanthrene bay could finally provide the long sought after aromatic fluoronium (**56**). What about other remarkable fluoronium type species that are a bit more fantastic? One is a dicationic fluorabenzene (**57**) that may ultimately be of pure theoretical interest. The trifluorine cation, on the other hand, could be experimentally feasible. All that is needed now is to reduce these ideas to practice!

Scheme 15. Possible Future Directions



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The authors declare no competing financial interest.

Biographies

Maxwell Gargiulo Holl was born in 1992 in New Haven, CT. He obtained his B.A. from Johns Hopkins University in 2014 and completed his Ph.D. research with Prof. Thomas Lectka in 2018. In addition to research on fluoronium ions, he has studied through-space activation effects on arenes by fluorine and other functional groups.

Cody Ross Pitts was born in 1988 and spent the early years of his life in Waterbury, CT. He obtained a B.S. in chemistry (with minors in physics and musical theatre) from Monmouth University (2010), completed his Ph.D. research at Johns Hopkins University with Prof. Thomas Lectka thereafter (2011−2017), and most recently conducted research abroad as an ETH Zürich postdoctoral fellow in the laboratory of Prof. Antonio Togni (2017−2019). Currently, he is an NIH postdoctoral fellow in the laboratory of Prof. Phil S. Baran at the Scripps Research Institute. His research in synthetic and physical organic chemistry (beyond the fluoronium ions) has also included the development of radical based C−H and C−C fluorination methods, mechanistic studies thereof, and work on mild oxidative polyfluorination of heteroatoms.

Thomas Lectka is a native of Detroit and a graduate of Oberlin College (1985). After Ph.D. studies with John McMurry at Cornell (1985−1990), he was an Alexander von Humboldt postdoctoral fellow at Heidelberg (1991) and an NIH postdoctoral fellow at Harvard (1992−1994). Since 1994, he has been a faculty in the Department of Chemistry at Johns Hopkins, where he is now the Jean and Norman Scowe Professor of Chemistry. His interests have spanned a diverse set of topics in physical and synthetic organic chemistry including the catalysis of amide isomerization, asymmetric catalysis, catalyzed and promoted site-selective aliphatic fluorination, and the study of close interactions of C−F bonds with other functional groups.

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