

Energetics of Fluoroalkene Double Bond Isomerizations

Patrick E. Lindner*¹ and David M. Lemal*

Contribution from the Department of Chemistry, Dartmouth College, Hanover, New Hampshire 03755

Received November 1, 1996[⊗]

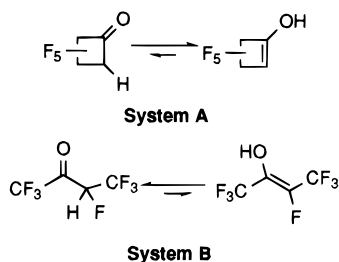
Abstract: The effects of perfluorination of cyclic and acyclic alkenes have been examined. Quantum mechanical calculations and experimental data show that vicinally substituted *acyclic* alkenes are strongly destabilized by fluorination. Quite surprisingly, the destabilization is much smaller for *cyclic* analogs. This difference has dramatic consequences for the relative stabilities of double bond isomers in cyclic and acyclic systems, including keto–enol systems. The reasons for the contrast are discussed in detail.

Introduction

Fluorine substitution at double bonds has produced many surprising results. For example, thermochemical and computational data indicate that a single fluorine stabilizes a double bond by about the same amount as a methyl group (ca. 2–3 kcal/mol), but that multiple fluorination is strongly destabilizing.² Another paradoxical finding is the greater thermodynamic stability of the *cis* isomer of 1,2-difluoroethylene over the *trans*.³ This is, of course, contrary to expectation based on simple steric or electrostatic considerations and has been termed the “*cis* effect”, for which several rationalizations have been set forth.⁴

Despite the attention that has been paid to the fluorine atom, the effects of perfluoroalkyl substitution at double bonds have remained relatively unexplored. The only perfluorinated olefins that have known structures are perfluorobutadiene,⁵ perfluorocyclobutene,⁶ and perfluoropropene,⁷ and that of the last is only partially solved. Thermochemically, it is apparent that a single CF₃ group modestly destabilizes² ethylene, but the consequences of further CF₃ substitution are not known. Furthermore, there are no data available on the thermochemical effects of fluorine substitution on simple cyclic alkenes.

The previous paper in this Journal described an unexpected contrast in the enol stabilities in highly fluorinated cyclic and acyclic keto–enol pairs. We demonstrated that in system **A** the enol is much more stable thermodynamically than its ketone, but that the opposite is true in system **B**.



[⊗] Abstract published in *Advance ACS Abstracts*, April 1, 1997.

(1) Walter H. Stockmayer Fellow, 1995–96.

(2) For a review of this subject, see: Smart, B. E. In *Molecular Structure and Energetics*; Liebman, J. F., Greenberg, A., Eds.; VCH Publishers: Deerfield Beach, FL, 1986; Vol. 3, p 141.

(3) Craig, N. C.; Piper, L. G.; Wheeler, V. L. *J. Phys. Chem.* **1971**, *75*, 1453.

(4) (a) Viehe, H. G. *Chem. Ber.* **1963**, *93*, 953. (b) Pitzer, K. S.; Hollenberg, J. L. *J. Am. Chem. Soc.* **1976**, *98*, 461. (c) Bingham, R. C. *J. Am. Chem. Soc.* **1976**, *98*, 535. (d) Epiotis, N. D.; Cherry, W. R.; Shaik, S.; Yates, R. L.; Bernardi, F. *Top. Curr. Chem.* **1977**, *70*, 1. (e) Hollein, H. C.; Snyder, N. H. *J. Mol. Struct.* **1982**, *84*, 83. (f) Wiberg, K. B.; Murcko, M. A.; Laidig, K. E.; McDougall, P. J. *J. Phys. Chem.* **1990**, *94*, 6956.

(5) Chang, C. H.; Andreassan, A. L.; Bauer, S. H. *J. Org. Chem.* **1971**, *36*, 920.

(6) Chang, C. H.; Porter, R. F.; Bauer, S. H. *J. Mol. Struct.* **1971**, *7*, 89.

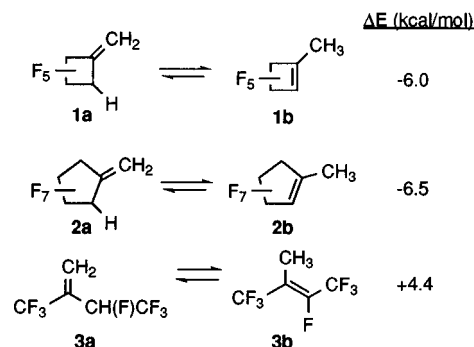


Figure 1. Ab initio energy differences for fluoroalkene isomerizations.

Quantum mechanical calculations indicated that destabilization of acyclic relative to cyclic enols by fluorine substitution is responsible for this contrast.

It seemed possible that this finding is not unique to keto–enol systems, but is rather a reflection of fundamental differences in the effects of fluorination on cyclic and acyclic alkenes. We have now tested this hypothesis both computationally and experimentally.

Results and Discussion

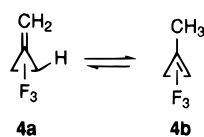
The Cyclic–Acyclic Contrast. Does the oxygen atom in the highly fluorinated keto–enol systems play an important role in determining the positions of the keto–enol equilibria? This question was addressed by replacing the oxygen atoms in the enols and ketones with methylene groups as shown in Figure 1. The energy differences were calculated at the HF/6-31G**/6-31G** level of theory.

In all of the alkene isomerizations the forward reaction is favored far more than in their ketone-to-enol counterparts, a reflection of the greater strength of the carbonyl double bond. The $\Delta\Delta E$ of 10.4 kcal/mol for system **1** versus system **3** agrees closely, though, with that calculated for the corresponding keto–enol pairs (11.1 kcal/mol for system **A** versus **B**). Agreement is equally good for system **2** versus **3**: $\Delta\Delta E = 10.9$ kcal/mol as compared with 11.4 kcal/mol for the analogous keto–enol pairs. These similarities indicate clearly that properties unique to oxygen are not responsible for the contrast between cyclic and acyclic keto–enol systems discussed in the preceding paper.

Comparison of the fluoroalkene isomerizations with those of the corresponding hydrocarbons is also informative. In particular, the contrast between the cyclic and acyclic systems is not seen in the corresponding hydrocarbons: the exo-endo

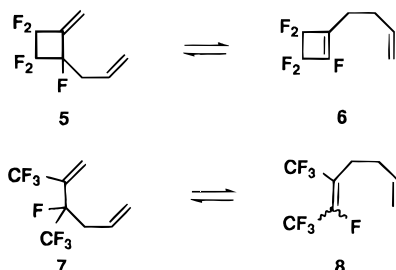
(7) Lowrey, A. H.; George, C.; D’Antoni, P.; Karle, J. *J. Mol. Struct.* **1979**, *53*, 189.

energy gaps for the hydrocarbon parents of systems **1**, **2**, and **3** were calculated to be 0.0, -3.8, and -1.0 kcal/mol, respectively, at the HF/6-31G**//6-31G** level of theory.⁸ Comparison of the pair of fluorinated isomers **4a** and **4b** with their hydrocarbon



analog is striking because the difference in their isomerization energies is so large. The exo-endo energy gap for system **4** is calculated to be -2.6 kcal/mol, while that of its hydrocarbon counterpart is +12.3 kcal/mol at the HF/6-31G**//6-31G** level of theory.⁹

The results in Figure 1 could be subjected to experimental test, of course, by measuring the heats of hydrogenation for each isomer. Since this seemed tedious from a synthetic point of view, we decided to take a different approach. Cope rearrangement should make it possible to equilibrate the double bond in each system as shown below.

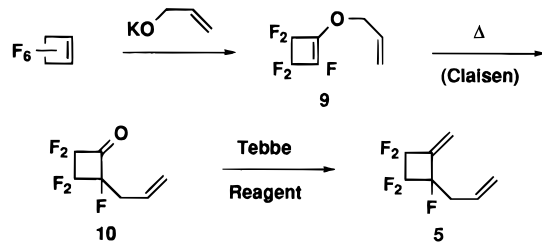


Dienes **6** and **8** (*E* and *Z*) were prepared by the reaction of the Grignard reagent from 4-bromo-1-butene with perfluorocyclobutene and (*E*)- and (*Z*)-perfluoro-2-butene (4:1), respectively.¹⁰ When a mixture of (*Z*)-**8** and (*E*)-**8** (1.3:1.0) was heated in a sealed ampule at 300 °C for 3 h, Cope rearrangement took place smoothly to diene **7**. The ratios of the three isomers were measured by gas chromatography to give these equilibrium constants (300 °C): $K_{(Z)8/7} = 0.015 \pm 0.002$ and $K_{(E)8/7} = 0.012 \pm 0.002$. The same values were obtained after 5 h of heating at 300 °C.

The energetics of the cyclic system were very different, however, as heating cyclobutene **6** for 3 h at 300 °C generated only a small amount (3.5%) of a compound whose ¹⁹F NMR spectrum was consistent with structure **5**. To confirm the assignment, independent synthesis of methylenecyclobutane **5** was pursued. It seemed possible that one could move an endocyclic double bond to the exocyclic position by taking advantage of the fact that aliphatic Claisen rearrangements strongly favor the ketone by virtue of the greater strength of the C=O as compared with the C=C bond.¹¹ A Wittig-type reaction should then generate methylenecyclobutane **5** as shown in Scheme 1.

In light of our finding that in system **A** the enol is more stable than its ketone, the equilibrium in the case of **9** and **10** undoubtedly favors the ketone less than would be the case in

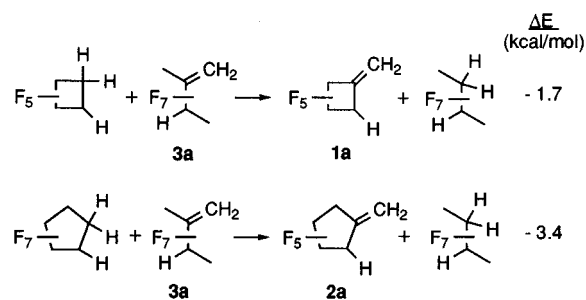
Scheme 1



their hydrocarbon counterparts. Quantum mechanical calculations on a model system predict that the ketone will be still favored, however, for 2-methylperfluorocyclobutanone is 14.2 kcal/mol more stable than 1-methoxyperfluorocyclobutene at the HF/6-311G**//3-21G level of theory. Indeed, rearrangement of allyl vinyl ether **9** was complete after heating for 1 h at 200 °C in 1,2,4-trichlorobenzene. Treating the ketone with methylenetriphenylphosphorane generated compound **5**, but in very low yield. Fortunately, much better yields were obtained with the Tebbe reagent.¹²

Heating compound **5** at 300 °C in a sealed tube for 10 h confirmed the ratio of **6**-to-**5** obtained above; i.e., $K_{6/5}$ (300 °C) = 28 ± 2 . The $\Delta\Delta G$ for dienes **6/5** versus (*Z*)-**8/7** is therefore 8.6 ± 0.2 kcal/mol at 300 °C. The similarity of this value to the $\Delta\Delta E$ for systems **1** and **3** in Figure 1 (10.4 kcal/mol) may well be fortuitous, as entropy, heat capacity, and zero-point energy differences are not accounted for in this comparison of numbers referenced to 573 and 0 K, respectively. Nonetheless, the experimental results constitute strong evidence for the sharp contrast between systems **1** and **3** predicted in Figure 1.

Reasons for the Cyclic-Acyclic Contrast. What is the origin of the differences in energetics of double bond isomerization between the cyclic and acyclic systems discussed above? Are highly fluorinated methylenecyclobutanes and -pentanes less stable than their acyclic counterparts? The isodesmic equations below (calculated at the HF/6-31G**//6-31G** level) indicate that the answer is no, with saturated compounds for reference.



Highly fluorinated acyclic alkenes must therefore be less stable than cyclic analogs. The following isodesmic reactions, which employ the same reference compounds, confirm this surmise. It is somewhat surprising that these acyclic fluoroalkenes (and enols), which are free to choose their configuration, are less stable than their configurationally locked cyclic counterparts. Since the difference is just as large when the hydroxyl group of the enol is replaced with a methyl group, it seems unlikely that resonance effects are important. We therefore turned our attention to the possibility that F-F nonbonded repulsions present in acyclic alkenes are relieved in cyclic alkenes.

The magnitude of the fluorine steric effect has been the subject of much debate over the years, partly because different

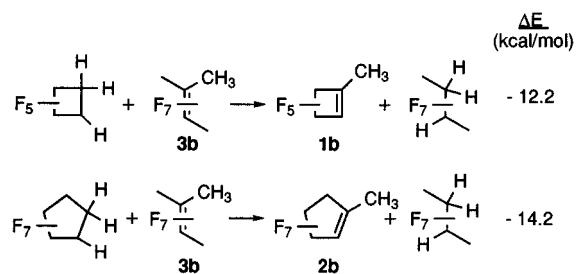
(8) The experimental values in the liquid phase (no solvent) are as follows: -0.54 ± 0.24 , -3.65 ± 0.22 , and -1.82 ± 0.06 kcal/mol, respectively. Wiberg, K. B.; Hao, S. *J. Org. Chem.* **1991**, *56*, 5108.

(9) The experimental value in the gas phase is $+10.3 \pm 0.52$ kcal/mol. Wiberg, K. B.; Fenoglio, R. A. *J. Am. Chem. Soc.* **1968**, *90*, 3395.

(10) For reactions of fluoroalkenes with Grignard reagents, see: Park, J. D.; Croft, T. S.; Anerson, R. W. *J. Organomet. Chem.* **1974**, *64*, 19.

(11) Bond energies for C=O and C=C are typically 173–181 and 146–151 kcal/mol, respectively. March, J. *Advanced Organic Chemistry*, 4th ed.; Wiley: New York, 1992; p 24.

(12) (a) Tebbe, F. N.; Parshall, B. W.; Reddy, G. S. *J. Am. Chem. Soc.* **1978**, *100*, 3611. (b) Pine, S. H.; Shen, G. S.; Huang, H. A. *Synthesis* **1991**, 165.



values for the van der Waals radius for fluorine have been employed. The old Pauling value¹³ of 1.35 Å is often used, along with the notion that F and H are about the same size. This has recently been challenged by Smart,¹⁴ who argues based on X-ray crystal data, modified Taft steric parameters, and Bondi's more recent van der Waals radius value¹⁵ of 1.47 Å for fluorine that "F and O, not H, are very nearly isosteric." The sum of the van der Waals radii for two F atoms is therefore approximately 3 Å, but the potential function for their interaction is not known with accuracy.¹⁶ In addition, a contribution to the fluorine steric effect undoubtedly arises from the fact that C–F bonds are about 20% longer than C–H bonds.²

Supporting the view that F–F steric effects are significant in fluoroalkenes are *ab initio* calculations carried out by Dixon on perfluoro-1,3-butadiene,¹⁷ using a DZ+D_C basis set. According to Dixon, the most stable conformer has an s-skew structure with a dihedral angle of 58.4° (relative to the s-cis angle of 0°). In the s-cis conformer, the C–C bond lengths by nearly 0.02 Å and the C=C–C bond angle increases by 4° relative to the same parameters in the s-skew conformer. These geometric variations apparently arise from strong steric repulsion between the two inner fluorine atoms in the s-cis form, which are only 2.51 Å apart as opposed to 3.04 Å in the s-skew form. Steric effects were also evident from the optimized parameters in the s-trans conformer, as the distance between the cisoid fluorines at C₁(C₂) and C₃(C₄) is only 2.62 Å. It was estimated that the F–F nonbonded repulsions in each planar conformer (Figure 2) are responsible for more than 5 kcal/mol of destabilization energy relative to their hydrocarbon parents.

More relevant to our study is the magnitude of steric effects in trifluoromethyl-substituted alkenes. It is commonly accepted that strain in 1,2-dialkylalkenes manifests itself in two ways:¹⁸ higher energy content and larger C=C–R angle in the cis form as compared with the trans. One would expect these characteristics to be prominent in 1,2-trifluoromethyl-substituted alkenes. This is not the case, however, as the cis-minus-trans enthalpy difference for perfluoro-2-butene¹⁹ is essentially the same as that for 2-butene (0.8 vs 1.0 kcal/mol, respectively),²⁰ despite the fact that CF₃ is much larger than CH₃.²¹ Perhaps even more surprising are the results of an electron diffraction study on *cis*- and *trans*-2,3*H*-perfluoro-2-butene that found the C=C–CF₃ bond angles to be nearly identical in the two isomers.²²

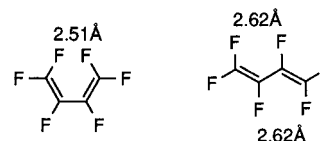


Figure 2. F–F distances (Å) in planar conformations of perfluoro-1,3-butadiene (Dixon).

Thus, there appeared to be a contradiction as to the amount of destabilization that results from F–F nonbonded repulsions in highly fluorinated alkenes. In an effort to settle this conflict we have carried out quantum mechanical calculations on the *cis* and *trans* isomers of perfluoro-2-butene and 2,3*H*-perfluoro-2-butene.²³ Initial optimizations were carried out at the HF/3-21G level of theory; the coordinates were used as starting points for the higher level optimizations. Final optimizations were performed at the HF/6-31+G** level because it has been shown that inclusion of diffuse functions is necessary to correctly order the relative energies of fluorinated alkenes.^{4f} Computational studies using similar basis sets have shown that calculations at the Hartree–Fock level underestimate C=C, C–F, and C–C bond lengths by 0.01–0.02 Å. Since bond angles, on the other hand, compare favorably with experiment, meaningful conclusions can be drawn from the optimized structures.²⁴ The important geometric parameters for perfluoro-2-butene are shown in Table 1.

The *cis*/*trans* energy gap for perfluoro-2-butene is calculated to be 1.8 kcal/mol at the HF/6-31+G**//6-31+G** level of theory, which agrees reasonably well with the experimental value of 0.8 kcal/mol.¹⁹ Virtually all of the important geometric parameters for *cis*- and *trans*-perfluoro-2-butene can be rationalized on the basis of minimizing F–F repulsions; interatomic distances are shown in Figure 3.

In the *cis* isomer, the C₄–F₄ bond bisects the angle made by F₁–C₁–F_{1'} so as to increase the distance between F₄ and F₁ (and F_{1'}) to 2.68 Å. The interaction between the CF₃ groups is reflected in the large C₁C₂C₃ and C₂C₃C₄ bond angles; there is no evidence for out-of-plane twisting about the C=C bond. The shortest F–F interatomic distance is between F_{1'} and F₂ (2.46 Å). Severe repulsion between these fluorines is reflected in the smaller C₁C₂C₃ angle (relative to C₂C₃C₄) and the fact that the C₁C₂ bond is 0.020 Å longer than the C₃C₄ bond.

trans-Perfluoro-2-butene also has several close F–F interatomic distances. The consequences of these interactions are displayed in the surprisingly large C₁C₂C₃ and C₂C₃C₄ bond angles, which are only 1.2° and 2.2° smaller than the respective angles in the *cis* isomer. This reflects the F–F nonbonded repulsions between F₁–F₃ and F₂–F₄, which are separated by only 2.59 Å. The F₁–F₃ or F₂–F₄ distances cannot be increased, of course, without decreasing the F₁(F_{1'})–F₂ or F₄–

(22) Burger, H.; Pawelka, P.; Oberhammer, H. *J. Mol. Struct.* **1982**, *84*, 49.

(23) Care was taken to be reasonably sure that global minima had been located for *cis*- and *trans*-perfluoro-2-butene by using different input geometries for the geometry optimizations. For example, for the *trans* isomer the dihedral angles made by F₁C₁C₂F₂ (*t*₁) and F₄C₄C₃F₃ (*t*₂) were varied. The different combinations for (*t*₁, *t*₂) were (0°, 0°), (0°, 60°), (60°, 60°), and (30°, 30°). The (60°, 60°) structure gave the lowest energy. The (*t*₁, *t*₂) torsion angles for the optimized structure were then changed to (30°, 30°) and the structure was reoptimized. Convergence back to the original (60°, 60°) optimized structure provides strong evidence for the location of a global minimum. The bond rotations were performed in the Spartan graphics module. A similar technique was used in the optimization of *cis*-perfluoro-2-butene, *cis*- and *trans*-2,3*H*-perfluoro-2-butene, and enol B. All vibrational frequencies were shown to be real for all minima.

(24) (a) Hehre, W. J.; Radom, L.; Schleyer, P. v. R.; Pople, J. S. *Ab Initio Molecular Orbital Theory*; Wiley: New York, 1986. (b) Reference 14. (c) Mack, H. G.; Steger, B.; Oberhammer, H. *Chem. Phys. Lett.* **1986**, *129*, 582. (d) Oberhammer, H.; Bauknight, C. W.; DesMarteau, D. D. *Inorg. Chem.* **1989**, *28*, 4340.

(13) Pauling, L. *The Nature of the Chemical Bond*, 3rd ed.; Cornell University Press: Ithaca, NY, 1960.

(14) Smart, B. E. In *Organofluorine Chemistry*; Banks, R. E., Smart, B. E., Tatlow, J. C., Eds.; Plenum Press: New York, 1994; p 80.

(15) Bondi, A. *J. Phys. Chem.* **1964**, *68*, 441.

(16) Allinger, N. L. *J. Am. Chem. Soc.* **1977**, *99*, 8127. For recent progress with this problem, see: Holt, D. R.; Farmer, B. L.; Macturk, K. S.; Eby, R. K. *Polymer* **1996**, *37*, 1847.

(17) Dixon, D. A. *J. Phys. Chem.* **1986**, *90*, 2038.

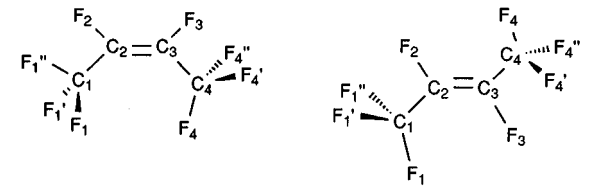
(18) (a) Zefirov, N. S.; Sokolov, V. I. *Russ. Chem. Rev.* **1967**, *36*, 87.

(b) Robin, M. B.; Taylor, G. N.; Kuevleer, N. A. *J. Org. Chem.* **1973**, *38*, 1049. (c) Ermer, O.; Lifson, S. *Tetrahedron* **1974**, *30*, 2425.

(19) Schlag, E. W.; Kaiser, E. W. *J. Am. Chem. Soc.* **1965**, *87*, 1171.

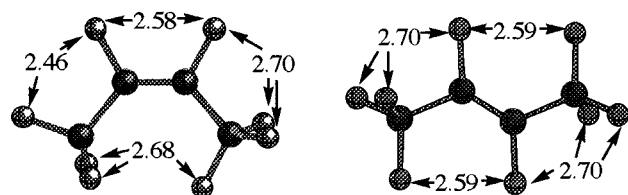
(20) Conant, J. B.; Kistiakowsky, G. B. *Chem. Rev.* **1937**, *20*, 181.

(21) It has been estimated that a CF₃ group is at least as large as an isopropyl group (ref 14).

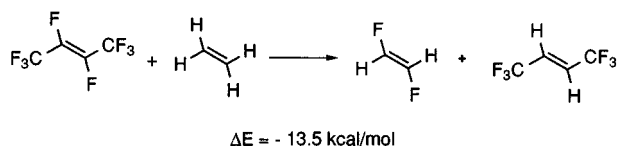
Table 1. Geometric Parameters for *cis*- and *trans*-Perfluoro-2-butene at the HF/6-31+G** Level


parameter ^{a,b}	<i>cis</i>	<i>trans</i>
$r(\text{C}_1\text{C}_2)$	1.523	1.510
$r(\text{C}_2\text{C}_3)$	1.313	1.312
$r(\text{C}_3\text{C}_4)$	1.503	1.510
$r(\text{C}_1\text{F}_1)$	1.317	1.310
$r(\text{C}_1\text{F}_1')$	1.317	1.318
$r(\text{C}_1\text{F}_1'')$	1.311	1.318
$r(\text{C}_2\text{F}_2)$	1.315	1.318
$r(\text{C}_3\text{F}_3)$	1.317	1.318
$r(\text{C}_4\text{F}_4)$	1.311	1.310
$r(\text{C}_4\text{F}_4')$	1.316	1.318
$r(\text{C}_4\text{F}_4'')$	1.316	1.318
$\theta(\text{C}_1\text{C}_2\text{C}_3)$	129.1	127.9
$\theta(\text{C}_2\text{C}_3\text{C}_4)$	130.1	127.9
$\theta(\text{C}_1\text{C}_2\text{F}_2)$	112.6	110.4
$\theta(\text{C}_4\text{C}_3\text{F}_3)$	110.7	110.4

^a Bond distances in angstroms. ^b Bond angles in degrees.

**Figure 3.** F–F distances (Å) in *cis*- and *trans*-perfluoro-2-butene.

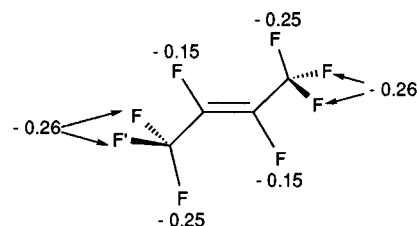
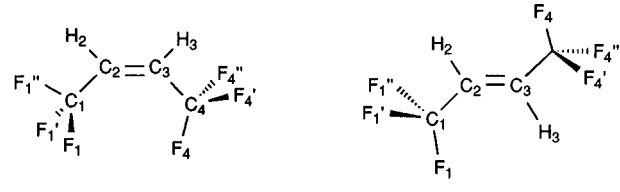
(F_4'')– F_3 distances, which are only 2.70 Å in the optimized structure. Perfluoroalkenes therefore suffer from a type of buttressing effect even in the *trans* configuration. The F–F nonbonded repulsions in *trans*-perfluoro-2-butene are at least partly responsible for the surprisingly small *cis*/*trans* energy gap.²⁵ The large magnitude of the steric effects in the *trans* isomer is demonstrated by the strongly exothermic isodesmic reaction shown below, calculated at the HF/6-31G**//6-31G** level of theory.



We note that the “steric” destabilization is not purely steric, as in the interpenetration of two neutral atoms. There is an additional electrostatic contribution. For example, the fluorine atoms of *trans*-perfluoro-2-butene all bear substantial negative charges. The charges shown in Figure 4 were calculated at the HF/6-31+G** level of theory; they are based on fits to electrostatic potentials, but the charges based on natural orbitals lead to the same conclusion.

Since *trans*-2,3*H*-perfluoro-2-butene does not suffer from F–F nonbonded repulsions, it is difficult to understand the claim that its C=C–CF₃ bond angle is nearly identical to the corresponding angle in the *cis* isomer. A strong possibility is that the structures obtained by electron diffraction are in error. The optimized parameters for the quantum mechanical calcula-

(25) There may also be a contribution from the “*cis* effect” as in 1,2-difluoroethylene (ref 3).

**Figure 4.** Electrostatic potentials at fluorine in *trans*-perfluoro-2-butene.**Table 2.** Geometric Parameters for *cis*- and *trans*-2,3*H*-Perfluoro-2-butene at the HF/6-31+G** Level


parameter ^{a,b}	<i>cis</i>		<i>trans</i>	
	expt	6-31+G**	expt	6-31+G**
$r(\text{C}_1\text{C}_2)$	1.492(5)	1.506	1.481(5)	1.498
$r(\text{C}_2\text{C}_3)$	1.310(16)	1.316	1.296(20)	1.312
$r(\text{C}_3\text{C}_4)$	1.492(5)	1.507	1.481(5)	1.498
$r(\text{C}_1\text{F}_1)$		1.325		1.320
$r(\text{C}_1\text{F}_1')$		1.317		1.324
$r(\text{C}_1\text{F}_1'')$		1.323		1.324
$r(\text{C}_2\text{H}_2)$		1.074		1.074
$r(\text{C}_3\text{H}_3)$		1.074		1.074
$r(\text{C}_4\text{F}_4)$		1.318		1.321
$r(\text{C}_4\text{F}_4')$		1.325		1.324
$r(\text{C}_4\text{F}_4'')$		1.325		1.324
$\theta(\text{C}_1\text{C}_2\text{C}_3)$	126.0(0.5)	128.2	125.1(1.6)	122.7
$\theta(\text{C}_2\text{C}_3\text{C}_4)$	126.0(0.5)	128.2	125.1(1.6)	122.7
$\theta(\text{C}_1\text{C}_2\text{H}_2)$		113.0		114.7
$\theta(\text{C}_4\text{C}_3\text{H}_3)$		113.0		114.7

^a Bond distances in angstroms. ^b Bond angles in degrees.

tions shown in Table 2 support this suggestion: the C=C–CF₃ bond angle for the *trans* isomer is calculated to be 122.7°, as opposed to 128.2° in the *cis* isomer.

The fact that virtually all of the calculated bond lengths are longer than the experimental values casts further doubt upon the validity of the electron diffraction study. Numerous calculations have shown that similar levels of theory underestimate bond lengths slightly.²⁴ The steric interactions of the CF₃ groups in the *cis* isomer not only widen the C=C–CF₃ angle but also result in twisting of the C=C bond by 1.2° out of planarity.

The size of the steric effect for the two CF₃ groups is large, for the *cis*/*trans* energy gap is 6.6 kcal/mol at the HF/6-31+G**//6-31+G** level. This suggests that a CF₃ group is effectively much larger than an isopropyl group though smaller than a *tert*-butyl group,²⁶ as the *cis*/*trans* energy differences for 1,2-diisopropylethylene²⁷ and 1,2-di-*tert*-butylethylene²⁸ are 1.9 and 9.3 kcal/mol, respectively.

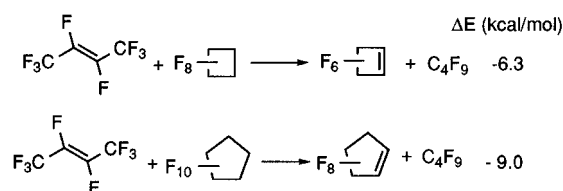
We now return to the question of why acyclic alkenes are more destabilized than cyclic analogs, as further documented

(26) This agrees nicely with modified Taft steric parameters, e.g., E_s: H = 0; CH₃ = –1.24; (CH₃)₂CH = –1.71; CF₃ = –2.40; (CH₃)₃C = –2.78. Unger, S. H.; Hansch, C. In *Progress in Physical Organic Chemistry*; Taft, R. W., ed.; John Wiley & Sons: New York, 1976; Vol. 12, pp 91–118. It should be noted, however, that the A value for CF₃ is much smaller than that for (CH₃)₃C. Eliel, E. L.; Wilen, S. H. In *Stereochemistry of Organic Compounds*; Wiley: New York, 1994; p. 697.

(27) Turner, R. B.; Jarrett, A. D.; Goebel, P.; Mallon, B. *J. Am. Chem. Soc.* **1973**, *95*, 790.

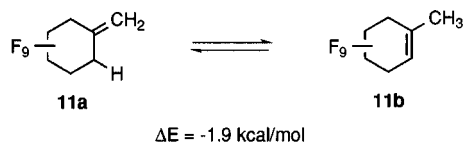
(28) Turner, R. B.; Nettleton, D. E.; Perelman, M. *J. Am. Chem. Soc.* **1958**, *80*, 1430.

by the isodesmic reactions shown below (HF/6-31G**//6-



31G**). Confining the double bond to a ring can remove the steric interactions present in the acyclic alkenes. For example, in perfluorocyclobutene and perfluorocyclopentene the calculated distance between the vinyl fluorine and the fluorines at the adjacent saturated carbon is 3.15²⁹ and 2.82 Å, respectively, at the HF/6-31G**//6-31G** level. In contrast, *trans*-perfluoro-2-butene has several F–F distances that are calculated at the same level of theory to be much shorter than the sum of the van der Waals radii. This represents a particularly striking and surprising manifestation of the fluorine steric effect and clearly indicates that steric considerations should be entertained when discussing highly fluorinated alkenes. In focusing on *cis*- and *trans*-perfluoro-2-butene, we have considered the simplest acyclic, internal perfluoroalkenes; higher homologs must suffer from greater steric repulsions.

In the above discussion cyclic systems have been represented only by 3- to 5-membered rings; what about 6-membered rings? The isomerization energy shown below for system **11** was



calculated at the HF/6-311G**//3-21G level of theory. Values for systems **1**, **2**, and **3** at the same level are very similar to those in Figure 1: -6.5, -6.8, and +4.3 kcal/mol, respectively. Thus, the gap between cyclic and acyclic systems is still large (6.2 kcal/mole) when the ring is 6-membered, though significantly less than when it is 4- or 5-membered. We believe the reason for the intermediate position of the 6-ring system is that the steric interactions which so strongly influence acyclic systems are beginning to manifest themselves at this ring size. Support for this surmise is presented below.

Implications for Keto–Enol Systems. In the preceding paper, we concluded on the basis of quantum mechanical calculations that the greater thermodynamic stability of highly fluorinated cyclic enols with respect to their ketones as compared with acyclic enols can be attributed to greater destabilization of the latter by perfluorination. The reason for this difference is now clear: since oxygen and fluorine are isosteric,¹⁴ the steric arguments developed above for alkenes are all applicable to the keto–enol systems. Alkene and keto–enol systems parallel each other nicely, as the enol form is substantially more stable than the ketone in the 4-ring (system **A**) and 5-ring cases, is close in energy to the ketone in the 6-ring case, and is very strongly disfavored in acyclic examples such as system **B**.

The intermediate location of the 6-ring in this hierarchy may seem awkward at first, but in fact it helps to confirm the interpretation we have offered for the cyclic–acyclic contrast. When the geometries of the 4-, 5-, and 6-ring enols are optimized at the 3-21G level, a striking trend appears in the distances between the heteroatoms on the double bond and the fluorines vicinal to each (Figure 5). In the planar cyclobutenol

(29) The experimental value according to electron diffraction data is 3.20 Å (ref 6).

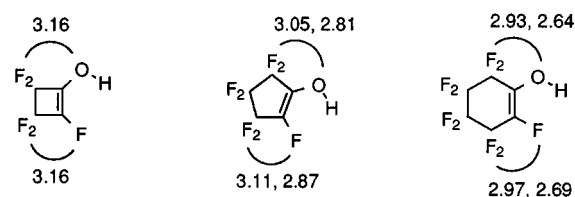


Figure 5. O–F and F–F distances (Å) in cyclic enols.

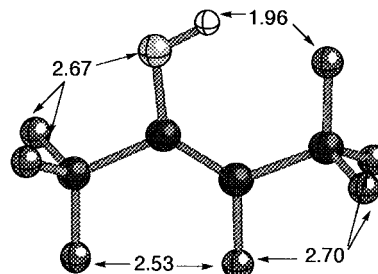


Figure 6. Nonbonded contacts (Å) in perfluoro-2-butenol.

all four distances are well beyond the sums of the van der Waals radii, which are 2.94 Å for the F–F and 2.99 Å for the O–F interactions.¹⁵ In the somewhat puckered cyclopentenol, two of those distances fall within the sums of radii, but they are not short enough to cause substantial steric strain. However, in the cyclohexenol three of the four distances lie within, and two of them well within, the sums of radii. This trend can account for the stability order 4- > 5- >> 6-ring, that we found for the cyclic enols. In other words, steric repulsions, which are powerful in acyclic examples, start to be significant in cyclic cases as well when the ring size increases to 6-membered. The same idea serves to explain the smaller isomerization energy revealed by calculation for the 6-ring alkenes **11a** and **11b** as compared with the homologous 4- and 5-ring systems.

That steric interactions are very important in acyclic perfluoro enols is apparent in Figure 6, which depicts the lowest energy conformer for enol **B** (perfluorobut-2-en-2-ol) at the 6-31G** level of theory. As in *trans*-perfluoro-2-butene, a large part of the destabilization is probably electrostatic in nature. This is true for both F–F and F–O interactions, as the oxygen atom of enol **B** bears a large (-0.57) negative charge. The charges at the two fluorines that are only 2.67 Å away from the oxygen are -0.23.

One must consider the possible role of H–F intramolecular hydrogen bonding³⁰ in stabilizing enol **B**, however, especially since the hydrogen and a fluorine of the CF₃ group are only 1.96 Å apart at the HF/6-31G** level; the sum of the (Bondi¹⁵) van der Waals radii for H and F is 2.67 Å.³¹ To assess the strength of this hydrogen bond, the energies of conformers rotated about the C–O bond were examined at the HF/6-31G** level of theory (without reoptimization). Of course, this rotation also results in some loss of resonance energy as an oxygen lone pair is no longer ideally aligned to interact with the C=C bond. For comparison, the C–O bond of enol **A** was also rotated. Here, the geometric constraints of the four-membered ring make H–F hydrogen bonding impossible, so differences in the two

(30) For evidence of intramolecular H–F hydrogen bonding in fluorinated enols, see: (a) Dixon, D. A.; Smart, B. E. *J. Phys. Chem.* **1991**, *95*, 1609. (b) Kitazume, T. *Synthesis* **1986**, 855. (c) Kitazume, T. *J. Fluorine Chem.* **1987**, *35*, 287.

(31) Recently, an extremely extensive study of nonbonded contact distances in crystal structures has provided strong support, in general, for Bondi's values of van der Waals radii. The only notable exception is hydrogen, for which the value 1.10 Å is recommended to replace Bondi's 1.20 Å. The discrepancy arises because H···X (e.g. X = N, O, F) contacts tend to be shorter than H···H, which formed the basis for Bondi's value. Rowland, R. S.; Taylor, R. *J. Phys. Chem.* **1996**, *100*, 7384.

Table 3. Torsional Potential about the C–O Bond for Enols **A** and **B** at the HF/6-31G**//6-31G** Level of Theory

<i>t</i> (deg)	relative energy (kcal/mol)	
	enol A	enol B
0	0	0
30	+1.04	+1.22
90	+3.74	+4.05

rotational potentials should reflect the strength of the hydrogen bond in the acyclic enol. The values in Table 3 clearly show that a 90° rotation results in approximately the same amount of destabilization in each enol, thereby indicating that H–F bonding in enol **B** must be really weak. If there were no hydrogen bonding, the very short H–F distance would imply substantial steric repulsion; the similarity of the **A** and **B** rotational potentials thus suggests that hydrogen bonding and steric repulsion are simultaneously present but that they nearly cancel one another, resulting in little *net* bonding. Stated differently, the H–F interaction in enol **B** appears to be a compressed hydrogen bond.³²

Conclusion

Our research on highly fluorinated enols revealed that acyclic enols are much more destabilized by perfluorination than their cyclic counterparts, as described in the preceding paper. The present work has demonstrated that this finding is not unique to enols, but is rather a consequence of severe steric repulsions that are present in highly fluorinated, vicinally substituted acyclic alkenes in general. Examination of HF/6-31+G* optimized geometries for *cis*- and *trans*-perfluoro-2-butene, for example, shows that both isomers have F–F interatomic distances that are much shorter than the sum of the van der Waals radii for two fluorine atoms. This effect is absent or small in cyclic alkenes up to 6-membered, depending on ring size, for constraining the double bond to a ring reduces the steric repulsions by increasing interatomic distances. The cyclic–acyclic contrast represents one of the more surprising manifestations of the fluorine steric effect and further dispels the myth that fluorine and hydrogen atoms are about the same size.

In earlier papers³³ we reported that the remarkable stability enjoyed by cyclic perfluoro enols is attributable to destabilization of the keto form by fluorine substitution, an electronic effect. The same phenomenon characterizes acyclic fluoroketones, but here the effect of higher ketone energy on K_{EK} is negated to a large degree by the steric destabilization of their enols.

Experimental Section

¹⁹F NMR spectra were recorded at 282.2 MHz, and all chemical shifts are reported on the ϕ scale (ppm from internal trichlorofluoromethane, upfield negative). The ¹H NMR spectra were recorded at 300 MHz with tetramethylsilane as an internal standard, and chemical shifts are reported on the δ scale. Carbon tetrachloride was distilled from phosphorus pentoxide and acetonitrile from calcium hydride. Diethyl ether and tetrahydrofuran were distilled from potassium benzophenone ketyl. Tebbe reagent was purchased from Aldrich as a

(32) A new investigation of fluorine as a hydrogen-bond acceptor has found, using *ab initio* calculations, an optimal H···F distance of 1.9 Å for the complex of water with fluoromethane, but a search of intramolecular C(sp³)–F···H–O contacts in the Cambridge Structural Database System revealed an average distance >2.20 Å for the handful of examples found. Only molecules in which a single fluorine is bound to carbon were considered. Howard, J. A. K.; Hoy, V. J.; O'Hagen, D.; Smith, G. T. *Tetrahedron* **1996**, *52*, 12613.

(33) Correa, R. A.; Lindner, P. E.; Lemal, D. M. *J. Am. Chem. Soc.* **1994**, *116*, 10795. Lindner, P. E.; Correa, R. A.; Gino, J.; Lemal, D. M. *J. Am. Chem. Soc.* **1996**, *118*, 2556. Lindner, P. E.; Lemal, D. M. *J. Org. Chem.* **1996**, *61*, 5109.

0.5 M solution in toluene. Perfluorocyclobutene was prepared according to literature methods³⁴ and perfluoro-2-butene was purchased from PCR, Inc.

Analytical gas chromatograms were obtained with a 25-m methyl silicone capillary column and flame ionization detector. The standard program was as follows: carrier pressure 25 psi; injector temperature 150 °C; detector temperature 200 °C; column temperature noted in the text. Isolation of pure compounds was done by preparative GC using a thermal conductivity detector. The column used was 20 ft × 1/8 in, 10% OV-101 on 80/100 mesh Chromosorb-P. The standard program was as follows: injector temperature 190 °C; detector temperature 210 °C; column temperature noted in the text. Elemental analyses were performed by Schwarzkopf Microanalytical Laboratory, Woodside, NY.

Pyrolyses were performed in a metal pipe wrapped with heating tape. The temperature was controlled by an Omega temperature controller, Model 149, and monitored by means of a thermocouple attached to the ampule.

The quantum mechanical calculations were performed on Gaussian 92, Revision C.3.³⁵ The electrostatic point charges were calculated with the Spartan package of programs.³⁶

2-(Allyl)-1-methyleneperfluorocyclobutane (5). To a 50-mL 3-neck round-bottom flask was added 5.0 mL of Tebbe's reagent (0.5 M in toluene, 2.5 mmol). The toluene was removed in vacuo and replaced with 5.0 mL of 1,2,4-trichlorobenzene. With the system under an atmosphere of nitrogen, a solution containing 0.50 g (2.5 mmol) of ketone **10** in 10 mL of 1,2,4-trichlorobenzene was added slowly to the Tebbe reagent.¹² After 15 min, the volatile products were passed through two traps connected in series (–13 and –196 °C). The former contained only solvent and the latter contained crude **5**, which was isolated by preparative GC (110 °C; t_R = 6.0 min). A total of 110 mg was recovered (22% yield). ¹⁹F NMR (CDCl₃): –113.5, –115.8 (AB q, J = 227 Hz, 2F); –127.4, –130.6 (AB q, J = 215 Hz, 2F); –163.2 (m, 1F). ¹H NMR: 6.11 (m, 1H); 5.96 (m, 1H); 5.89 (m, 1H); 5.31 (d, J = 8 Hz, 1H); 5.27 (d, J = 17 Hz); 2.79 (m, 1H); 2.56 (m, 1H). IR (thin film, cm^{–1}): 3090, 3029, 2991, 2915, 1693, 1640, 1421, 1353, 1148, 1064, 1003, 958, 829. MS *m/e*: 198 (M⁺), 127 (base). Anal. Calcd for C₈H₇F₅: C, 48.50; H, 3.56. Found: C, 48.26; H, 3.61.

1-(4-Butenyl)perfluorocyclobutene (6). The Grignard reagent was prepared by adding 1.4 mL (14 mmol) of 4-bromo-1-butene to 0.34 g of magnesium metal (14 mmol) in dry ethyl ether. In a separate vessel, 2.0 g (12 mmol) of perfluorocyclobutene was distilled into 10 mL of ether with the aid of a dry ice condenser. With the vessel cooled to 0 °C, the Grignard reagent was added dropwise to the olefin over a 5-min period. The contents of the flask were stirred well for 3 h and then treated with 20 mL of a saturated NH₄Cl solution. The ether layer was separated and washed with 10 mL of water, dried over MgSO₄ and filtered. The solvent was removed by short-path distillation to leave a yellow oil. All high boiling material was left behind with a vacuum transfer at 25 °C/30 mTorr. Compound **6** was isolated pure by preparative GC (120 °C, t_R = 6.0 min). A total of 450 mg (19% yield) was recovered. ¹⁹F NMR (CDCl₃): –115.4 (t, J = 10 Hz, 1F); –117.4 (m, 2F); –118.2 (s, 2F). ¹H NMR: 5.79 (m, 1H); 5.11 (d, J = 19 Hz, 1H); 5.09 (d, J = 11 Hz, 1H); 2.4 (broad m, 4H). IR (thin film, cm^{–1}): 3081, 2932, 1727, 1644, 1399, 1329, 1241, 1110, 926. MS *m/e*: 198 (M⁺).

Equilibration of Alkenes 5 and 6. A heavy-walled Pyrex ampule containing 30 mg (0.31 mmol) of cyclobutene **6** was connected to the vacuum line. The sample was frozen at liquid nitrogen temperature and the ampule was evacuated to 30 mTorr. The glass bomb was then sealed with a torch and heated at 300 °C for 2 h. After the tube had been cooled to liquid nitrogen temperature it was cracked open. The clear, colorless liquid was dissolved in chloroform-*d*; its ¹⁹F NMR spectrum revealed the presence of dienes **5** and **6** in a 3.5:96.5 ratio.

(34) Buxton, M. W.; Ingram, D. W.; Smith, F.; Stacey, M.; Tatlow, J. C. *J. Chem. Soc.* **1952**, 3830.

(35) Gaussian 92, Revision C.3 (Frisch, M. J.; Trucks, G. W.; Head-Gordon, M.; Gill, P. M. W.; Wong, M. W.; Foresman, J. B.; Johnson, B. G.; Schelegel, H. B.; Robb, M. A.; Replogle, E. S.; Gomperts, R.; Andres, J. L.; Raghavachari, K.; Binkley, J. S.; Gonzalez, C.; Martin, R. L.; Fox, D. J.; Defrees, D. J.; Baker, J.; Stewart, J. J. P.; Pople, J. A. Gaussian, Inc.; Pittsburgh, PA, 1992).

(36) Spartan; Hehre, W.; Wavefunction, Inc.: 188401 Von Karman, Suite 370, Irvine, CA 92717.

Similar values were obtained when **6** was heated at 300 °C for 10 h and also when the equilibrium was approached from compound **5**. The average of the three experiments gave $K_{65}(300^\circ\text{C}) = 28 \pm 2$.

2-(4-Butenyl)perfluoro-2-butene (8). The Grignard reagent was prepared by gently warming a mixture of 1.2 mL (12 mmol) of 4-bromo-1-butene and 0.30 g (12 mmol) of magnesium metal in 10 mL of dry triglyme. In a separate vessel, 2.0 g (10 mmol) of (*E*)- and (*Z*)-perfluoro-2-butene (4:1) was distilled into 10 mL of triglyme with the aid of a dry ice condenser. With the vessel cooled to 0 °C, the Grignard reagent was added dropwise over a 5-min period. The contents of the flask were stirred well for 3 h, then the flask was attached to the vacuum line. Volatiles were passed through two traps connected in series, with the first cooled to -10 °C (ice, NaCl) and the second chilled to -78 °C (dry ice/2-propanol). The second trap contained the *Z* and *E* isomers of compound **8**, along with approximately 10% of perfluoro-2-butene. The *Z* and *E* isomers (1.3:1.0) were separated by preparative GC (90 °C, $t_{\text{R}}^{\text{Z}} = 6.0$ min; $t_{\text{R}}^{\text{E}} = 6.4$ min).³⁷ ¹⁹F NMR (CDCl₃): *Z* isomer: -62.1 (d, $J = 21$ Hz, 3F); -66.9 (d, $J = 7$ Hz, 3F); -115.8 (m, 1F). *E* isomer: -59.5 (m, 3F); -67.3 (m, 3F); -116.1 (s, 1F). ¹H NMR (both isomers): 5.77 (m, 1H); 5.13 (d, $J = 19$ Hz, 1H); 5.09 (d, $J = 11$ Hz, 1H); 2.46 (t, $J = 8$ Hz, 2H); 2.28 (m, 2H). IR (thin film, cm⁻¹): *Z* isomer: 3082, 2976, 2938, 1700, 1640, 1368, 1277, 1209, 1163, 1072, 996, 837. *E* isomer: 3082, 2984, 2931, 1696, 1640, 1345, 1309, 1223, 1172, 1057. Anal. Calcd for C₈H₇F₇: C, 40.70; H, 2.99; F, 56.31. Found: C, 40.61; H, 3.03; F, 56.13.

Equilibration of Alkenes 7 and 8. A mixture of (*Z*)- and (*E*)-**8** in a ratio of 1.3:1.0 was placed in a heavy-walled Pyrex ampule, frozen, evacuated, sealed, and heated at 300 °C for 2 h. After being cooled to liquid nitrogen temperature the tube was cracked open. Then 0.1 μL of the clear, colorless liquid was withdrawn with a syringe and injected into the analytical gas chromatograph (column temperature = 50 °C). Integration values for the appropriate peaks ($t_{\text{R}}^{\text{Z}8} = 4.9$ min, $t_{\text{R}}^{\text{E}8} = 5.3$ min, $t_{\text{R}}^7 = 6.1$ min) were recorded. This experiment was repeated twice by heating the tubes at 300 °C for 5 h. The average of the three experiments gave $K_{\text{Z}8/7} = 0.015 \pm 0.002$ and $K_{\text{E}8/7} = 0.012 \pm 0.002$. The uncertainties are standard deviations. ¹⁹F NMR of **7** (CDCl₃): -63.9 (s, 3F); -80.3 (s, 3F); -174.2 (s, 1F). ¹H NMR (CDCl₃): 6.23 (d, $J = 5$ Hz, 1H); 5.89 (s, 1H); 5.70 (m, 1H); 5.26 (d, $J = 10$ Hz,

1H); 5.21 (d, $J = 18$ Hz, 1H); 2.86 (m, 2H). IR of **7** (thin film, cm⁻¹): 3090, 2991, 2960, 2930, 1640, 1603, 1520, 1413, 1352, 1322, 1277, 1201, 1148, 1034, 996.

1-(Allyloxy)perfluorocyclobutene (9). A 25-mL round-bottom flask was charged with 1.2 mL (18 mmol) of allyl alcohol and 2.0 g (12 mmol) of perfluorocyclobutene. With the vessel chilled in an ice bath, a solution containing 1.3 g (24 mmol) of KOH dissolved in 5 mL of water was added slowly with efficient magnetic stirring. After 30 min the organic products were extracted with 3 × 10 mL of pentane. The organics were back-washed with brine and dried over MgSO₄. The pentane was removed by short-path distillation to leave 1.4 g of **9** as a clear, colorless oil (58% yield). Compound **9** was obtained analytically pure by preparative GC (110 °C, $t_{\text{R}} = 8.0$ min). ¹⁹F NMR (CDCl₃): -116.3 (m, 2F); -119.5 (m, 2F); -140.4 (m, 1F). ¹H NMR: 5.98 (m, 1H); 5.47 (d, $J = 19$ Hz, 1H); 5.42 (d, $J = 11$ Hz, 1H); 4.73 (d, $J = 7$ Hz, 2H). IR (thin film, cm⁻¹): 3097, 2999, 2961, 2908, 1760, 1648, 1398, 1360, 1239, 1110, 1011. Anal. Calcd for C₇H₅F₅O: C, 42.02; H, 2.52; F, 47.46. Found: C, 41.71; H, 2.48; F, 47.37.

2-Allylperfluorocyclobutanone (10). Compound **9** (0.10 g, 0.50 mmol) was heated to 200 °C in 3.5 mL of dry 1,2,4-trichlorobenzene under an atmosphere of nitrogen. After 2 h the vessel was cooled to room temperature and attached to a vacuum line. The volatiles were passed through two traps connected in series with the first cooled to -10 °C (ice, NaCl) and the second cooled to liquid nitrogen temperature. The latter contained 0.090 g (90% yield) of pure ketone **10**. ¹⁹F NMR (CDCl₃): -121.5, -128.6 (AB q, $J = 270$ Hz, 2F); -130.9, -134.3 (AB q, $J = 230$ Hz, 2F); -173.2 (d, $J = 17$ Hz, 1F). IR (gas phase, cm⁻¹): 3006, 1831, 1641, 1277, 1186, 1064, 996, 936, 890.

Acknowledgment. The authors thank the donors of the Petroleum Research Fund, administered by the American Chemical Society, and the National Science Foundation for support of this research. P.E.L. is grateful to DuPont for a Walter H. Stockmayer Fellowship. Thanks are due to the Keck and the Dreyfus Foundations for grants to the Chemistry Department, which enabled purchase of the computer hardware and software used in this work.

Supporting Information Available: Tables of final optimized geometries (6-31G**) in Cartesian coordinates for *cis*- and *trans*-perfluoro-2-butene and *cis*- and *trans*-2,3*H*-perfluoro-2-butene (2 pages). See any current masthead page for ordering and Internet access instructions.

JA963789F

(37) The (*Z*) and (*E*) isomers were assigned on the basis of their ¹⁹F NMR spectra. In particular, one of the CF₃ groups in the *Z* isomer is a doublet with a coupling constant of 21 Hz. This large splitting results from through-space coupling with the vinyl fluorine that is *cis* to that CF₃ group. This coupling is also seen in the vinyl fluorine which appears as a multiplet. In contrast, the CF₃ groups for the *E* isomer are multiplets because they are through-space coupled to one another. The vinyl fluorine appears as a singlet because it is *cis* to the hydrocarbon chain. For a similar analysis, see compound **11** in the preceding paper.