the first water molecule; (3) the somewhat larger charge transfer to Me$_4$N$^+$.

One consequence of the above results is that any solvent that cannot participate in a solvent-to-solvent hydrogen bond, i.e., Me$_2$CO and MeCl, would be expected to have a lower ΔE_D for the second attachment than for the first. This is in complete agreement with the experimental results for Me$_2$CO.

Interaction of Me$_4$N$^+$ with Polyfunctional Ligands. The interaction energies of Me$_4$N$^+$ with ethers are shown in reactions 8–10, Table I. The interaction with the large, polarizable *n*-Bu$_2$O is somewhat stronger than that with H$_2$O and MeOH. While the triether in reaction 9 has a polarizability comparable to that of *n*-Bu$_2$O, the polydentate ligands give considerably more stable complexes than any of the other ligands. The strong interaction energies here clearly indicate multiple interactions of the ion with the multiple ether functions. A comparison of the *n*-Bu$_2$O and diglyme complexes indicates 8 kcal mol^{-1} for the second plus third CH$^{\delta+}$···O interactions in the latter. Interestingly, this is similar to the 10 kcal mol^{-1} interactions assigned for two CH$^{\delta+}$···O interactions in the analogous complex Me$_3$NH$^+$·diglyme.[3] Apparently the exchange of a strong vs. weak bond in the first interaction does not significantly affect the additional CH$^{\delta+}$···O interactions.

The question of multiple interactions is also of interest in the complex of Me$_4$N$^+$ with the amino acid derivative CH$_3$CONH-CH$_2$COOCH$_3$, i.e., CH$_3$CO-gly-OCH$_3$ (reaction 11, Table I). The geometries of the two carbonyl functions in CH$_3$CO-gly-OCH$_3$ are similar to those in two adjacent amide oxygens in a peptide. Therefore, the complex simulates the interactions of the charged group in acetylcholine with two adjacent amide links in an unionized peptide. The experimental results show that the interaction, 20 kcal mol^{-1}, is substantial. However, a comparable interaction energy is observed with a single amide group, i.e., dimethylacetamide (reaction 12, Table I). Molecular models

demonstrate that, with the oxygen of one amide group of CH$_3$CO-gly-OCH$_3$ in a three-proton cavity of Me$_4$N$^+$, the other carbonyl oxygen could interact only with the same protons, whose charge is already delocalized in the first interaction. This may explain the small effects of multiple interaction for this group as opposed to the polyether complexes, for which the oxygen can interact with different groups of methyl protons. Of course, with larger peptides, non-neighboring amide groups could bind the ion in a more efficient polydentate conformation.

Complexes of Other Quaternary Ions. Complexes of Quaternary Ions with π-Donors. With increased size of the quaternary ion, i.e., in Et$_4$N$^+$, the charge on each proton should be decreased and, consequently, the interaction energy with solvents should decrease. Indeed, the interaction energies with H$_2$O and Me$_2$CO are decreased by 2 kcal mol^{-1} (reactions 18 and 19, Table I) compared to the analogous complexes of Me$_4$N$^+$.

Finally, the complex Me$_3$O$^+$·Me$_2$O is comparable in stability to the complexes of Me$_4$N$^+$ with oxygen ligands. This suggests that there are similar charge distributions in the quaternary ammonium and tertiary oxonium ions. Reactions 16 and 17 show that the interaction of Me$_4$N$^+$ with the π-donor ligands benzene and toluene is similar to that with the oxygen and nitrogen ligands. Interestingly, benzene also shows an interaction energy with Me$_3$NH$^+$ comparable to regular n-donor ionic hydrogen bonds as will be seen in the following article in this issue. The complexes of Me$_4$N$^+$ with benzene and toluene represent a special type, CH$^{\delta+}$···π, where no n-donors are involved on either side.

Unconventional Ionic Hydrogen Bonds. 2. NH$^+$···π. Complexes of Onium Ions with Olefins and Benzene Derivatives

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Abstract: Unconventional strong ionic hydrogen bonds of the –XH$^+$···π type, where the electron donor is a π-bond or an aromatic π-system, are formed in the clustering reactions of NH$_4$$^+$ and MeNH$_3$$^+$ with C$_2$H$_4$ and benzene derivatives. The interaction energies range from 10 to 22 kcal mol^{-1}. The experimental results and ab initio calculations on C$_2$H$_4$·NH$_4$$^+$, C$_6H_6$·NH$_4$$^+$, and C$_6H_5$F·NH$_4$$^+$ indicate that the interaction is primarily electrostatic in nature with little π-donation into the bond. The most stable structure of C$_2$H$_4$·NH$_4$$^+$ is the conformer where one N–H$^+$ bond points at the center of the double bond. For C$_6$H$_6$·NH$_4$$^+$ and C$_6$H$_5$F·NH$_4$$^+$, the lowest energy π-dimers have two NH$_4$$^+$ hydrogens directed toward the ring. The F···H–NH$_3$$^+$ σ-complex was studied also for C$_6$H$_5$F·NH$_4$$^+$. The latter complex is the more stable of the two at this level of calculation.

Conventional ionic hydrogen bonds between protonated and neutral n-donors –XH$^+$···X–, ranging in strength up to 30 kcal mol^{-1}, have been investigated extensively.[1,2] Fewer studies have been done on unconventional ionic bonds, i.e., those where the atom on either side of the bonding, partially charged hydrogen is not an n-donor. In the preceding paper[3] we observed unconventional bonds of the CH$^{\delta+}$···X type ranging in strength from

6 to 18 kcal mol^{-1}. In the present paper we examine strong ionic hydrogen bonds where the proton is attached to an n-donor, but the electron donor is a π-bond or an aromatic π-system.

While the thermochemistry of XH$^+$···π interactions has not been investigated previously, Sunner and co-workers[4] have studied a

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† National Bureau of Standards.
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Unconventional Ionic Hydrogen Bonds. 2

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Figure 1. van't Hoff plots for NH⁺⋯π complexes. Numbers refer to Table I. •, O, and ■ in reaction 1 indicate points obtained in neat C_6H_6 and 1.1% and 0.1% C_6H_6 in cyclohexane, respectively.

$K^+⋯\pi$ system in the clustering of K^+ with benzene. The authors found that the attachment energy of 19.2 kcal mol⁻¹ was due mainly to electrostatic interactions, especially the attraction between K^+ and the benzene equatorial $\pi(-)$-ring(+)–$\pi(-)$ quadrupole. One point of interest in the present work is to compare the interaction of the potentially hydrogen bonding NH_4^+ vs. that of K^+ with a π-donor. More broadly it is of interest whether the $XH^+⋯\pi$ bonds involve significant π-donation into the bond or whether these hydrogen bonds are due mostly to electrostatic interactions.

In addition to their physical interest, $XH\delta^+⋯\pi$ interactions may also play a role in peptide conformation. Lilly[5] has found evidence that hydrogen bonds involving π-systems and partially charged amide hydrogens significantly affect the conformational thermochemistry of small peptides with aromatic constituents.

Experimental and Computational Details

Measurements were performed with use of the NBS pulsed high pressure mass spectrometer and standard procedures.[6] In most studies of the alkyl-ammonium ion–π-donor complexes, trace amounts of the amine (0.01%) and 0.1–2% of the ligand in cyclohexane or iso-C_4H_{10} as the carrier gas were used. In some systems, such as NH_4^+, $MeNH_3^+$, and $Me_3NH^+⋅C_6H_6$ and $NH_4^+⋅C_2H_4$, neat C_6H_6 or C_2H_4, respectively, was the carrier gas. In the $NH_4^+⋅C_6H_6$ system, points obtained in neat C_6H_6 and 0.1% to 18% C_6H_6 in c-C_6H_{12} all fit on the same van't Hoff plot (Figure 1) demonstrating that variation of the partial pressure of C_6H_6 by a factor of 10³ does not affect the equilibrium constant. Similarly, checks were made in other equilibria to show that variation of the partial pressure of the ligand, usually by factors of 2–5, and of the total source pressure, in the range 0.2–0.8 torr, did not affect the measured equilibrium constant.

Trace concentrations of amines ensured that the formation of amine dimer ions was not significant above ca. 380 K. At lower temperatures and with weak ligands (especially C_2H_4), $(NH_3)_2H^+$ formed quickly and the reaction prevented accurate $NH_4^+⋅$donor equilibrium determinations.

The theoretical calculations were carried out ab initio by using the GAUSSIAN 80 system of programs[7] on a VAX 11/780 computer. $C_2H_4⋅NH_4^+$, $C_2H_5NH_3^+$, $C_6H_6⋅NH_4^+$, and $C_6H_5F⋅NH_4^+$ were investigated. Three sets of calculations were performed. For the first set the structures of C_2H_4,[8] NH_4^+,[8] C_6H_6,[8] C_6H_5F,[9] $C_2H_4⋅NH_4^+$, $C_6H_6⋅NH_4^+$,

and $C_6H_5F⋅NH_4^+$ were optimized by using the STO-3G basis set.[10a] The geometries of C_2H_4,[8] NH_4^+,[8] C_6H_6,[8] and $C_6H_5F^9$ were optimized completely. These structures were retained in the complexes and only selected bond lengths between NH_4^+ and C_2H_4, C_6H_6, or C_6H_5F were varied. This approach of partially optimizing the structures of complexes has been utilized by many research groups.[11]

Since the 3-21G basis[10b] has been shown to yield more accurate reaction energies and relative molecular energies than the STO-3G basis, for the second set of calculations the total energies of all of the above species were obtained at the 3-21G level by using the STO-3G optimized geometries, i.e., 3-21G//STO-3G computations were performed. In order to check the accuracy of the 3-21G//STO-3G calculations, for the third set of computations C_2H_4[8] and NH_4^{+8} were completely optimized and $C_2H_4⋅NH_4^+$ was partially optimized at the 3-21G level. Only calculations of the third type, i.e., 3-21G//3-21G, were carried out for $C_2H_5NH_3^+$.
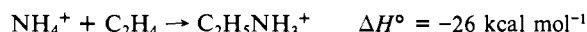
Geometry optimizations were done by the force relaxation method.[12] Reported bond lengths represent convergence to 0.001 Å and bond angles to 0.1°.

Results and Discussion

1. **The Thermochemistry of Condensation vs. Clustering Reactions.** In principle, attachment of onium ions to unsaturated hydrocarbons could yield either noncovalently bonded cluster ions or covalent condensation products. The thermochemistry can be obtained or estimated from ion and neutral heat-of-formation data.[13] The condensation reaction for $NH_4^+ + C_2H_4$ is strongly exothermic.

$$NH_4^+ + C_2H_4 \rightarrow NH_4⋅C_2H_4$$

$$\Delta H°(\text{exptl}) = -10 \text{ kcal mol}^{-1}$$
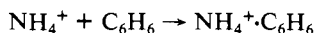
$$NH_4^+ + C_2H_4 \rightarrow C_2H_5NH_3^+ \qquad \Delta H° = -26 \text{ kcal mol}^{-1}$$

Similar results on the thermochemistry are obtained by ab initio calculations (see below).

Experimentally, in a system of trace NH_3 in C_2H_4 at temperatures above ~320 K we observed the irreversible buildup of m/z 46 (i.e., probably $C_2H_5NH_3^+$) and its clusters with NH_3. However, at ~290 K and below the relative intensities of m/z 18 (NH_4^+) and of the adduct ion at m/z 46 appeared to reach a constant ratio, suggesting reversible clustering. This observation is tentative, however, since both ions reacted further under these conditions to form higher clusters with NH_3. From these observations we conclude tentatively that NH_4^+ condenses with C_2H_4 to form $C_2H_5NH_3^+$ as the thermochemistry suggests. But the condensation proceeds only at elevated temperatures and thus seems to involve an activation energy, which would be expected considering the requisite electronic rearrangement. Irreversible condensation was also observed in the reaction of $CH_3NH_3^+$ with C_2H_4.

In contrast, the thermochemistry of onium ion–aromatic hydrocarbon systems favors noncovalent clustering.[14]

$$NH_4^+ + C_6H_6 \rightarrow NH_4^+⋅C_6H_6$$

$$\Delta H°(\text{exptl}) = -19.3 \text{ kcal mol}^{-1}$$

$$NH_4^+ + C_6H_6 \rightarrow c\text{-}C_6H_7NH_3^+ \qquad \Delta H° \approx 8 \text{ kcal mol}^{-1}$$
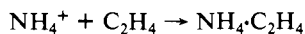
The condensation reaction is unfavorable since it eliminates the aromatic stabilization of benzene. In agreement with the thermochemistry, the association reactions of onium ions with aromatic

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(14) For the thermochemistry of the condensation reaction we estimate $\Delta H°_f(c\text{-}C_6H_7NH_2) = \Delta H°_f(c\text{-}C_6H_{11}NH_2) + \Delta H°_f(1,3\text{-}c\text{-}C_6H_8) - \Delta H°_f(c\text{-}C_6H_{12}) = 33$ kcal mol⁻¹; $PA(c\text{-}C_6H_7NH_2) = PA(c\text{-}C_6H_{11}NH_2) + PA(CH_2CHCH_2NH_2) - PA(n\text{-}C_3H_7NH_2) = 218$ kcal mol⁻¹.

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ligands were reversible under all conditions, suggesting clustering rather than condensation.

2. Complexes of Onium Ions with π-Donors. The ammonium ion forms a strong complex with benzene. The interaction may involve charge transfer between the ion and the π-system or, alternatively, only electrostatic forces may be significant.

The $K^+\cdot C_6H_6$ complex may be a model for the electrostatic factors in $NH_4^+\cdot C_6H_6$, since the ionic radii are similar (1.33 and 1.43 Å, respectively). Sunner et al.[4] found that K^+ is positioned axially (analogous to **1**), aligned with the benzene quadrupole, which results from the π(−) and carbon nuclear charges (+) system. They also found that the main electrostatic factors are


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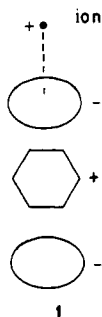
the ion-quadrupole, ion-induced dipole, dispersion, and electronic repulsion energies leading to the interaction energy $E = \Delta H^\circ_D = 19.2$ kcal mol^{-1}. The $NH_4^+\cdot C_6H_6$ interaction energy is similar which suggests that this complex is bound by similar forces. It is interesting that the homogeneous charge distribution of K^+ vs. the tetrahedral, localized fractional charges of NH_4^+ does not seem to make a significant difference. Indeed, this also applies to higher clusters of $NH_4^+\cdot nL$, where L is C_6H_6 or polar ligands, as will be shown elsewhere.

As may be expected from electrostatic considerations, ΔH°_D decreases as the fractional charge decreases, in going from NH_4^+ to $MeNH_3^+$ and Me_3NH^+, and increases with increasing polarizability of the ligand, in going from C_6H_6 to $1,3,5\text{-}(CH_3)_3C_6H_3$. However, this does not prove the electrostatic nature of the bond, since the same trend would be expected also for $H^+\cdots\pi$ charge delocalization energies. The weak interaction energies when ethylene, cyclohexene, and cyclohexane replace C_6H_6 as the ligand are also consistent with either electrostatic ion-quadrupole or $NH^+\cdots\pi$ delocalization considerations.

A somewhat more complex but revealing effect is observed with the fluorinated benzene ligands. With regard to $NH^+\cdots\pi$ charge transfer, we note that the ionization potentials of C_6H_5F and $1,4\text{-}C_6H_4F_2$ are similar to that of C_6H_6, and the $NH^+\cdots\pi$ interactions should be comparable in C_6H_6 vs. the fluorinated species. In contrast, electrostatic interactions should be weakened, since the dipole in C_6H_5F and the F-ring-F quadrupole in $1,4\text{-}C_6H_4F_2$ induce positive charge toward the center of the ring making **1** less stable. Experimentally, fluorination substantially destabilizes the complexes, again suggesting the electrostatic forces predominate (reaction 1 vs. reactions 3 and 4, Table I).

Indeed, with fluorinated benzenes the structure of the complex may be reoriented due to the partial negative charge on the fluorine atoms (**4**). ΔH°_f of $NH_4^+\cdot C_6H_5F$, is similar in magnitude to that

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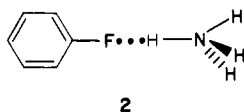
of the analogous complex $MeNH_3^+\cdot MeF$ (11.8 kcal mol^{-1}).[2] Similar considerations may apply also to $MeNH_3^+\cdot$pyrrole, where both the axial π-complex and the equatorial $MeNH_3^+\cdots NC_4H_4$ analogue of **2** are possible. Ab initio results on the stabilities of the benzene complexes will be presented below.

3. Ab Initio Calculations: $NH_4^+\cdot C_2H_4$. Additional support for the predominantly electrostatic nature of the $XH^+\cdots\pi$ complexes is provided by ab initio calculations. Two conformations of the $NH_4^+\cdot C_2H_4$ complex were investigated. The conformers differ

Table I. Thermochemistry of Dissociation of Complexes of Onium Ions AH$^+$ with π-Donors

AH$^+$	B	$\Delta H^\circ_D{}^a$	$\Delta S^\circ_D{}^a$	ΔG°_D $(T)^a$
1. NH$_4^+$	C$_6$H$_6$	19.3	23.3	
2.	c-C$_6$H$_{12}$	$<9^b$	(20)	≤ 2.8 (317)
3.	C$_6$H$_5$F	14.4	18.0	
4.	1,4-C$_6$H$_4$F$_2$	13^b	(20)	5.1 (395)
5.	1,3,5-(CH$_3$)$_3$C$_6$H$_3$	21.8	21.2	
6.	C$_2$H$_4$	10^b	(20)	3.7 (294)
7. CH$_3$NH$_3^+$	C$_6$H$_6$	18.8	25.1	
8.	c-C$_6$H$_{10}$	11.6	16.9	
9.	pyrrole	18.6	21.0	
10. (CH$_3$)$_3$NH$^+$	C$_6$H$_6$	15.9	27.7	
11. C$_2$H$_5$OH$_2^+$	C$_6$H$_6$	21^b	(25)	8.7 (491)

^{a} ΔH°, ΔG° in kcal mol^{-1}, ΔS° in cal mol^{-1} K^{-1}, T in K. Error estimates: ΔH° ± 1 kcal mol^{-1}; ΔS° ± 2 cal mol^{-1} K^{-1}. ^{b} Obtained from ΔG° measured at one temperature and ΔS° as estimated from analogous reactions.

in the orientation of the hydrogens of the NH_4^+ group with respect to the double bond of C_2H_4. In conformation **3a** one of the hydrogens points directly at the double bond, the others point away from it. In conformation **3b** two hydrogens are astride of the double bond and two are directed away from it. For both complexes the NH_4^+ nitrogen was positioned above the double bond midway between the two carbons and the distance between the nitrogen and the double bond (R_{X-N}) was optimized at the STO-3G and 3-21G basis set levels. The geometries of NH_4^+ and C_2H_4 are from ref 8.

3a   **3b**
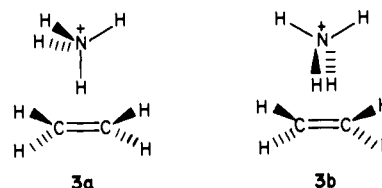
In order to compare the stabilization energies of the π-complexes with those of the covalently bonded protonated ethylamine, the structure of $C_2H_5NH_3^+$ was completely optimized. The geometrical parameters are given in the drawing below.
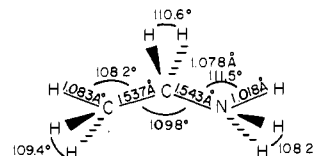

STO-3G//STO-3G, 3-21G//STO-3G, and 3-21G//3-21G total energies of NH_4^+, C_2H_4, $C_2H_5NH_3^+$(3-21G//3-21G only), conformer **3a**, and conformer **3b** and stabilization energies of $C_2H_5NH_3^+$(3-21G//3-21G only), **3a**, and **3b** are give in Table II. The table includes STO-3G and 3-21G optimized X–N bond distances also. A comparison of the data presented in Table II shows the following. (1) In general the 3-21G stabilization energies are closer to the experimental values than the STO-3G energies are and the optimized 3-21G X–N bond lengths are longer by 0.3 Å (**3a**) and 0.1 Å (**3b**). The former observation agrees with that of Binkley et al.[10b] from their work on hydrogenation reactions. (2) The 3-21G//STO-3G and 3-21G//3-21G total energies, stabilization energies, and relative energies are very similar, indicating that the 3-21G//3-21G results for these complexes are reproduced quite well by the 3-21G//STO-3G calculations. The maximum difference in the two sets of energies is about 3 kcal and corresponds to the 0.3-Å variation in X–N bond length. (3) **3a**, i.e., the conformer with the hydrogen pointing directly at the C_2H_4 double bond, is the more stable complex. However, **3a** is approximately 21 kcal less stable than $C_2H_5NH_3^+$. The latter result is in reasonable agreement with the experimental value.

Charge distributions from Mulliken population analysis[15] are given in Figure 2 for the monomers, dimers, and $C_2H_5NH_3^+$. Only

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Table II. Structures, Total Energies (E_T), and Stabilization Energies (ΔE) of $NH_4 \cdot C_2H_4$ Complexes

molecule	calculational level	$E_T{}^a$	$-\Delta E^b$	$R_{X\text{-}N}{}^c$
$NH_4{}^+$	STO-3G//STO-3G^d	-55.86885		
	3-21G//STO-3G	-56.23232		
	3-21G//3-21G^d	-56.23386		
C_2H_4	STO-3G//STO-3G^d	-77.07395		
	3-21G//STO-3G	-77.60069		
	3-21G//3-21G^d	-77.60099		
$C_2H_4 \cdot NH_4{}^+$, 3a	STO-3G//STO-3G	-132.95736	9.14	2.969
	3-21G//STO-3G	-133.84724	8.93	2.969
	3-21G//3-21G	-133.85224	10.91	3.255
$C_2H_4 \cdot NH_4{}^+$, 3b	STO-3G//STO-3G	-132.94736	2.86	3.257
	3-21G//STO-3G	-133.84340	6.52	3.257
	3-21G//3-21G	-133.84564	6.77	3.348
$C_2H_5NH_3{}^+$	3-21G//3-21G	-133.88511	31.54	

^{a} Energies in atomic units. ^{b} Energies in kcal. ^{c} Bond lengths in Å. $R_{X\text{-}N}$ is the distance between the $NH_4{}^+$ nitrogen and the center of the C–C double bond. ^{d} Reference 8.
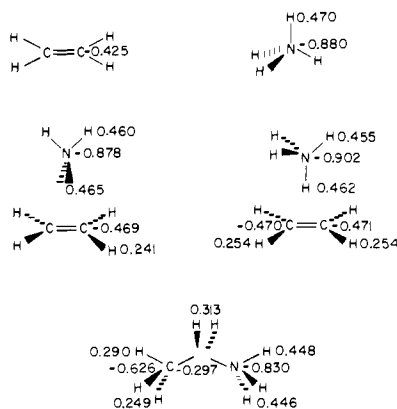
Figure 2. Atomic charges from population analysis for $NH_4{}^+$, C_2H_4, $NH_4{}^+ \cdot C_2H_4$, and $C_2H_5NH_3{}^+$.

the 3-21G//3-21G atomic charges have been included in the figure, since analysis of all three sets of charge distributions leads to the same conclusions. The results support the suggestion that the stabilization of the dimers arises primarily from electrostatic effects. The figure shows that there is little charge transfer from ethylene to $NH_4{}^+$ in either complex and that forming the complexes produces a moderate amount of charge redistribution within C_2H_4. Overall the ethylene carbons gain negative charge in the complexes, whereas the ethylene hydrogens lose negative charge. However, examination of the electron density in the individual atomic orbitals of the carbon atoms shows that the carbon π-orbitals lose electron density in the dimers as expected. The net gain of electron density on the carbons is due to the transfer of electrons from the ethylene hydrogens to the C_{2s} and C_{2p_z} atomic orbitals, i.e., the p orbitals in the plane of the ethylene molecule perpendicular to the C=C bond. The charge transferred to $NH_4{}^+$ is distributed such that each atom in $NH_4{}^+$ gains negative charge upon complex formation, with the exception of the nitrogen in 3b. Most of the charge density is transferred to the hydrogens that point away from the double bond. Except for the hydrogen(s) directed toward the C=C bond, these charge changes are characteristic of hydrogen-bonded complexes.[11] The observed gain in negative charge on the hydrogen(s) that interact(s) with the double bond is unusual in stable hydrogen-bonded complexes.[11,16]

$NH_4{}^+$ polarizes the electron density on C_2H_4 more extensively in 3a than in 3b, which partially accounts for the lower energy of 3a. The difference in polarization is not obvious from the atomic charges, however, since the bigger charge gains in the C_{2s} (0.041 e vs. 0.026 e) and C_{2p_z} (0.049 e vs. 0.036 e) orbitals on 3a are compensated for by the smaller charge loss in the C_π orbitals on 3b (-0.041 e vs. -0.013 e). Other reasons 3a is more stable than 3b include the following. First, from their electrostatic potential

(16) Del Bene, J. E.; Marchese, F. T. *J. Chem. Phys.* 1973, 58, 926.

calculations on $C_2H_2 \cdot HF$ and $C_2H_4 \cdot HF$, Kollman et al.[17] have shown that the most favorable line of approach of the hydrogen-bonded proton is along the center of the π-bond. Second, the charge transfer from C_2H_4 to $NH_4{}^+$ is greater for 3a than for 3b. Third, the overlap between the ethylene carbons and the hydrogen(s) directed toward the double bond is much larger for 3a than for 3b (0.0618 vs. 0.0109). This result is due to the shorter distance between the ethylene carbons and the pertinent NH_4 hydrogen(s) (2.328 Å vs. 2.956 Å), the more extensive mixing of the $NH_4{}^+$ and C_2H_4 atomic orbitals, and the larger charge transfer in 3a.

As in other donor-acceptor systems,[11,17] the electron donor, C_2H_4, molecular orbitals (MO's) are all stabilized in 3a and 3b and the proton donor, $NH_4{}^+$, MO's are all destabilized in the complexes. For both conformers the electron pair donor shifts dominate the electron pair acceptor shifts. The increases and decreases in the MO energies and, therefore, the average absolute value of the energy changes are larger for 3a than for 3b. This is also indicative of the stronger interaction between $NH_4{}^+$ and C_2H_4 in the former complex.[18]

4. Ab Initio Calculations: $NH_4{}^+ \cdot C_6H_6$ and $NH_4{}^+ \cdot C_6H_5F$. Four conformations of $C_6H_6 \cdot NH_4{}^+$ were investigated. For each $C_6H_6 \cdot NH_4{}^+$ conformer considered, the analogous $C_6H_5F \cdot NH_4{}^+$ complex was examined also. In addition to the above complexes for $C_6H_5F \cdot NH_4{}^+$, where $NH_4{}^+$ interacts with the π-system, the dimer (2) with $NH_4{}^+$ interacting with the fluorine was studied. Three conformations of $C_6H_6 \cdot NH_4{}^+$ and $C_6H_5F \cdot NH_4{}^+$ have the nitrogen atom on the C_6 axis. They differ from each other with respect to the number of hydrogens, one (4a), two (4b), or three (4c), directed toward the benzene ring. The fourth conformer (4d) has the nitrogen positioned above the center of one of the C–C bonds. On the basis of the $C_2H_4 \cdot NH_4{}^+$ results (section 3),

4b   4b'
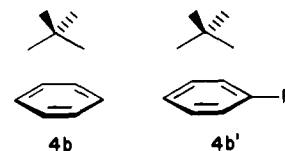
only the case where one hydrogen points directly at the C–C bond was considered here. Although the C–C bonds in $C_6H_5F \cdot NH_4{}^+$ are not all equivalent, the energies of the three possible 4d' complexes are within 1 kcal of each other. Thus, only the data for the most stable complex ($NH_4{}^+$ interacting with C_2–C_3 or C_5–C_6) will be reported.

For dimers 4a–d and 4a'–d' the distance between the nitrogen and the center of the ring or C–C bond ($R_{X\text{-}N}$) was optimized. It was found to make no difference in the total energy whether the hydrogens interacting with the ring in 4b and 4c point toward the carbons or between them. For complex 2 the hydrogen bond was assumed to be linear and only the distance between the fluorine and hydrogen-bonded proton ($R_{F\text{-}H}$) was optimized. The geometry of C_6H_5F is from ref 9. The structures of $NH_4{}^+$ and C_6H_6 are from ref 8.

Table III tabulates the STO-3G total energies, stabilization energies, and optimum X–N bond distances as well as the 3-21G//STO-3G total energies and stabilization energies of the complexes. The total energies of $NH_4{}^+$, C_6H_6, and C_6H_5F are included also for reference. The table shows that the most stable $C_6H_6 \cdot NH_4{}^+$ and $C_6H_5F \cdot NH_4{}^+$ dimers differ for the STO-3G (4d and 4d') and 3-21G//STO-3G (4b and 4b') calculations. Since the 3-21G basis set has been found to yield more reliable relative molecular energies than the STO-3G basis,[10b] the 3-21G//STO-3G results will be used in this article. However, based on the comparison of 3-21G//3-21G and 3-21G//STO-3G data given in the preceding section, the total energies of the complexes are sufficiently close that the relative energies could be reordered after optimization at the 3-21G level or higher levels. The close agreement among the total energies of the various conformers of

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(18) Kollman, P. A.; Allen, L. C. *J. Am. Chem. Soc.* 1970, 92, 6101.

Table III. Structures, Total Energies (E_T), and Stabilization Energies (ΔE) of $C_6H_6 \cdot NH_4^+$ and $C_6H_5F \cdot NH_4^+$

molecule	calculational level	$E_T{}^a$	$-\Delta E^b$	$R_{X-N}{}^c$
NH_4^+	STO-3G//STO-3G^d	−55.86885		
	3-21G//STO-3G	−56.23232		
C_6H_6	STO-3G//STO-3G^d	−227.89136		
	3-21G//STO-3G	−229.41896		
C_6H_5F	STO-3G//STO-3G^e	−325.35111		
	3-21G//STO-3G	−327.74265		
4a	STO-3G//STO-3G	−283.77117	6.88	3.104
	3-21G//STO-3G	−285.67545	15.17	3.104
4b	STO-3G//STO-3G	−283.77212	7.48	2.911
	3-21G//STO-3G	−285.67719	16.26	2.911
4c	STO-3G//STO-3G	−283.76946	5.81	2.992
	3-21G//STO-3G	−285.67464	14.66	2.992
4d	STO-3G//STO-3G	−283.77547	9.58	2.981
	3-21G//STO-3G	−285.67393	14.21	2.981
4a′	STO-3G//STO-3G	−381.22732	4.62	3.151
	3-21G//STO-3G	−383.99249	11.00	3.151
4b′	STO-3G//STO-3G	−381.22816	5.15	2.948
	3-21G//STO-3G	−383.99305	11.35	2.948
4c′	STO-3G//STO-3G	−381.22564	3.56	3.061
	3-21G//STO-3G	−383.99149	10.37	3.061
4d′	STO-3G//STO-3G	−381.23179	7.43	3.004
	3-21G//STO-3G	−383.99141	10.32	3.004
2	STO-3G//STO-3G	−381.25111	19.55	2.442
	3-21G//STO-3G	−384.00166	16.75	2.442

^{a} Energies in atomic units. ^{b} Energies in kcal. ^{c} Bond lengths in Å. R_{X-N} is the distance between the nitrogen and the center of the ring, the center of a C–C bond, or the fluorine. ^{d} Reference 8. ^{e} Reference 9.

4 and 4′ indicates that the orientation of NH_4^+ above the ring does not make a significant difference in the stability of the complexes. This supports the suggestion that similar forces are involved in the interactions between K^+ and C_6H_6 and NH_4^+ and C_6H_6.

The data given in Table III also demonstrate that (1) again the 3-21G//STO-3G stabilization energies are in better accord with the experimental values than the STO-3G energies are, (2) the stabilization energies of the π-complexes are uniformly lower than the $\Delta H°_D$ values, (3) the 3-21G//STO-3G relative stabilization energies of the $C_6H_6 \cdot NH_4^+$ complexes compared to the analogous $C_6H_5F \cdot NH_4^+$ complexes (i.e., 4b vs. 4b′) are in excellent agreement with the experimental difference of ∼5 kcal, and (4) complex 2 is calculated to be the most stable $C_6H_5F \cdot NH_4^+$ dimer but its computed stabilization energy is too high. Other researchers have also obtained overestimated stabilization energies for F⋯H–A dimers studied by using split-valence basis sets.[11,17]

Figure 3 displays the 3-21G//STO-3G atomic charges from population analysis[15] for NH_4^+, C_6H_6, C_6H_5F, 2, 4a–d, and 4a′–d′. The amount of charge transferred from C_6H_6 to NH_4^+ is 0.033, 0.066, 0.047, and 0.057 e for 4a, 4b, 4c, and 4d, respectively. The amount of charge transferred from C_6H_5F to NH_4^+ decreases slightly for each corresponding complex. It is 0.032, 0.058, 0.041, and 0.051 e for 4a′, 4b′, 4c′, and 4d′, respectively. For complex 2, 0.089 e is donated to NH_4^+. The additional electron density is distributed primarily to the nitrogen and to the hydrogens directed away from the ring.

Charge is also redistributed within C_6H_6 and C_6H_5F, from the ring hydrogens to the C_{2s} and C_{2p_z} orbitals. One measure of this polarization is the total gain in σ-electron density on the carbons. The gross carbon σ-electron density increases by 0.225, 0.218, 0.211, and 0.201 e for 4a, 4b, 4c, and 4d, respectively. The gains for 4a′, 4b′, 4c′, and 4d′ are 0.197, 0.190, 0.176, and 0.173 e, respectively, again showing a slight drop.

The charge density is polarized much less in the proton donor, NH_4^+, upon π-complex formation. In fact, only complexes 4a, 4a′, 4d, and 4d′ exhibit a gain in positive charge on the hydrogen-bonded proton. For 4b and 4b′ there is no change in electron density on the hydrogens interacting with the ring, and for 4c and 4c′ these hydrogens become more negatively charged. As stated previously, the latter result is unusual for stable hydrogen-bonded dimers.[11,16]
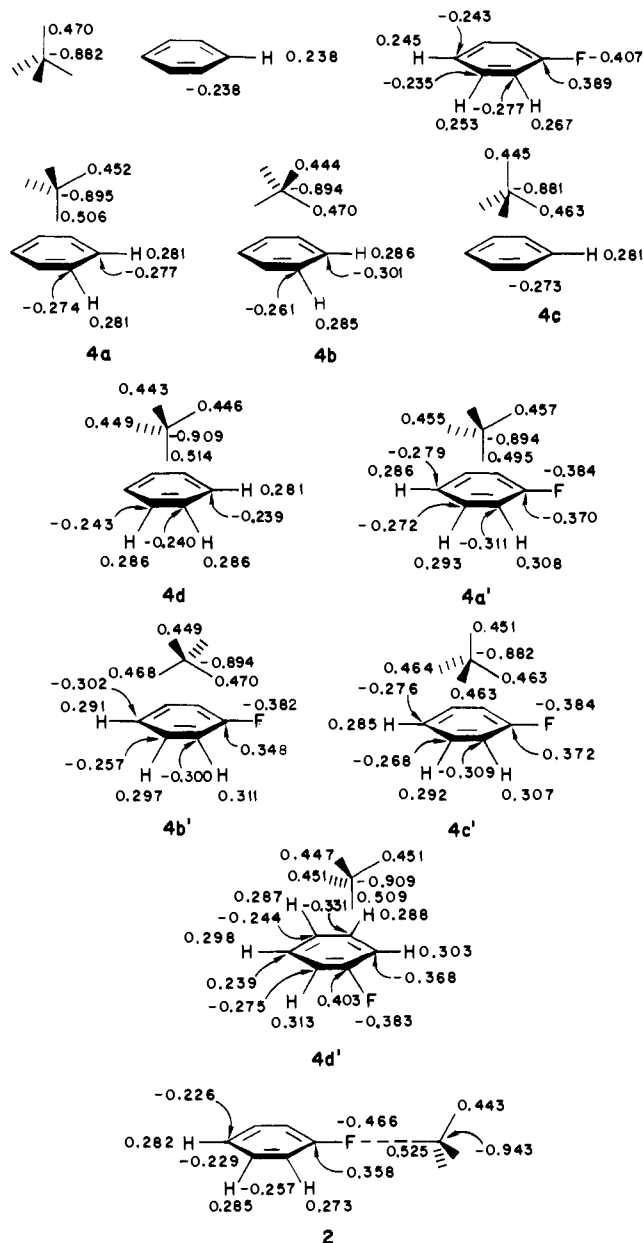

Figure 3. Atomic charges from population analysis for NH_4^+, C_6H_6, C_6H_5F, 2, 4a–d, and 4a′–d′.

All of the charge redistributions obtained for 2 are characteristic of hydrogen-bonded complexes. The hydrogen-bonded proton, carbons (except C_1), and C_6H_5F hydrogens lose electrons. The nitrogen, remaining NH_4^+ hydrogens, C_1, and F gain electrons.

The stabilities of the $C_6H_6 \cdot NH_4^+$ dimers fall in the order 4b > 4a > 4c > 4d, although the energies of 4a, 4c, and 4d are very close. Comparing the actual order of stabilities to the orders predicted by the charge-transfer data, polarization data, MO energy shifts, and X–N distances leads to some interesting results. On the basis of the X–N distances and the amount of charge transfer, the expected ranking would be 4b > 4d > 4c > 4a. The arrangement from most stable to least stable on the basis of polarization would be 4a > 4b > 4c > 4d. The average absolute values of the MO energy shifts of the donor and acceptor orbitals lead to the same predicted order as the X–N distances and charge-transfer data, 4b (0.1315 au) > 4d (0.1274 au) > 4c (0.1247 au) ≈ 4a (0.1244 au). Apparently electrostatic and exchange repulsion effects play an important role in determining the relative stabilization energies of these π-complexes as well as in their overall stabilization energies.[11,19] Note, however, that 4b is positioned first in four out of the five categories. The low relative stability of 4d can be explained by the observation of Kollman

et al.[17] on C₆H₆·HF that the electrostatic potential is most negative along the C_6 axis. Of course, since the differences in the above data are generally quite small, the trends could change after reoptimization at the 3-21G or higher basis set level. Nevertheless, these trends do seem to follow the patterns found for other donor–acceptor complexes.[11,19]

Identical rankings are obtained for **4a′–d′** in each of the aforementioned categories.[20] Thus, the same conclusions apply to the C₆H₅F·NH₄⁺ π-dimers. Of course, the calculations predict that **2** is the most stable C₆H₅F·NH₄⁺ complex. However, since the computed total energies of the π-dimers are uniformly too low and the computed total energy of **2** is too high, it is probable that the deviations in their energies are smaller than the values found in this work. Consequently, more accurate calculations are required to distinguish between them, although the results of other researchers would suggest that **2** is the lowest energy form of C₆H₅F·NH₄⁺.[11,19]

(19) Yamabe, S.; Morokuma, K. *J. Am. Chem. Soc.* **1975,** *97,* 4458.
(20) The average absolute values of the MO energy changes for the C₆H₅F·NH₄⁺ π-dimers are **4a′,** 0.1226; **4b′,** 0.1295; **4c′,** 0.1236; **4d′,** 0.1258.

The decrease in ΔE observed for the C₆H₅F·NH₄⁺ π-complexes compared to the C₆H₆·NH₄⁺ complexes can be accounted for as follows. The highest occupied molecular orbitals in C₆H₆ are a pair of degenerate π-orbitals. Substituting a hydrogen with a fluorine stabilizes these MO's by 0.00496 and 0.0182 au, respectively. This makes the π-electrons on C₆H₅F more tightly bound than those on C₆H₆ (primarily due to the second energy lowering) and leads to weaker electrostatic, polarization, and charge-transfer interactions for C₆H₅F.[17] The diminutions in the latter two interactions were noted above.

If C₆H₅F·NH₄⁺ is π-complex, the reduced availability of the π-electrons explains the smaller $\Delta H°_D$ value for C₆H₅F·NH₄⁺ than for C₆H₆·NH₄⁺. In contrast, if C₆H₅F·NH₄⁺ is a σ-complex, more accurate calculations are needed to explain the smaller $\Delta H°_D$.

Mechanism of the Oxidation of NADH by Quinones. Energetics of One-Electron and Hydride Routes

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Abstract: The kinetics of NADH oxidation by 7 *o*-benzoquinones and 14 *p*-benzoquinones were studied by using buffered aqueous solutions and UV/vis spectroscopy. For each quinone the rate law was first order in NADH and first order in quinone. The rate constants varied from 0.0745 to 9220 M⁻¹ s⁻¹. Variation of the pH from 6 to 8 gave no change in rate. The use of 4-D- and 4,4-D₂-NADH revealed kinetic isotope effects. The dideuterio data gave k_H/k_D in the range 1.6–3.1 for *p*-quinones and 4.2 for 3,5-di-*tert*-butyl-*o*-quinone. When *p*-quinones were used, the log k was a linear function of $E°$ for the quinone/hydroquinone monoanion (Q/QH⁻) couple with a slope of 16.9 V⁻¹. *o*-Quinones reacted about 100 times more rapidly, but the same linear relationship with a slope of 16.4 V⁻¹ was observed. Comparisons to data for one-electron-transfer reactions indicate that such mechanisms are not involved. A hydride-transfer mechanism accommodates all the data, and rate-limiting hydrogen atom transfer followed by electron transfer cannot be ruled out.

The mechanisms by which the coenzyme NADH is oxidized to NAD⁺ by chemical oxidants stand at a curious intersection of several scientific frontiers. These mechanisms are obviously interesting to enzymologists and biochemists.[1] In addition they provide an instructive case with regard to the currently fashionable investigation of the intercession of single-electron transfer in organic reactions.[2] Indeed, previous work has provided some mechanistic controversy.[3] Finally studies of NADH electrochemistry[4] have indicated a need for further mechanistic understanding, including a careful treatment of the classical one-electron vs. two-electron dichotomy.

In this paper we elucidate the chemistry of benzoquinones with NADH. These reactions were of interest to us because of previous
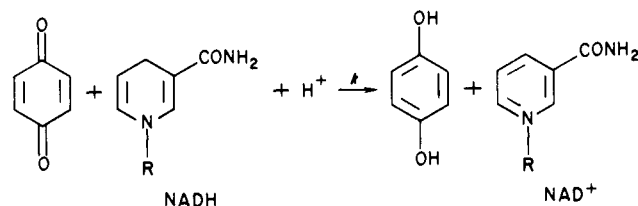
electrochemical studies which showed that *o*-quinones were efficient mediators for the electrooxidation of NADH,[5] i.e., the quinone catalyzed the electrooxidation as shown in Figure 1. Our long-term goal has been to design a chemically modified electrode which would use a surface-attached quinone-type mediator to effectively catalyze NADH oxidation at potentials near the NADH/NAD⁺ $E°$ (−310 mV, NHE). We reasoned that a

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