Hydrogen bond acceptor capability of carbonyl π-electrons—case study of the hydrogen-bonded urea dimer

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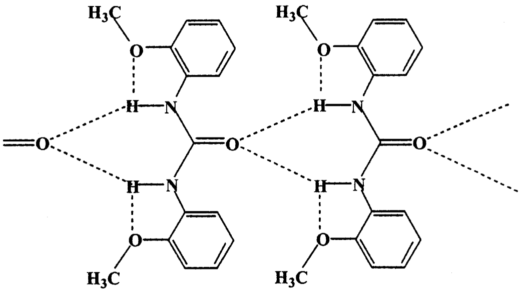
Due to hydrogen bonding interactions, the urea–urea dimer has two relatively stable rotamers, having dihedral angles of either 0° (coplanar) or 90° (perpendicular). These two rotamers may be respectively explained by contributions to the bridging hydrogen bonds by either electrons from the nonbonding orbital or from the π orbital of oxygen. We find that the π-donor substituents can enhance effectively the hydrogen bond acceptor capability for the carbonyl π-electrons. The urea dimer has nearly equal bonding energies for the two rotamers with $\theta = 0°$ and $\theta = 90°$. For the dimer series R_2CO·urea (R = H, F, Cl, NH_2, NHCH), the energy difference between the two conformers can be accounted for by the HOMO–LUMO type non-Coulombic interaction.

An interesting feature of the urea crystal is that a given oxygen atom participates in a total of four hydrogen bonds with neighboring molecules.[1–5] Both of the nonbonding electron pairs on oxygen can act as ligands to hydrogen atoms of a neighboring coplanar urea molecule. It is further bonded through carbonyl π-electrons with two other hydrogen atoms from neighboring molecular chains with anti-parallel axes, one above and one below. The planes of the two neighboring chains are perpendicular to that of the main chain, as shown in **1**. Probably better evidence for the hydrogen bond acceptor capability of π-electrons of the urea molecule is the crystal structure of the sterically bulky urea analog, 1,3-bis(*o*-anisoyl)urea.[6] It is bidentate rather than tetradentate through either the in-plane nonbonding orbital, structure **2**, or the π orbital, structure **3**. The two structures are here referred to as the $\theta = 0°$ and $\theta = 90°$ rotamers, respectively.



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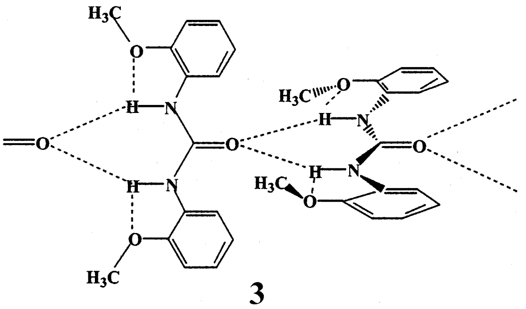

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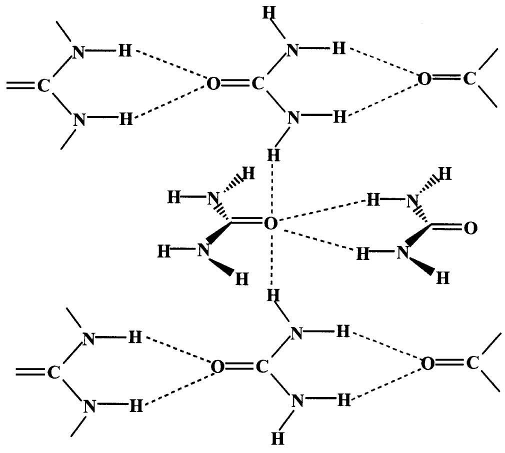


1

These interesting π-directional hydrogen bonding structures prompted us to perform some theoretical study on the dihedral angle (θ) dependence of hydrogen bond strength for the urea dimer. The results of $\theta = 0°$, 30°, 60°, 90° are given in Table 1, as determined by the GAUSSIAN 98 program[7] with the theoretical method B3LYP/6-31+G**.[8] The optimized geometry is constrained with both monomers to be planar with local C_{2v} symmetry within the dimer. We are concerned mostly with the relative energy around the dihedral angle θ; therefore, the BSSE correction has not been considered extensively. We have checked the BSSE values for the 0° and 90° rotamers of the urea dimer. The corresponding values are 0.400 and 0.478 kcal mol^{-1} with ΔBSSE = 0.078 being much smaller than the ΔE of 0.421 kcal mol^{-1} between the two rotamers. The energy curve is nearly flat with a maximum difference of only 0.5 kcal mol^{-1} *vs.* the total binding energy of about 7 kcal mol^{-1}. The Mulliken charges $q(O_1, H_1, H_2)$ refer to the atoms involved in the hydrogen bonds. In Table 2, a comparison is given between the current method and previous theoretical and experimental results for some properties of the urea

Table 1 Hydrogen bonding energies, distances and Mulliken charges for the urea dimers from B3LYP/6-31 + G** level calculations[a]

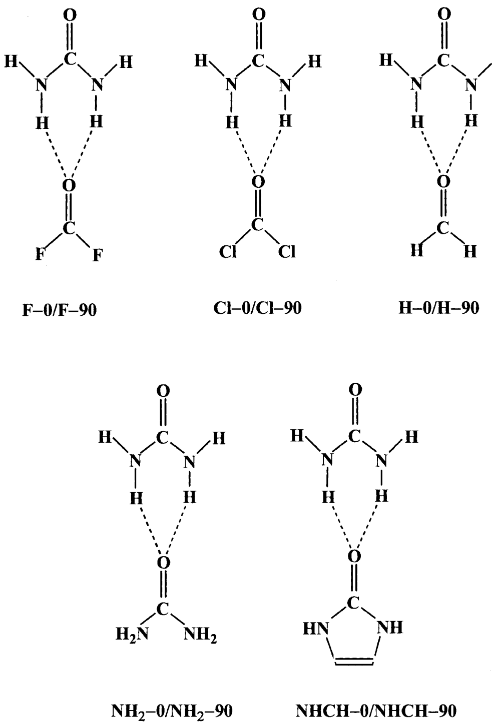
$\theta/°$	E_{HB}[b] /kcal mol^{-1}	R^c(O···H)/Å	q(O$_1$, H$_2$, H$_3$)[d]
0	−7.476	2.160	(−0.65, 0.34, 0.34)
30	−6.929	2.152	(−0.68, 0.34, 0.34)
60	−7.030	2.153	(−0.67, 0.34, 0.34)
90	−7.055	2.150	(−0.67, 0.34, 0.34)

[a] θ is the dihedral angle between two monomer planes. The monomer is constrained to be planar with C_{2v} symmetry. [b] $E_{HB} = E[(\text{urea})_2] - 2 \times E[\text{urea}]$. [c] R(O···H) is the intermolecular hydrogen bond distance between the carbonyl oxygen and amino hydrogen. [d] Mulliken charge.

molecule. The present theoretical results compare reasonably well with the experimental results. From the studies of isotopomers of urea by microwave spectroscopy, the geometrical parameters derived indicate the molecule is almost planar with the amino groups somewhat pyramidal. Therefore, the molecule has actually either C_2 or C_s symmetry. In our calculation it is constrained to have C_{2v} local symmetry in the bonded dimer for simplicity.

Substituent effects

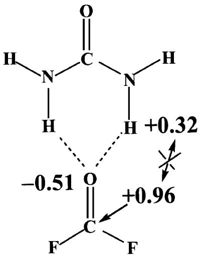
The main focus of this study is the strong π-directional hydrogen bond as shown in Table 1 for E_{HB} at 90°. It raises two interesting questions — Is the flat potential surface along the dihedral angle θ unique to the urea dimer or does it exist for other substituted carbonyl dimers as well? — What effect do various substituents have on the carbonyl π-electron hydrogen bonding capability? To answer these questions, we carried out calculations on the hydrogen bond dimer series R$_2$CO·urea (R = H, F, Cl, NH$_2$, NHCH) as shown in Scheme 1 for both $\theta = 0°$ and $\theta = 90°$ rotamers. The results are shown in Table 3. We found that variation of the substituent group has a significant effect on the hydrogen bonding energies, which range from −1.6 kcal mol^{-1} for the **F–90** rotamer to −8.2 kcal mol^{-1} for the **NHCH–90** rotamer. For any given substituent, the energy difference between $\theta = 0°$ and $\theta = 90°$ rotamers is not great, the biggest difference being 1.1 kcal mol^{-1} for the formaldehyde·urea rotamers (**H–0** and **H–90**). The most interesting result in Table 3 is probably that the energy difference between the 0° and 90° rotamers has a different sign for the R = NHCH system than for the others (the 90° rotamer is the most stable in this one case). This can be attributed to the fact that this substituent is the strongest π-donor studied here. It will be discussed in more detail later in Fig. 2.



Scheme 1

In the last column of Table 3, there is also an important variation of the charge on the carbonyl oxygen with different substituents. It appears that the average hydrogen bonding energies ($\bar{E}_{HB}$) and the Mulliken charges on the oxygen [$\bar{q}$(O)] correlate well with the exception of the F-substituted compound: $-\bar{E}_{HB}$ (kcal mol^{-1}) = 2.019 (**Cl**) < 4.100 (**H**) < 7.266 (**NH$_2$**) < 8.063 (**NHCH**); $-\bar{q}$(O) = 0.23 (**Cl**) < 0.36 (**H**) < 0.66 (**NH$_2$**) < 0.72 (**NHCH**). Here, both $\bar{E}_{HB}$ and $\bar{q}$(O) are the average values of the values at $\theta = 0°$ and $\theta = 90°$, which are close. The fluorine-substituted compound has a large average $\bar{q}$(O) of −0.51 and a small average $\bar{E}_{HB}$ of −1.85 kcal mol^{-1}. This is because the carbon is highly positively charged with q(C) = 0.96 as shown in Scheme 2. From the dipole moment of the fluorine-substituted compound, we obtain an effective oxygen charge[12] $\bar{q}_{\text{eff}}$(O) = −0.087.

Fig. 1 shows good correlation between $\bar{E}_{HB}$ values and $\bar{q}$(O) values with a correlation coefficient 0.9724. Another interesting result from Table 3 is the good correlation between E_{HB} and R(O···H), the hydrogen bond distance. The average distances for the $\theta = 0°$ and $\theta = 90°$ rotamers are 2.413 (**F**) > 2.386 (**Cl**) > 2.309 (**H**) > 2.155 (**NH$_2$**) > 2.142 (**NHCH**).



Scheme 2

Table 2 A comparison between the present results and previous theoretical and experimental results for the urea monomer

	B3LYP/6-31+G**[a]	HF/6-311++G(d,p)[b]	MP2/6-311++G(d,p)[b]	B3PW91/D95++**[c]	Exptl[b]
d(C=O)/Å	1.225	—	1.218	1.226	1.221
d(C–N)/Å	1.386	—	1.391	1.377	1.378
μ/D	3.981	4.503	3.473	—	3.846[d]
Rotational constants/MHz	A = 11 132.9	A = 11 549.1	A = 11 053.8	—	A = 11 233.3
	B = 10 282.7	B = 10 530.8	B = 10 373.8		B = 10 369.4
	C = 5385.6	C = 5550.8	C = 5408.2		C = 5416.6

[a] The present results. [b] Ref. 9. [c] Ref. 10. [d] Ref. 11.

Table 3 B3LYP/6-31+G** calculated values for the hydrogen-bonded R_2CO·urea dimer

System[a]	E_{HB}[b]/kcal mol^{-1}	R(O···H)/Å	q(O, H, C)
F–0	−2.134	2.394	(−0.52, +0.32, +0.97)
F–90	−1.568	2.432	(−0.50, +0.31, +0.95)
Cl–0	−2.266	2.360	(−0.25, +0.32, −0.0067)
Cl–90	−1.772	2.411	(−0.21, +0.32, −0.034)
H–0	−4.654	2.259	(−0.36, +0.32, +0.10)
H–90	−3.546	2.359	(−0.36, +0.32, +0.11)
NH$_2$–0	−7.476	2.160	(−0.65, +0.34, +0.59)
NH$_2$–90	−7.055	2.150	(−0.67, +0.34, +0.64)
NHCH–0	−7.913	2.144	(−0.73, +0.34, +0.67)
NHCH–90	−8.212	2.139	(−0.71, +0.33, +0.65)

[a] See Scheme 1 for system definitions. [b] $E_{HB} = E[(\text{urea})_2] - 2 \times E(\text{urea})$.

This correlation between the average values of $\bar{E}_{HB}$, $\bar{R}$(O···H) and $\bar{q}$(O) points out the importance of Coulombic interactions in the hydrogen bond.

Our next concern is the basis for the difference in hydrogen bonding between $\theta = 0°$ and $\theta = 90°$ rotamers. From Fig. 2, it appears that there is some correlation between the hydrogen bond strengths of $\theta = 0°$ and $\theta = 90°$ rotamers and the energy levels of the accepting orbital,[13] n or π. For example, the R = **NHCH** compound has a very high π orbital relative to the n orbital and a stronger hydrogen bond for the 90° rotamer than for the 0° rotamer. This reminds us to consider these hydrogen bond acceptor orbitals as the electronic donor (ϕ_d) in the HOMO–LUMO type interaction involving the neighboring urea molecule. The perturbation contribution to the donor–acceptor pair is given as $\langle\phi_d|\hat{\mathbf{H}}|\phi_a\rangle^2/(\varepsilon_d - \varepsilon_a)$. The electronically donating orbital is the nonbonding orbital on oxygen for the $\theta = 0°$ rotamer and the carbonyl π orbital for the $\theta = 90°$ rotamer. We find the population of the oxygen nonbonding $2p_x$ orbital is 1.8, close to the value of the oxygen $2p_y$ π orbital, 1.6, in the urea molecule (the molecular axis is the z axis).

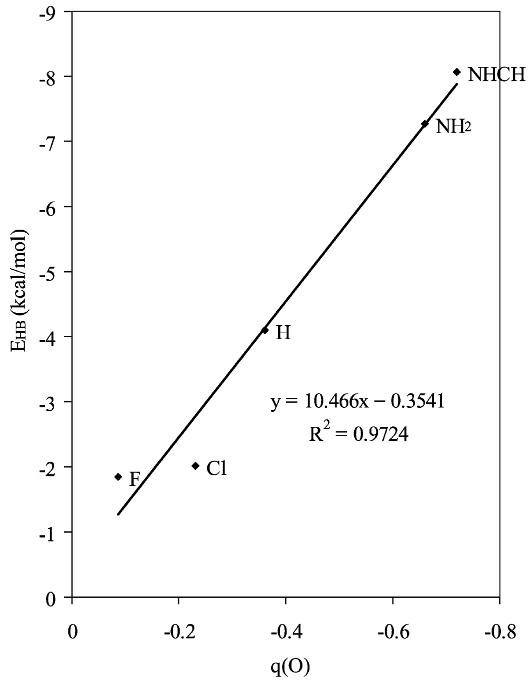


Fig. 1 The correlation between the averaged hydrogen bond energies, $\bar{E}_{HB}$ (E_+), and the averaged Mulliken charge on the carbonyl oxygen atoms, $\bar{q}$(O), for the R_2CO urea series shown in Scheme 1.

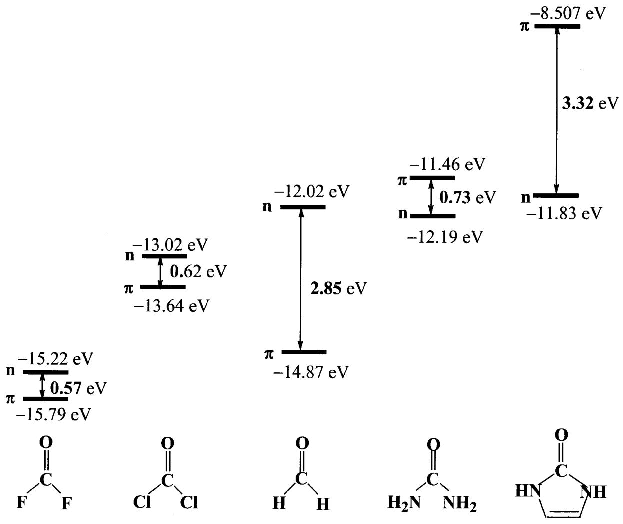


Fig. 2 RHF/6-31+G** calculated energies of the n and π orbitals of substituted carbonyl compounds.

Therefore, we made an approximation that the matrix element $\langle\phi_d|\hat{\mathbf{H}}|\phi_a\rangle$ is the same for the two different type of donor orbitals. The ϕ_a is the same σ^*_{NH} orbital for the two rotamers. Therefore, we expect that the ΔE_{HB} between the two rotamers is proportional to $[1/\varepsilon_n - \varepsilon_a] - 1/(\varepsilon_\pi - \varepsilon_a)]$. The expression within the bracket is termed $\Delta(1/\varepsilon) = [1/\varepsilon(0°) - 1/\varepsilon(90°)]$ in Table 4. The correlation coefficient square between ΔE_{HB} and $\Delta(1/\varepsilon)$ is found to be 0.9628, as shown in Fig. 3. Another donor-acceptor interaction between the symmetric nonbonding a_1 orbital on oxygen and σ^*_{NH} is independent of the dihedral angle, which is ignored here. The orbital energies are taken from the Hartree–Fock calculation instead of the B3LYP results. It is known that the former can be compared qualitatively with the ionization potential results and the latter should be adjusted through some unknown scaling scheme.[13]

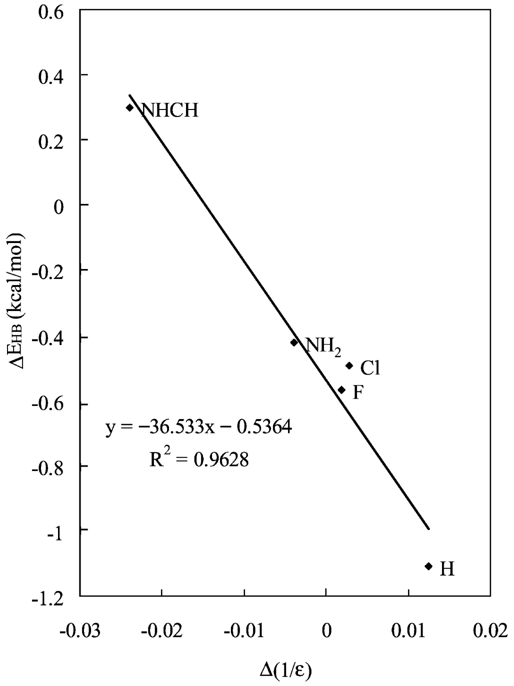


Fig. 3 The correlation between the hydrogen bonding energy and the inverse of the HOMO–LUMO gap. Both quantities are the difference between the $\theta = 0°$ and $\theta = 90°$ rotamers for the R_2CO·urea series shown in Scheme 1. See text and Table 4 for details of the HOMO–LUMO gap ε. ΔE_{HB} is also defined as $2E_-$ in the conclusion section.

Table 4 R_2CO·urea hydrogen bonding energies, E_{HB}, ΔE_{HB}, and the HOMO–LUMO gap, ε, calculated at the B3LYP/6-31 + G** level

System[a]	E_{HB}/kcal mol^{-1}	ΔE_{HB}[b] /kcal mol^{-1}	$1/\varepsilon$[c] /eV^{-1}	$\Delta 1/\varepsilon$[d] /eV^{-1}
F–0/F–90	−2.134/−1.568	−0.566	0.0589/0.0570	0.00192
Cl–0/Cl–90	−2.266/−1.772	−0.494	0.0677/0.0650	0.00273
H–0/H–90	−4.654/−3.546	−1.108	0.0726/0.0602	0.0125
NH_2–0/NH_2–90	−7.476/−7.055	−0.421	0.0717/0.0757	−0.00397
NHCH–0/NHCH–90	−7.913/−8.212	+0.299	0.0736/0.0975	−0.0239

[a] See Scheme 1 for system definitions. [b] $\Delta E_{HB} = E_{HB}(0°) - E_{HB}(90°)$. [c] $1/\varepsilon = 1/[\varepsilon(\text{LUMO}) - \varepsilon(\text{HOMO})] = 1/[1.748 - \varepsilon(\text{HOMO})]$, $\varepsilon(\text{LUMO}) = \varepsilon(\sigma^*_{NH}) = 1.748$ eV. For $\theta = 0°$, the HOMO is a nonbonding orbital n. For $\theta = 90°$, the HOMO is a carbonyl π orbital. [d] $\Delta 1/\varepsilon = 1/\varepsilon(0°) - 1/\varepsilon(90°) = [1/(\varepsilon_n - \varepsilon_a) - 1/(\varepsilon_\pi - \varepsilon_a)]$, $\varepsilon_a = 1.748$ eV.

Conclusion

For the R_2CO·urea (R = H, F, Cl, NH_2, NHCH) series, the hydrogen bonding interaction between the R_2CO acceptor and the urea donor results from the superposition of two components. One component is the Coulombic interaction between the carbonyl oxygen and amine hydrogen as manifested by the average energy (E_+) of the $\theta = 0°$ and $\theta = 90°$ rotamers for each compound. The second component is an electronic donor-acceptor interaction like the familiar HOMO–LUMO type interaction. The electronically donating orbital (ϕ_d) is the oxygen nonbonding orbital for $\theta = 0°$ and the carbonyl π orbital for $\theta = 90°$. In other words, if the $E_\pm$ values are defined as $[E_{HB}(0°) \pm E_{HB}(90°)]/2$, then $E_{HB}(0°)$ and $E_{HB}(90°)$ can be re-expressed as $E_+ \pm E_-$. The first component, E_+, can be interpreted as due to a Coulombic-type interaction as shown in Fig. 1, and E_- can be accounted for by a electronic donor-acceptor type interaction, as shown in Fig. 3. Interestingly, the concept of both Coulombic and non-Coulombic types of interactions in hydrogen bonding situations have been well received among the chemistry community.[14–17] Recently, there appeared to be some overemphasis by physicists[18] on the second component, termed "covalent" interaction.

In terms of orbital energies, the substituents have two effects on the hydrogen bond strength. Firstly, there is the shift of the n and π orbital energies as a whole, accompanied by a change in the charge on oxygen, $\bar{q}(\text{O})$ (Table 3 and Fig. 2). In fact, the correlation coefficient square between $\bar{q}(\text{O})$ and the quantity $(\varepsilon_n + \varepsilon_\pi)$ is found to be 0.899 for the series. Secondly, there is the splitting of the energy gap between n and π. The π-donor substituent can enhance effectively the hydrogen bonding capability of the carbonyl π-electrons. The urea molecule is quite unique in that $\theta = 0°$ and $\theta = 90°$ have nearly equal hydrogen bond strengths, resulting in small values of E_- relative to E_+, such that the structures **2** and **3** are both observable.

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