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Push−Pull Electronic Effects in Charge-Transfer Complexes: The Case of N−H and N−Me Lactams

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Equilibrium constants and Gibbs free energy changes for 1:1 charge-transfer complexes between I_2 or ICl and lactams (N−H and N−Me) were determined in solution by means of UV−visible spectroscopic techniques. As expected, ICl-complexes are stronger than I_2 complexes. Ab initio calculations at HF/LANL2DZ* and MP2(full)/LANL2DZ* levels of theory were carried out to investigate the structures of the complexes and the nature of the intermolecular interaction between the lactams and I_2 or ICl. Systematically, the global minimum of the potential energy surface corresponds to the approach of I_2 or ICl to the carbonyl oxygen, with the exception of aziridinone. Two stable conformers, syn and anti with respect to the ring nitrogen atom, were found to be local minima of the potential energy surface. For unsubstituted lactams, the syn conformer is stabilized through a hydrogen-bond interaction with the NH group of the base and becomes the most stable structure. Conversely, for the *N*-methyl substituted derivatives, where this interaction is not possible, the anti conformer is the most stable one. Experimental free energies in solution and gas-phase theoretical values follow a good linear relationship. Both the experimental results and the ab initio calculations showed that, contrary to lactones, lactams are more basic than cyclic ketones with respect to ICl and I_2.

Introduction

Lactams have attracted a great deal of attention because of the presence of the β-lactam moiety in a number of major antibiotics.[1−5] A number of studies on their bactericidal properties,[4,5] molecular structures, and reactivity, including gas-phase basicities toward proton,[6−10] have been carried out. The latter allow us to assess, inter alia, the effects of cyclization on basicity, relative to aliphatic and cyclic amides.[11] Despite this wealth of information, there seems to be no systematic studies about their charge-transfer complexation with molecular iodine and iodine monochloride. It is well-known that this kind of study was triggered by the experimental determination of the existence, stoichiometry, and thermodynamic stability of 1:1 complexes between hydrocarbons and iodine in solution.[12] Soon after, substantial experimental studies of charge-transfer (CT) complexes, involving both π and n donor bases, were reported. Interestingly, the basic concept of Mulliken's theory[13,14] has played a key role in the development of Drago's classical quantitative empirical model of reactivity.[15]

Our longstanding interest in CT complexes involving n-donor bases and molecular iodine and iodine monochloride[16−21] has prompted us to extend this study to lactams. More precisely, we have studied the following compounds: aziridinone (**1**), azetidin-2-one (**2**), pyrrolidin-2-one (**3**), δ-valerolactam (**4**), and ϵ-caprolactam (**5**); see Figure 2.

For this purpose, we have measured the equilibrium constants K_c (I_2 or ICl) for reaction 1 in solution, from which we derived their solution basicity toward I_2 and ICl, as measured respectively by $\Delta_r G°$ (I_2 or ICl), the standard Gibbs energy change for reaction 1

$$B + I_2(ICl) \rightleftharpoons B\cdots I_2(ICl) \quad K_c\ (I_2 \text{ or ICl}) \quad \Delta_r G°\ (I_2 \text{ or ICl}) \qquad (1)$$

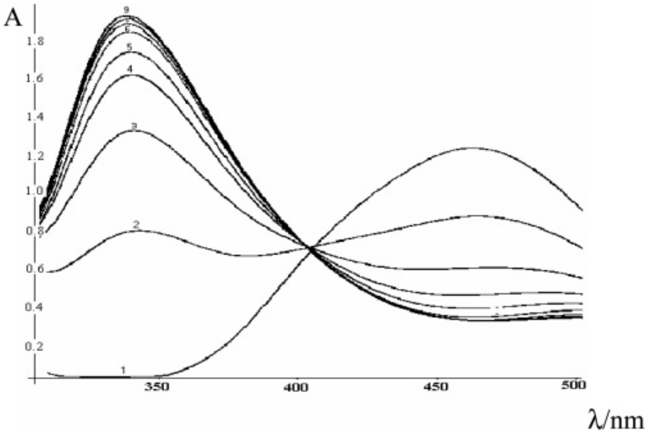

Figure 1. UV/vis spectra of pyrrolidin-2-one + ICl, measured in CCl_4 at 25 °C, with constant ICl concentration ($c = 7.8 \times 10^{-3}$ mol/L) and increasing pyrrolidin-2-one concentration: **1** ($c = 0$), **2** ($c = 1.07 \times 10^{-3}$), **3** ($c = 2.04 \times 10^{-3}$), **4** ($c = 2.96 \times 10^{-3}$), **5** ($c = 4.09 \times 10^{-3}$), **6** ($c = 6.08 \times 10^{-3}$), **7** ($c = 8.20 \times 10^{-3}$), **8** ($c = 10.27 \times 10^{-3}$), and **9** ($c = 12.3 \times 10^{-3}$ mol/L).

The combination of experimental thermodynamic data with ab initio calculations[19−21] is proving to be a powerful tool for the understanding of the structure and reactivity of molecules and ions. Thus, parallel to the experimental work, free and I_2- and ICl-complexed lactams were studied at the HF/LANL2DZ* and MP2/LANL2DZ* levels of theory.

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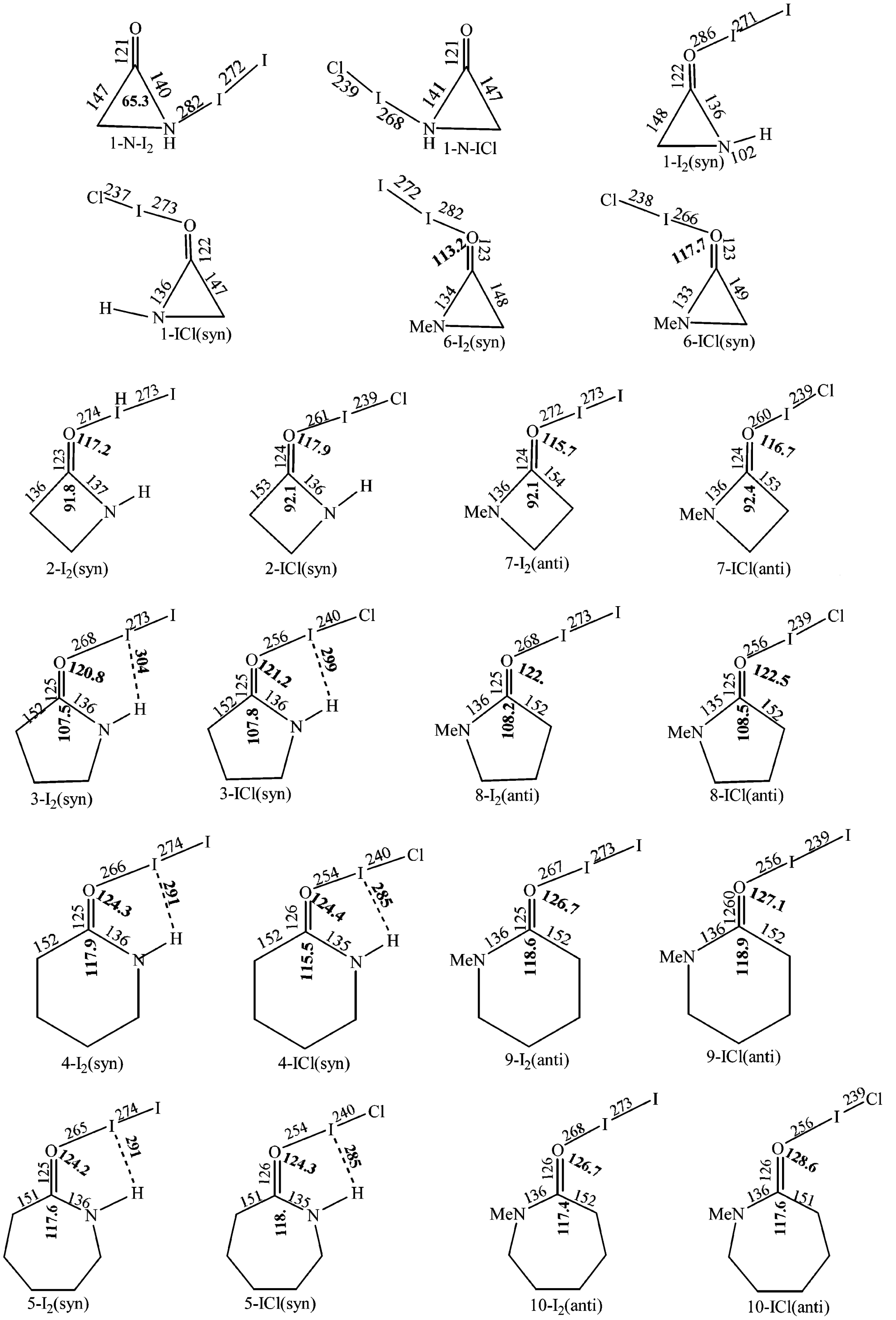

Figure 2. MP2(full)/LANL2DZ* optimized geometries of the most stable complexes lactams-I_2 and lactams-ICl. Distances in pm and angles in degree.

To investigate the possible effect of *N*-methyl substitution on the characteristics of these CT complexes, the following derivatives were added to the previous set in our theoretical analysis: 1-methylaziridinone (**6**), 1-methylazetidin-2-one (**7**), 1-methylpyrrolidin-2-one (**8**), 1-methyl-δ-valerolactam (**9**), 1-methyl-ϵ-caprolactam (**10**). From these, only compounds **8** and **10**, which are the only ones commercially available, have been also included in our experimental scrutiny.

Experimental Section

Compounds studied in this work were of commercial origin. Solvent (CCl_4) of spectrograde quality was purified according to the literature.[22] A Cary 219 spectrophotometer was used to determine the equilibrium constants for the 1:1 association between the lactam compounds and I_2 and ICl in solution by means of UV−visible spectroscopy. The spectrophotometric measurements were carried out using 1 cm matched silica cells kept at 25 °C. The formation of the complex was detected easily by the displacement of the acceptor band. The single isobestic point located between the bands of free and associated I_2 or ICl in the UV−visible spectrum (see Figure 1) indicates that the CT complexes formed have a 1:1 stoichiometry as shown in the literature.[23] The procedure used for calculating the equilibrium constants, K_c, for reaction 1, is Drago's method,

TABLE 1: Experimental Values of the Equilibrium Constants K_c (L mol^{-1}) and the Gibbs Energies $\Delta_r G°$(kcal·mol^{-1}) of Lactams-I_2(ICl), Cyclic Ketones-I_2(ICl), and Lactones-I_2(ICl) 1:1 Complexes, Determined in CCl_4 at 25 °C

no.	compound	K_c (I_2)	$\Delta_r G°$ (I_2)	K_c (ICl)	$\Delta_r G°$ (ICl)
2	azetidin-2-one	2.50 ± 0.3	−0.55 ± 0.07	50 ± 10	−2.30 ± 0.12
3	pyrrolidin-2-one	7.5 ± 0.9	−1.20 ± 0.07	300 ± 34	−3.40 ± 0.07
4	δ-valerolactam	10.40 ± 0.34	−1.40 ± 0.02	520 ± 20	−3.73 ± 0.03
5	ϵ-caprolactam	10.77 ± 0.1	−1.42 ± 0.06	558 ± 10	−3.77 ± 0.01
8	1-methylpyrrolidin-2-one	10.24 ± 1.5	−1.39 ± 0.09	604 ± 24	−3.82 ± 0.03
10	1-methyl-ϵ-caprolactam	12.81 ± 1.8	−1.52 ± 0.08	822 ± 45	−4.00 ± 0.04
12	cyclobutanone[a]	0.44 ± 0.11	0.49 ± 0.15	13.04 ± 0.86	−1.53 ± 0.04
13	cyclopentanone[a]	0.94 ± 0.18	0.04 ± 0.12	35.00 ± 1.00	−2.12 ± 0.02
14	cyclohexanone[a]	1.47 ± 0.13	−0.23 ± 0.05	44.00 ± 2.00	−2.26 ± 0.03
15	cycloheptanone[a]	1.43 ± 0.16	−0.21 ± 0.07	55.17 ± 0.01	−2.39 ± 0.01
17	β-propiolactone[a]	0.08 ± 0.03	1.50 ± 0.23	0.90 ± 0.10	0.06 ± 0.07
18	γ-butyrolactone[a]	0.52 ± 0.10	0.39 ± 0.12	17.33 ± 0.84	−1.70 ± 0.03
19	δ-valerolactone[a]	0.95 ± 0.02	−0.03 ± 0.02	44.34 ± 0.61	−2.26 ± 0.01
20	ϵ-caprolactone[a]	1.01 ± 0.12	−0.06 ± 0.07	41.46 ± 0.20	−2.22 ± 0.01

[a] From reference 21.

and his simplifications in the case of the small values of K_c have already been described in detail.[24,25] The equilibrium constants reported in Table 1 are the average of at least three independent measurements.

Computational Details

Quantum chemical calculations were performed with the aid of the Gaussian 98 set of programs.[26] All structures were fully optimized at SCF and MP2 (full) levels. All these calculations have been carried out using a LANL2DZ* basis set[27,28] which includes an effective core potential (ECP) for all atoms except those of the first row. The ECP used is that proposed by Hay and Wadt,[27] which for iodine incorporates the mass velocity and Darwin relativistic effects. The LANL2DZ basis set corresponds to a Dunning/Huzinaga full double ξ-basic (D95)[29] for first row elements and to an effective core potential plus double-ξ basis for iodine atoms.[30] This basis set was augmented by one set of six d polarization function (LANL2DZ*) with the following exponents[31−33] $\alpha_C = \alpha_O = \alpha_N = 0.8$, $\alpha_{Cl} = 0.75$, and $\alpha_I = 0.29$. This basis set has been shown to yield reasonably good results for I_2 complexes with pyridine,[34] thiocarbonyl derivatives,[19] and carbonyl−ICl complexes.[20] Even though the effective core potential used takes into account the two most important relativistic effects, spin−orbit interactions are not included in the model. Nevertheless, since the ground states of the systems under consideration are always singlets, these spin−orbit corrections will be zero at the first order. The basis set superposition error (BSSE) which affects these binding energies was estimated by means of counterpoise method.[35] The harmonic vibrational frequencies were determined at the HF/LANL2DZ* level to confirm that the optimized structures found correspond to real minima on the potential energy surface and to evaluate the zero-point energy (ZPE) corrections, thermal corrections, and the entropy. The ZPE and the stretching frequencies were scaled by the empirical factor 0.893.[36]

The natural bond orbital (NBO)[30] analysis and the atoms in molecules (AIM) theory,[37−40] which is based on a topological analysis of electron charge density and its Laplacian, have been used to get more information on the electronic-charge density and to characterize some intramolecular hydrogen bonds.

Experimental Results. Table 1 lists the thermodynamic data obtained in this work for the formation of complexes of some lactams with molecular iodine and iodine monochloride in tetrachloromethane as solvent. To extend the range of correlation and for a possible comparisons, these results are presented together with data of cyclic ketones and lactones from previous studies.[21] Equilibrium constants were determined as indicated above.

We recall that the uncertainties on $\Delta_r G°$ values were estimated using the differentiation of eq 2 under constant T (experimental conditions) leading to eq 3

$$\Delta_r G°[(I_2,\mathrm{ICl});\ \mathrm{soln}] = -RT \ln K_c \tag{2}$$

$$\delta\Delta_r G°[(I_2,\mathrm{ICl});\ \mathrm{soln}] = -RT\, \delta K_c/K_c \tag{3}$$

Computational Results

The total energies evaluated at the MP2(full)/LANL2DZ* level, as well as the scaled ZPE and entropy values, evaluated at the HF/LANL2DZ* level, of different bases and complexes considered in this work, are reported in Tables 1s and 2s of the Supporting Information.

Discussion

A. Structures. To gain further insight on the characteristics of these CT complexes, it is necessary to investigate their equilibrium structures, as well as the relative stabilities of the possible conformations. The MP2(full)/LANL2DZ* optimized molecular structures for the most stable I_2 and ICl CT complexes, showing the most significant structural parameters, are depicted in Figure 2. The complexation at the carbonyl oxygen is the most favorable process exception made in the case of aziridinone where the basic center is the nitrogen atom instead of the oxygen one. It is worth noting that, for the other unsubstituted lactams, the sy*n* conformer, defined by the relative position of I_2 or ICl with respect to the ring nitrogen atom, was found to be more stable than the anti conformer which is more stable in the case of *N*-methyl-lactams, where I_2 (ICl) approaches the opposite site of the NCH_3. The syn position is stabilized by the hydrogen-bond interaction between the incoming iodine and the hydrogen atom of the NH group (see later). More importantly, in the case of *N*-methyl-lactams and for the larger cycles, the syn structure collapses to the anti ones because the steric hindrance in the syn position with the substituent is larger for the larger cycles.

For the free compounds, the C═O bond length is on average about 123 pm and it lengthens upon complexation being about 124.8 pm. It is well-known that the interaction takes place between the lone pair of the donor and the antibonding orbital of iodine.[13,14,19−21,41,42] This is necessarily reflected in the

TABLE 2: Calculated Wavenumbers, $\tilde{\nu}$ (in cm^{-1}), of the C═O Stretching Vibration of Compounds 1−10 Evaluated at HF/LANL2DZ* and Its Displacements[a]

	B	$B-I_2$		B−ICl	
	$\tilde{\nu}_{C=O}$	$\tilde{\nu}_{C=O}$	$\Delta\tilde{\nu}_{C=O}$	$\tilde{\nu}_{C=O}$	$\Delta\tilde{\nu}_{C=O}$
1	1932.5	1916.4	16	1905.7	27
2	1824.4	1794.0	30	1777.1	47
3	1771.7	1738.7	33	1719.9	52
4	1738.7	1702.1	36	1684.2	55
5	1733.3	1701.2	32	1681.5	52
6	1916.4	1896.7	20	1886.0	30
7	1801.2	1768.1	33	1751.2	50
8	1746.7	1711.0	36	1691.3	55
9	1711.9	1670.8	41	1647.6	64
10	1704.7	1670.8	34	1651.2	54

[a] The calculated values were scaled by the empirical factor 0.893.[36]

lengthening of the I−I and I−Cl bond distances. Thus, the I−I and I−Cl bond lengths within the corresponding complexes (272.7 and 238.9 pm, respectively, on average) are longer than the distances in the free molecules (269.7 pm and 235.1 pm respectively) calculated at the same level of theory. The O···I−I and O···I−Cl arrangement is essentially linear with an angle close to 180°. The structural changes are also mirrored in the corresponding force field of the system and the C═O stretching frequencies appear systematically red shifted upon complexation (see Table 2). This variation is more pronounced in the case of ICl complexes and increases with the ring size. For five-, six-, and seven-membered rings the effect is nearly constant. The same behavior was found for the analogous lactones[21] and in the case of protonation.[43,44]

B. Energetics. We summarize in Table 3 the thermodynamic state functions pertaining to the formation of the various isomeric structures of the adducts in the gas phase as computed in this work. The calculations include BSSE, ZPE, and thermal corrections for the translational, rotational, and vibrational degrees of freedom. The Δ(PV) term is also taken into account. Because the experimental measurements are carried out in solution with 1 mol/L as standard state and the calculated values are referred to the gas phase at 298.15 K and 1 atm, which corresponds to a concentration of $c = 4.09\ 10^{-2}$ mol/L, we have corrected the calculated values by the 1.89 kcal/mol term.[21]

Some relevant features of these results are as follows:

(1) For all lactams, the complexation at the carbonyl oxygen is the most favorable process, as was found in other similar cases,[20,21,43−46] with the exception of aziridinone, where the basicity of the nitrogen atom is slightly higher than that of the carbonyl oxygen. The same results were found in the case of protonation.[12] However, it is to note that protonation leads to ring opening contrary to the I_2 or ICl complexation which occurs without affecting significantly the structure of the cycle. This is due to the fact that, in the case of I_2 or ICl complexation, the interaction is essentially electrostatic whereas in the case of protonation, the interaction leads to a covalent bond.

(2) For the majority of the adducts, the stabilities of the syn and anti (relative to the nitrogen atom) forms as measured by the corresponding $\Delta_r G°(I_2,ICl)$ values are quite comparable. Physically, this tends to imply that the CT complexes observed in solution frequently are equilibrium mixtures of both isomeric structures. In principle, one should experimentally observe reactions 4 and 5, characterized by the constants $K_c(I_2,ICl)_{syn}$ and $K_c(I_2,ICl)_{anti}$

$$B + I_2(ICl) \rightleftharpoons (B\cdots I_2(ICl))_{syn}\ K_c(I_2,ICl)_{syn}\ \Delta_r G°(I_2,ICl)_{syn} \tag{4}$$

$$B + I_2(ICl) \rightleftharpoons (B\cdots I_2(ICl))_{anti}\ K_c(I_2,ICl)_{anti}\ \Delta_r G°(I_2,ICl)_{anti} \tag{5}$$

Unfortunately, the experimental technique does not allow the equilibrium constants $K_c(I_2,ICl)_{syn}$ and $K_c(I_2,ICl)_{anti}$ to be determined separately. The actual experimental datum is the "experimental equilibrium constant" $K_c(I_2,ICl)$ (given in Table 1), equal to $K_c(I_2,ICl)_{syn} + K_c(I_2,ICl)_{anti}$, that is

$$K_c(I_2,ICl) = \{[(B\cdots I_2(ICl))_{syn}] + [(B\cdots I_2(ICl))_{anti}]\}/[B][I_2(ICl)] \tag{6}$$

For the purposes of comparing the experimental and computational results and getting further mechanistic insights, it is useful to define the dimensionless equilibrium constant K_{isom} pertaining to the isomerization reaction 7

$$[B\cdots I_2(ICl)]_{anti} \rightleftharpoons [B\cdots I_2(ICl)]_{syn}\ K_{isom}(I_2,ICl)\ \Delta_r G°_{isom}(I_2,ICl) \tag{7}$$

K_{isom} determines the position of equilibrium (7) and $\Delta_r G°_{isom}(I_2,ICl)$ is given by $-RT \ln K_{isom}$. Also, $\Delta_r G°_{isom}(I_2,ICl) = [\Delta_r G°(I_2,ICl)]_{syn} - [\Delta_r G°(I_2,ICl)]_{anti}$. It is of importance that, although K_{isom} and $\Delta_r G°_{isom}$ (I_2,ICl) cannot be experimentally measured, the computational results presented in Table 3 do allow a quantitative estimate of these properties. It can be easily shown that

$$K_c(I_2,ICl) = K_c(I_2,ICl)_{anti}[1 + K_{isom}(I_2,ICl)] \tag{8}$$

and

$$\Delta_r G°(I_2,ICl) = [\Delta_r G°(I_2,ICl)]_{anti} - RT \ln [1 + K_{isom}(I_2,ICl)] \tag{9}$$

Notice that, in cases such as in the complexes of (**1**) and on account of the very small differences in stability of the oxygen (syn or anti), or nitrogen-bound complexes, we must consider two isomerization processes. Then, two equilibrium constants are relevant, $K_{isom}(I_2,ICl)$, defined above, and $K'_{isom}(I_2,ICl)$, pertaining to reaction 10

$$[B\cdots I_2(ICl)]_{anti} \rightleftharpoons [B\cdots I_2(ICl)]_N\ K'_{isom}(I_2,ICl)\ \Delta_r G'°_{isom}(I_2,ICl) \tag{10}$$

In this case, $\Delta_r G'°_{isom}$ (I_2,ICl) is given by $-RT \ln K'_{isom}(I_2,ICl)$ and, $\Delta_r G'°_{isom}$ $(I_2,ICl) = [\Delta_r G°(I_2,ICl)]_N - [\Delta_r G°(I_2,ICl)]_{anti}$. From this, eq 11 readily follows

$$\Delta_r G°(I_2,ICl) = [\Delta_r G°(I_2,ICl)]_{anti} - RT \ln [1 + K_{isom}(I_2,ICl) + K'_{isom}(I_2,ICl)] \tag{11}$$

$\Delta_r G°(I_2,ICl)$ can be obtained experimentally (at least in solution, as it is the case here), whereas $[\Delta_r G°(I_2,ICl)]_{anti} - RT \ln [1 + K_{isom}(I_2,ICl)]$ and $[\Delta_r G°(I_2,ICl)]_{anti} - RT \ln [1 + K_{isom}(I_2,ICl) + K'_{isom}(I_2,ICl)]$ can be fully estimated by computational means. We present in Table 3 the computed values of K_{isom}, $\Delta_r G°_{isom}(I_2,ICl)$, and $\Delta_r G°(I_2,ICl$, computed). The experimental $\Delta_r G°(I_2,ICl)$ values are given in Table 1.

Figures 3 and 4 portray comparisons of the experimental values of $\Delta_r G°(I_2,ICl)$, from Table 1 against the corresponding computed values taken from Table 3.

TABLE 3: Calculated Values (in kcal mol^{-1}) of Selected Thermodynamic State Functions at the MP2 (full)/LANL2DZ* Level for Lactams, Cyclic Ketones, and Lactones-I_2/ICl Complexes

species	$\Delta_r H°$(I2,ICl)	$T\Delta_r S°$ (I2,ICl)	$\Delta_r G°$ (I2,ICl)	K_{isom}(I_2,ICl)	$\Delta_r G°_{isom}$(I_2,ICl)	$\Delta_r G°$(I2,ICl)$_{(computed)}$[a]
1−I_2(syn)	−5.70(−5.49)	−6.15	−1.44	1.11	−0.060	
1−I_2(anti)	−5.67(−5.37)	−6.18	−1.38			−2.08
1N−I_2	−6.20(−6.42)	−6.74	−1.35	1.16	−0.090	
1−ICl(syn)	−7.23(−7.65)	−7.67	−1.45	1.16	−0.090	
1−ICl(anti)	−7.20(−7.51)	−7.73	−1.36			−1.99
1N−ICl	−8.10(−8.97)	−8.36	−1.63	0.74	0.180	
2−I_2(syn)	−7.27(−7.58)	−6.95	−2.21	1.25	−0.130	
2−I_2(anti)	−7.02(−7.46)	−6.83	−2.08			−2.56
2−ICl(syn)	−9.59(−10.46)	−8.63	−2.85	1.07	−0.04	
2−ICl(anti)	−9.46(−10.28)	−8.54	−2.81			−3.24
2N−ICl	−8.00(−9.28)	−9.13	−0.75	11.8	−1.46	
3−I_2(syn)	−8.01(−8.69)	−6.98	−2.92	1.204	−0.110	
3−I_2(anti)	−7.81(−8,34)	−6.89	−2.81			−3.28
3−ICl(syn)	−10.72(−12.00)	−8.51	−4.10	1.526	−0.250	
3−ICl(anti)	−10.38(−11.48)	−8.42	−3.85			−4.40
4−I_2(syn)	−8.21(−9.11)	−6.98	−3.12	1.11	−0.060	
4−I_2(anti)	−8.06(−8.69)	−6.89	−3.06			−3.50
4−ICl(syn)	−11.03(−12.57)	−8.57	−4.35	1.63	−0.290	
4−ICl(anti)	−10.62(−11.94)	−8.45	−4.06			−4.63
5−I_2(syn)	−8.26(−9.20)	−7.10	−3.35	1.87	−0.370	
5−I_2(anti)	−7.89(−8.51)	−6.80	−2.98			−3.60
5−ICl (syn)	−11.08(−12.65)	−8.75	−4.22	1.63	−0.290	
5−ICl (anti)	−10.52(−11.64)	−8.48	−3.93			−4.50
6-I_2(syn)	−6.33(−6.41)	−6.41	−1.81	0.93	0.04	
6-I_2(anti)	−6.22(−6.09)	−6.29	−1.85			−2.25
6-ICl(syn)	−8.55(−8.89)	−7.88	−2.56	1.69	−0.31	
6-ICl(anti)	−8.15(−8.52)	−7.79	−2.25			−2.84
7−I_2(syn)	−6.76(−7.35)	−6.47	−2.18	0.45	0.47	−2.87
7−I_2(anti)	−7.65(−8.09)	−6.89	−2.65			
7−ICl(syn)	−9.26(−10.13)	−8.00	−3.15	0.46	0.46	−3.84
7−ICl(anti)	−10.17(−11.15)	−8.45	−3.61			
8−I_2(syn)	−7.19(−7.54)	−6.44	−2.64	0.44	0.48	
8−I_2(anti)	−8.21(−8.88)	−6.98	−3.12			−3.34
8−ICl(syn)	−9.56(−10.21)	−7.35	−4.11	0.53	0.38	
8−ICl(anti)	−11.05(−12.16)	−8.45	−4.49			−4.74
9−I_2(syn)	−	−	−		>3.52	−3.52
9−I_2(anti)	−8.40(−9.17)	−6.77	−3.52			
9−ICl(syn)	−	−	−		>4.69	
9−ICl(anti)	−11.21(−12.58)	−8.36	−4.69			−4.69
10−I_2(syn)	−	−	−			
10−I_2(anti)	−8.38(−9.06)	−6.98	−3.29		>3.29	−3.29
10−ICl(syn)	−	−	−			
10−ICl(anti)	−11.17(−12.30)	−8.54	−4.52		>4.52	−4.52
11−I_2	−5.73(−7.57)	−6.33	−1.29			−1.70
11−ICl	−7.37(−9.67)	−7.88	−1.38			−1.79
12−I_2	−6.31(−8.21)	−6.35	−1.85			−2.26
12−ICl	−8.15(−10.65)	−8.11	−1.73			−2.14
13−I_2	−6.55(−8.80)	−6.67	−2.01			−2.42
13−ICl	−8.57(−11.37)	−8.22	−2.24			−2.65
14−I_2	−6.64(−8.83)	−6.59	−1.94			−2.35
14−ICl	−8.70(−11.33)	−8.09	−2.50			−2.91
15−I_2	−6.66(−8.89)	−6.57	−2.06			−2.47
15−ICl	−8.75(−11.46)	−8.10	−2.69			−3.10
16−I_2 (anti)	−4.85(−6.35)	−5.82	−0.92			
16−I_2 (syn)	−4.74 (−6.09)	−5.63	−1.00	1.14	−0.08	−1.37
16−ICl (anti)	−6.10(−8.00)	−7.35	−0.64			
16−IC l (syn)	−6.00(−7.76)	−7.23	−0.66	1.03	−0.02	−1.06
17−I_2 (anti)	−5.58(−7.41)	−6.22	−1.25			
17−I_2 (syn)	−5.46(−6.99)	−6.11	−1.24	0.98	0.01	−1.66
17−ICl (anti)	−7.15(−9.47)	−7.75	−1.29			
17−ICl (syn)	−7.09(−9.06)	−7.43	−1.55	0.65	0.26	−1.85
18−I_2 (syn)	−6.37(−8.15)	−6.48	−1.79			
18−I_2 (anti)	−6.24(−8.26)	−6.43	−1.70	0.86	0.09	−2.16
18−ICl (syn)	−8.38(−10.64)	−7.95	−2.32			
18−ICl (anti)	−8.10(10.67)	−8.00	−1.99	0.57	0.33	−2.59
19−I_2 (syn)	−6.63(−8.93)	−6.69	−1.95			
19−I_2 (anti)	−6.50(−6.62)	−6.47	−1.83	0.82	0.12	−2.31
19−ICl (syn)	−8.80(−11.65)	−8.17	−2.52			
19−ICl (anti)	−8.50(−11.19)	−8.02	−2.37	0.78	0.15	−2.86
20−I_2 (syn)	−6.65(−9.08)	−6.73	−1.97			
20−I_2 (anti)	−6.56(−8.68)	−6.45	−1.72	0.66	0.25	−2.27
20−ICl (syn)	−8.81(−11.84)	−8.21	−2.49			
20−ICl (anti)	−8.52(−11.19)	−7.99	−2.42	0.89	0.07	−2.87

[a] Obtained from eqs 9 and 11 defined in the text. **11**: cyclopropanone. **16**: oxiran-2-one.

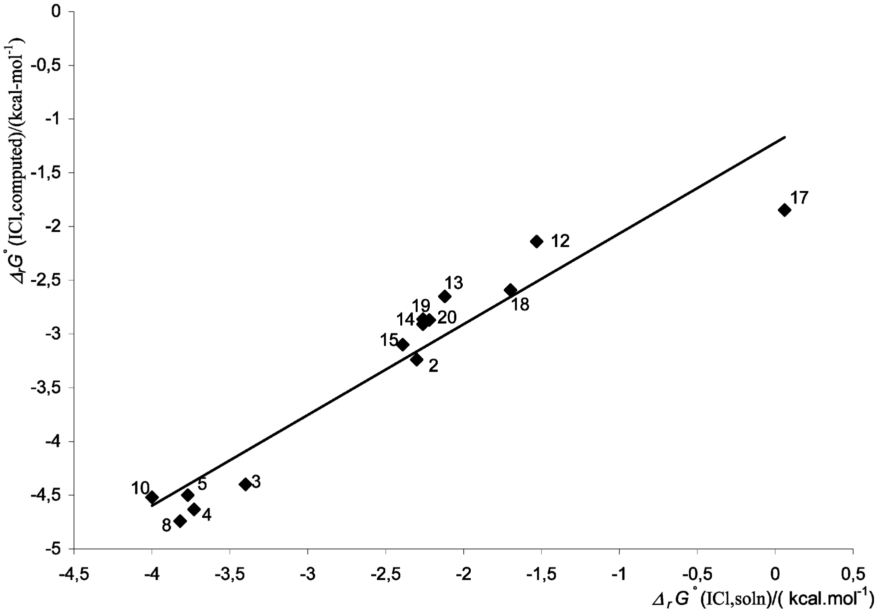

Figure 3. Correlation between $\Delta_r G°$(ICl,soln) and $\Delta_r G°$(ICl,computed).

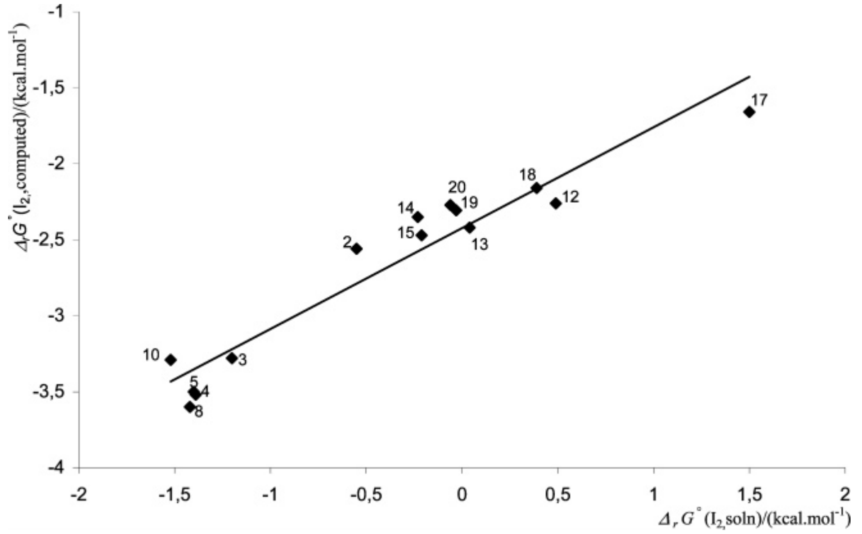

Figure 4. Correlation between $\Delta_r G°$(I_2,soln) and $\Delta_r G°$(I_2,computed).

The corresponding correlation equations are

$$\Delta_r G°(\text{ICl,computed}) = (0.84 \pm 0.08)\Delta_r G°(\text{ICl,soln}) - (1.22 \pm 0.21)$$

$$r = 0.953;\ \text{s.d.} = 0.31\ \text{kcal/mol.};\ n = 14 \qquad (12)$$

If we exclude the point corresponding to the β-propiolactone, **17**, which presents some experimental difficulties, the correlation equation becomes

$$\Delta_r G°(\text{ICl,computed}) = (1.05 \pm 0.06)\Delta_r G°(\text{ICl,soln}) - (0.61 \pm 0.16)$$

$$r = 0.985;\ \text{s.d.} = 0.16\ \text{kcal/mol.};\ n = 13 \qquad (13)$$

In the case of I_2, the correlation equation is

$$\Delta_r G°(\text{I}_2\text{,computed}) = (0.66 \pm 0.05)\Delta_r G°(\text{I}_2\text{,soln}) - (2.42 \pm 0.05)$$

$$r = 0.959;\ \text{s.d.} = 0.17\ \text{kcal/mol.};\ n = 14 \qquad (14)$$

The quality of the correlations is remarkably good, particularly considering the very small range of structural effects. This high degree of self-consistency strongly supports the experimental and computational techniques used in this work. Furthermore, the slopes of the correlation eqs 7−9 are very close to unity and this indicates a very good computational description of differential structural effects on reactivity. In some cases, the intercepts are small but not nil. We have attributed this fact, particularly in the case of I_2 complexation,[19,21] to the sizable error likely affecting the calculated entropies and consequently the calculated free energies. This is so because the complexes considered in this series of studies have six vibrational modes which are absent in the separate partners. Five of them have very low frequencies and are likely to be quite anharmonic and not adequately described by the harmonic approximation employed in ab initio calculations.[47] However, it is important that the possibility exists for the intercepts to be significant even in the absence of these errors. The origin of this situation is the influence of the solvent on the position of eq 1. Thus, even in the absence of significant "specific" solvent−solute interactions, van der Waals forces (largely dispersive) lead to quite sizable changes in $K_c(I_2,ICl)$[48,49] that can be quantitatively explained

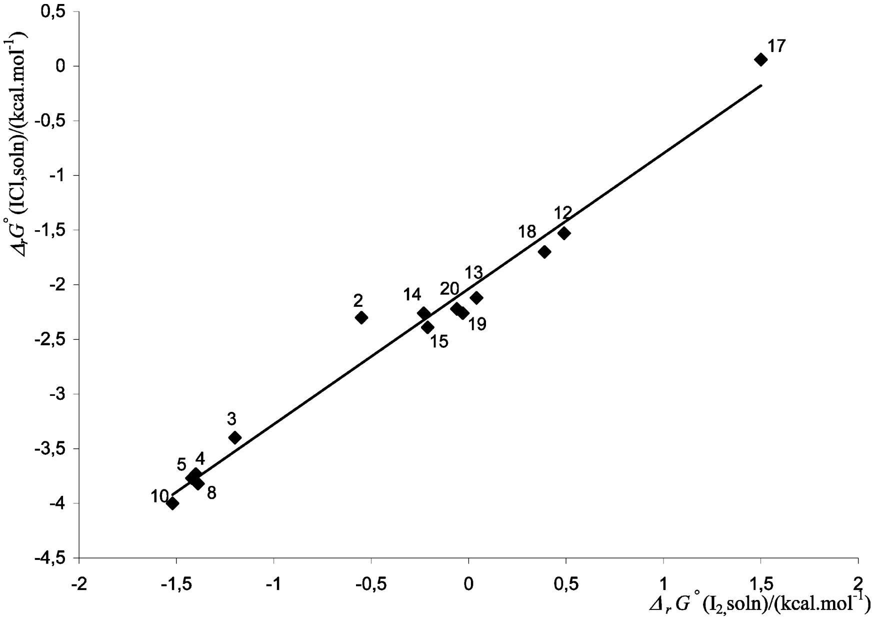

Figure 5. Correlation between the experimental values of $\Delta_rG°$(ICl,soln) and of $\Delta_rG°$(I_2,soln) for the complexed lactams, lactones, and cyclic ketones.

in a number of cases by means of Hildebrand's classical theory of regular solutions.[50−52]

(3) From the structural point of view, considering cyclic ketones, lactones, and lactams, the experimental $\Delta_rG°_m$ values reported in Table 1 show that these cyclic compounds exhibit a greater basicity toward ICl than toward I_2. Figure 5 portrays the correlation between their experimental Gibbs energy of association with I_2 and with ICl in solution

$$\Delta_rG°(\text{ICl,soln}) = (1.24 \pm 0.05)\Delta_rG°(\text{I}_2\text{,soln}) - (2.04 \pm 0.05)$$

$$r = 0.989;\ \text{s.d.} = 0.17\ \text{kcal/mol};\ n = 14 \qquad (15)$$

This result indicates the following: (i) Cyclic ketones, lactones, and lactams constitute a homogeneous family in terms of structural responses to interaction with I_2 and ICl; (ii) The slope of eq 15 shows that the complexation with ICl is more sensitive to structural effects than I_2 complexation. This behavior can be easily understood if we compare the dipole moments of I_2 (0.0)- and ICl (1.207 ± 0.003) in Debye units.[53,54] The latter polarizes strongly the base and, as a consequence, behaves as a stronger acceptor than molecular iodine; (iii) The intercept indicates a constant effect favoring the associations with ICl.

(4) For a given ring size, the ranking of basicities is lactone < cyclic ketone < lactam < *N*-methyl lactam. This is the order to be expected on the basis of a competition between a destabilizing field-inductive effect and a stabilizing resonance effect in an electron-demanding process, as it is the case for the gas-phase protonation of aliphatic ketones, esters, and amides.[55,56] This behavior similarity of I_2 and ICl is in agreement with the basicity scales of Gutmann[57] and Maria and Gal.[58] Gutmann uses the term of "donicity" as a measure of the ability to donate an electron pair and he has proposed the so-called donor number (DN) to establish an order of base strengths from the enthalpy of the reaction with the acid $SbCl_5$. He assumed that this order remains constant for all other acids, i.e., basicity scales referred to different acids must be linearly correlated. Maria and Gal proposed another scale of Lewis basicity defined as the enthalpy change ($-\Delta H°_{BF3}$) for the reaction between BF_3 and basic organic molecules. To confirm this rule, we have compared our $\Delta_rG°$(ICl,soln) with the DN of Gutmann and with the enthalpies measured by Maria and Gal for which the reference acid is BF_3. The linear relationships found in these cases fulfilled eqs 16 and 17, respectively, and show that DN and $-\Delta H°_{BF3}$ scales are more sensitive to the structural effects than $\Delta_rG°$(ICl,soln)

$$\Delta_rG°(\text{ICl,soln}) = (-0.22 \pm 0.02)\text{DN} + (2.01 \pm 0.37)$$

$$r = 0.981;\ \text{s.d.} = 0.38\ \text{kcal/mol};\ n = 9 \qquad (16)$$

$$\Delta_rG°(\text{ICl,soln}) = (-0.060 \pm 0.006)(-\Delta H°_{BF3}) + (2.52 \pm 0.61)$$

$$r = 0.969;\ \text{s.d.} = 0.46\ \text{kcal/mol};\ n = 7 \qquad (17)$$

These correlations show that, although the association energies change significantly, from quite small values when the reference acid is ICl to much larger values when the reference acids are $SbCl_5$ or BF_3, the different basicity scales follow the same trends.

(5) Notwithstanding their small size, structural effects present several interesting features:

(a) For all lactams, lactones, and cyclic ketones, the stability of the complexes (both with I_2 and ICl) increases with the ring size[19,55] (see Table 1). This is to be expected on the basis of Mulliken's VB model of CT complexes[13,14,59] According to it, the stability of the complex between an electron donor D and an acceptor A is essentially composed by the energetic contributions of the mesomeric structures I and II[13,14]

$$\underset{\text{I}}{\text{D}|\cdots\text{A}} \leftrightarrow \underset{\text{II}}{\text{D}^{\oplus}\cdots\text{A}^{\ominus}}$$

For a given electron acceptor, the former contribution is largely determined by the ionization energy of the base. For a series of carbonyl derivatives, this energy increases with the s character of the carbonyl carbon.We present in Table 4 the results of an NBO analysis of this property for the compounds considered herein. It can be seen that this character significantly decreases (thus becoming a better electron donor) as the size of the ring increases.

TABLE 4: Charge Transferred from Bases (Q_{CT}) to I_2 and ICl Deduced from the NBO Analysis and Polarizability for the Free Compounds (α_D)

	$-Q_{CT}(B-I_2)$	$-Q_{CT}$ (B−ICl)	$\alpha_D(\mathrm{Bohr})^3$
1	0.05325	0.07693	25.8
2	0.05668	0.07715	35.73
3	0.06669	0.08783	45.08
4	0.07080	0.09162	54.22
5	0.07150	0.09230	64.67
6	0.04291	0.06305	36.36
7	0.05464	0.07370	45.97
8	0.06082	0.08013	54.97
9	0.06239	0.08158	64.64
10	0.05876	0.07768	74.38
11	0.03585	0.05246	32.22
12	0.04508	0.05821	39.51
13	0.04558	0.06375	48.73
14	0.04418	0.06216	58.36
15	0.04419	0.06215	68.11
16	0.02179	0.03541	22.71
17	0.02758	0.04426	31.77
18	0.03636	0.05258	41.20
19	0.04302	0.06038	51.14
20	0.04399	0.06151	60.94

The contribution from the “no-bonding” structure I depends on the classical electrostatic and dispersive (London) interactions between D and A. For a given A, the strength of the former interactions increases with μ, the modulus of the molecular dipole moment of D. Again, for a given A, the strength of dispersive interactions increase linearly with the average polarizability, α, of D, as shown by London’s eq 18[59]

$$E_{\mathrm{dispersion}} = -3/2(I_A I_D/I_A + I_D)(\alpha_A \alpha_D/r_{AD}{}^6) \qquad (18)$$

where I_A and I_D can be approximated by the ionization energies of A and D, α_A and α_D stand for the average polarizabilities of A and D, respectively, and r_{AD} is the intermolecular distance. We have used in previous studies[19] the Taft-Topsom parameters σ_α[55,56] as descriptors of substituent polarizabilities. For our present purposes we obviously had to directly compute the α_D values using the static vibrational frequencies[59] computationally determined in this study given in Table 4. These calculated values have recently been proven to be extremely valuable for the study of charge-transfer complexes.[60] Within families, they are seen to smoothly increases with the ring size of D.

From the above, it follows that all of Mulliken’s VB mechanistic factors are consistent with the observed reactivity trend for the three families of carbonyl compounds. This is quite satisfactory, as it is the excellent agreement between the experimental and MO-computed energetics. At this point, however, a quantitative partitioning of the latter according to Mulliken’s model certainly requires a much wider experimental database.

(b) The case of 2-aziridinone (**1**) is interesting, because it behaves both as an oxygen and as a nitrogen base. This molecule is strained and the strain increases with the electronic interaction between the nitrogen lone pair of the nitrogen atom and the antibonding π^* orbital of the carbonyl group, as it is clearly shown by the shortening of the N−C(=O) distances in the structures of the complexes presented in Figure 2 [the N−C(=O) in free (**1**) is about 138 pm]. In general, this interaction favors the electronic transfer to the electron acceptor and the ensuing stabilization of the D−A complex. It is hampered in the case of the complexes of (**1**), whereas its nitrogen lone pair is more readily available for a direct electrophilic attack. This feature is akin to that present in 2,2-dimethylquinuclidin-one.[61,62] Notice however, that, as we have mentioned above, the gas-phase protonation of **1**[18] takes place at the nitrogen and leads to ring opening. In the case of compound **2**, gas-phase protonation occurs at the carbonyl oxygen but the consequence is an increase of the internal strain by some 20 kcal mol^{-1}.[18] In its CT complexes, the most basic site is also the oxygen, but its basicity is significantly lower than that of the homologous compounds **3**−**5**.

(c) As indicated above, in the case of unsubstituted lactams, I_2 (ICl) approaches syn to the NH group, this favoring the interaction between the incoming iodine and the hydrogen atom of this group. In fact, the I···H(NH) distance is about 300 pm in the cases of **3**−**5**. In the case of **2**, this distance is longer (330 pm) than the sum of van der Waals radii of H and I (318 pm). A topological analysis of the charge density of the complexes considered reveals the existence of bond critical points between the incoming iodine and the hydrogen atom of NH group in the case of larger cycles. We have also determined the charge density at the ring critical point (rcp), nicely indicating the existence of NH···I intramolecular hydrogen bond (see Table 5).

More importantly, the variation of the charge density at the rcp, in the case of ICl complexation, seems to correlate very well with the change in the free energies of complexation for the three largest cycles considered in this work. Some previous studies[63−66] indicate that the charge density at the rcp can be a reliable index to measure the strength of the hydrogen-bond interaction. Note that, in the case of I_2 complexation, the hydrogen-bond interaction was found only for the complexes of rings **4** and **5**. The charge density values and those at the rcp are relatively larger in the case of ICl than in that of I_2 complexation. This shows that the hydrogen-bond interaction is relatively more pronounced in the former case. We believe that these are examples of synergy between hydrogen-bonding and charge-transfer interactions. Indeed, as we show in Table 4, the amount of electronic charge transferred from D to A is significantly larger in the case of the syn isomers of the complexes of lactams **3**−**5**. The same happens in the cases of **1** and **2**, although no hydrogen bonding is detected by means of the topological analysis. However, it is likely that electrostatic interactions similar to those reported between the OH group and one of the F atoms in trifluoroethanol [67] are present. In the case of *N*-methyl lactams, the syn isomer of the adducts of **6** involves both a larger transfer of charge and a slightly larger stability of the adducts. The opposite holds for the adducts of compounds **7**−**10**. It seems reasonable to suspect that there is some competition between the stabilizing dispersive interactions

TABLE 5: Charge and Energy Density Evaluated at the Critical Point (cp) and Ring Critical Point (rcp)

	B−ICl				$B-I_2$			
	Cp		Rcp		cp		Rcp	
	ρ	$H(r)$	ρ	$H(r)$	ρ	$H(r)$	ρ	$H(r)$
pyrrolidin-2-one	0.00776	0.00559	0.00723	0.00644				
δ-valerolactam	0.00979	0.00704	0.00881	0.00781	0.00861	0.00611	0.00785	0.00662
ϵ-caprolactam	0.00983	0.00706	0.00883	0.00788	0.00800	0.00507	0.00786	0.00668

and the de-stabilizing steric hindrance originating in the methyl group being close to the nearest iodine atom.

Conclusion

(1) We have constructed a purely experimental database of equilibrium constants for the formation of 1:1 complexes between I_2 or ICl and lactames (N−H and N−Me), lactones, and cyclic ketones in tetrachloromethane solution at 25 °C.

(2) These complexes were examined computationally by means of LANL2DZ* method at HF and MP2(full). The level of consistency between experimental (solution) values and computational (gas phase) results is excellent. Also, the experimental results are fully consistent with the classical VB model of Mulliken.

(3) In the cases of lactones and lactams, the computational study reveals that the adduct species actually present in solution are generally mixtures of the syn and anti conformers (both of them are local minima on the corresponding potential energy surfaces of the complexes), very often in comparable amounts.

(4) For unsubstitued lactams, the syn conformers are generally stabilized through a hydrogen bond between the N−H group and the nearest iodine atom of I_2 or ICl. This entails an increase in the amount of electronic charge transferred from the base to these acidic species and strengthens the hydrogen bond (electronic push−pull). We are not aware of this synergistic effect being reported in the past.

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Supporting Information Available: The total energies evaluated at the MP2(full)/LANL2DZ* level, as well as the scaled ZPE and entropy values, evaluated at the HF/LANL2DZ* level, of different bases and complexes considered in this work. This material is available free of charge via the Internet at http://pubs.acs.org.

References and Notes

(1) Laws, A. P.; Page, M. I. *J. Chem. Soc., Perkin Trans. 2* **1989**, 1577−1581.

(2) Sweet, R. M.; Dahl, L. F. *J. Am. Chem. Soc.* **1970**, *92*, 5489−5507.

(3) Frau, J.; Coll, M.; Munoz, F.; Gracia-Blanco, F. *J. Mol. Struct. (THEOCHEM)* **1991**, *231*, 109−124.

(4) Izaki, K.; Matsuhashi, M.; Strominger, J. L. *J. Biol. Chem.* **1968**, *243*, 3180−3192.

(5) Kaminskaia, N. V.; Spingler, B.; Lippard, S. J. *J. Am. Chem. Soc.* **2000**, *122*, 6411−6422.

(6) Alcami, M.; Mó, O.; Yàñez, M.; Abboud, J.-L. M. *J. Phys. Org. Chem.* **1991**, *4*, 177−191.

(7) Abboud, J.-L. M.; Cañada, T.; Homan, H.; Notario, R.; Cativiela, C.; Villegas, M. D. D.; Bordejé, M. C.; Mó, O.; Yàñez, M. *J. Am. Chem. Soc.* **1992**, *114*, 4728−4736.

(8) DelBene, J. E. *J. Phys. Chem.* **1993**, *97*, 107−110.

(9) Cremer, D.; Dorofeeva, O. A.; Mastryukov, V. S. *J. Mol. Struct. (THEOCHEM)* **1981**, *75*, 225−240.

(10) Marstokk, K. M.; Mollendal, H.; Sandal, S.; Uggerud, E. *Acta Chem. Scand.* **1989**, *43*, 351−363.

(11) Bordejé, M. C.; Mó, O.; Yáñez, M.; Herreros, M.; Abboud, J.-L. M. *J. Am. Chem. Soc*. **1993**, *115*, 7389−7396.

(12) Benesi, H. A.; Hildebrand, J. H. *J. Am. Chem. Soc.* **1949**, *71*, 2703−2707.

(13) Mulliken, R. S. *J. Am. Chem. Soc.* **1950**, *72*, 600−608.

(14) Mulliken, R. S. *J. Am. Chem. Soc.* **1952**, *74*, 811−824.

(15) Drago, R. S. *Applications of Electrostatic-Covalent Models of Chemistry*; Surfside: Gainesville, 1994.

(16) Abboud, J. L.-M.; Notario, R.; Berthelot, M.; Claramunt, R. M.; Capildo, P.; Elguero, J.; El Ghomari, M. J.; Bouab, W.; Mokhlisse, R.; Guihéneuf, G. *J. Am. Chem. Soc.* **1991**, *113*, 7489−7493.

(17) Bouab, W.; Esseffar, M.; Abboud, J.-L. M. *J. Phys. Org. Chem.* **1997**, *10*, 343−346.

(18) Roux, M. V.; Jiménez, P.; Davalos, J. Z.; Castaño, O.; Molina, M. T.; Notario, R.; Abboud, J.-L. M. *J. Am. Chem. Soc.* **1996**, *118*, 12735−12737.

(19) Esseffar, M.; Bouab, W.; Lamsabhi, A.; Abboud, J.-L. M.; Notario, R.; Yàñez, M. *J. Am. Chem. Soc.* **2000**, *122*, 2300−2308.

(20) Lamsabhi, A.; Bouab, W.; Esseffar, M.; Alcami, M.; Yàñez, M.; Abboud, J.-L. M. *New J. Chem.* **2001**, *25*, 509−517.

(21) El Firdoussi, A.; Esseffar, M.; Bouab, W.; Lamsabhi, A.; Abboud, J.-L. M.; Mó, O.; Yàñez, M. *New J. Chem.* **2003**, *27*, 1741−1747.

(22) Riddick, J. A.; Bunger, W. B. *Organic Solvents: Techniques of Chemistry*, 3th ed.; John Wiley: London, 1970; Vol. 2.

(23) Berthelot, M.; Helbert, M.; Laurence, C. *Can. J. Chem.* **1985**, *63*, 958−962.

(24) Guihéneuf, G.; Abboud, J.-L. M.; Bouab, W. *Can. J. Chem.* **1987**, *65*, 2106−2108.

(25) Rose, N. J.; Drago, R. S. *J. Am. Chem. Soc.* **1959**, *81*, 6138−6141.

(26) Frisch, M. J.; Trucks, G. W.; Schlegel, H. B.; Scuseria, G. E.; Robb, M. A.; Cheeseman, J. R.; Zakrzewski, V. G.; Montgomery, J. A., Jr.; Stratmann, R. E.; Burant, J. C.; Dapprich, S.; Millam, J. M.; Daniels, A. D.; Kudin, K. N.; Strain, M. C.; Farkas, O.; Tomasi, J.; Barone, V.; Cossi, M.; Cammi, R.; Mennucci, B.; Pomelli, C.; Adamo, C.; Clifford, S.; Ochterski, J.; Petersson, G. A.; Ayala, P. Y.; Cui, Q.; Morokuma, K.; Malick, D. K.; Rabuck, A. D.; Raghavachari, K.; Foresman, J. B.; Cioslowski, J.; Ortiz, J. V.; Stefanov, B. B.; Liu, G.; Liashenko, A.; Piskorz, P.; Komaromi, I.; Gomperts, R.; Martin, R. L.; Fox, D. J.; Keith, T.; Al-Laham, M. A.; Peng, C. Y.; Nanayakkara, A.; Gonzalez, C.; Challacombe, M.; Gill, P. M. W.; Johnson, B. G.; Chen, W.; Wong, M. W.; Andres, J. L.; Head-Gordon, M.; Replogle, E. S.; Pople, J. A. *Gaussian 98*, revision A.3; Gaussian, Inc.: Pittsburgh, PA, 1998.

(27) Hay, P. J.; Wadt, W. R. *J. Chem. Phys.* **1985**, *82*, 270−283.

(28) Hay, P. J.; Wadt, W. R. *J. Chem. Phys.* **1985**, *82*, 299−310.

(29) Wong, M. W.; Gill, P. M. W.; Nobes, R. H.; Radom, L. *J. Phys. Chem.* **1988**, *92*, 4875−4880.

(30) Reed, A. E.; Curtiss, L. A.; Weinhold, F. *Chem. Rev.* **1988**, *88*, 889−926.

(31) Franch, M. M.; Pietro, W. J.; Hehre, W. J.; Binkley, J. S.; Gordon, M. S.; Defrees, D. J.; Pople, J. A. *J. Chem. Phys.* **1982**, *77*, 3654−3665.

(32) Huzinaga, S. *J. Chem. Phys.* **1965**, *42*, 1293−1302.

(33) Schneider, W.; Thiel, W. *J. Chem. Phys.* **1987**, *86*, 923−936.

(34) Reiling, S.; Bernardi, M.; Bopp, Ph. A. *J. Phys. Chem. A* **1997**, *101*, 4409−4415.

(35) Boys, S. F.; Bernardi, F. *J. Mol. Phys.* **1970**, *19*, 553−560.

(36) Pople, J. A.; Scott, A. P.; Wong, M. W.; Radom, L. *Isr. J. Chem.* **1993**, *33*, 345−550.

(37) Bader, R. F. W.; Bielger-Koning, F. W.; Tang, T. H. *J. Comput. Chem.* **1982**, *3*, 317−328.

(38) Bader, R. F. W.; MacDougall, D. J.; Lau, C. D. H. *J. Am. Chem. Soc.* **1984**, *106*, 1594−1605.

(39) Bader, R. F. W.; Essén, H. *J. Chem. Phys.* **1984**, *80*, 1943−1960.

(40) Bader, R. F. W. *Atoms In Molecules: A Quantum Theory*; Oxford University Press: New York, 1990.

(41) Rafik, C.; Abboud, J.-L. M.; Guihéneuf, G. *J. Org. Chem.* **1983**, *48*, 4761−4763.

(42) Laurence, C.; Guiéneuf, G.; Wojtkowiak, B. *J. Am. Chem. Soc.* **1979**, *101*, 4793−4801.

(43) Bouchoux, G.; Drancourt, D.; Leblanc, D.; Yàñez, M.; Mó, O. *New J. Chem.* **1995**, *19*, 1243−1257.

(44) Bouchoux, G.; Leblanc, D.; Mó, O.; Yàñez, M. *J. Org. Chem.* **1997**, *62*, 8439−8448.

(45) Esseffar, M.; Mó, O.; Yàñez, M. *Mol. Phys.* **2003**, *101*, 1249−1258.

(46) Berthelot, M.; Helbert, M.; Laurence, C. *Can. J. Chem.* **1985**, *63*, 958−962.

(47) Mρller, C.; Plesset, M. S. *Phys. Rev.* **1934**, *46*, 618−622.

(48) Christian, S. D.; Childs, J. D.; Lane, E. H. *J. Am. Chem. Soc.* **1972**, *94*, 6861−6862.

(49) Uruska, I.; Koschmidder, M. *J. Chem. Soc., Perkin Trans. 2* **1987**, 1713−1718.

(50) Hildebrand, J. H.; Prausnitz, J. M.; Scott, R. L. *Regular and Related Solution*; Van Nostrand-Reinhold: New York, 1970.

(51) Barton, A. F. M. *Handbook of Solubility Parameters and Other Cohesion Parameters*, 2nd ed.; CRC Press: Boca Raton, FL, 1991.

(52) Reichardt, C. *Solvents and Solvents Effects in Organic Chemistry*, 3rd ed.; Wiley-VCH: Weinheim, Germany, 2003; Chapter 5.

(53) Durand, A.; Leison, J.-C.; Vigué, J. *J. Chem. Phys.* **1997**, *106*, 477−484.

(54) Bazalgette, G.; White, R.; Trénec, G.; Audouard, E.; Büchner, M.; Vigué, J. *J. Phys. Chem. A* **1998**, *102*, 1098−1105.

(55) Taft, R. W.; Topsom, R. D. *Prog. Phys. Org. Chem.* **1987**, *16*, 1−80.

(56) Hansch, C.; Leo, A.; Taft, R. W. *Chem. Rev.* **1991**, *91*, 165−195.

(57) (a) Mayer, U.; Gutmann, V. *Struct. Bonding (Berlin)* **1972**, *12*, 113−140. (b) Gutmann, V.; Schmid, R. *Coord. Chem. Rev.* **1974**, *12*, 263−293. (c) Gutmann, V. *Coord. Chem. Rev.* **1976**, *18*, 225−255. (d) Gutmann, V.; Resch, G. *Comments Inorg. Chem.* **1982**, *1*, 265−278.

(58) (a) Maria, P.-C.; Gal, J.-F. *J. Phys. Chem.* **1985**, *89*, 1296−1304. (b) Regarding the quantitative comparison between the scales of Gutmann and Maria and Gal, see, e.g., Abboud, J.-L. M.; Notario, R. *Pure Appl. Chem.* **1999**, *71*, 645−718.

(59) Hirschfelder, J. O.; Curtiss, C. F.; Bird, R. B. *Molecular Theory of Gases and Liquids*, 4th ed.; Wiley: New York, 1967; Chapter 13.

(60) Chiappe, C.; Detert, H.; Lenoir, D.; Pomelli, C. S.; Ruasse, M.-F. *J. Am. Chem. Soc.* **2003**, *125*, 2864−2865.

(61) Pracejus, H.; Kehlen, M.; Kehlen, H.; Mastchiner, H. *Tetrahedron* **1965**, 2257−2264.

(62) Pracejus, H. *Chem. Ber.* **1959**, *92*, 988−992.

(63) Gonzalez, L.; Mó, O.; Yàñez, M. *J. Comput. Chem.* **1997**, *18*, 1124−1135.

(64) Gonzalez, L.; Mó, O.; Yàñez, M. *J. Chem. Phys.* **1998**, *109*, 139−150.

(65) Mó, O.; Yàñez, M. *J. Phys. Chem. A* **1998**, *102*, 8174−8181.

(66) Sanz, P.; Mó, O.; Yàñez, M. *Phys. Chem. Chem. Phys.* **2003**, *5*, 2942−2947.

(67) Molina, M. T.; Bouab, W.; Esseffar, M.; Herreros, M.; Notario, R.; Abboud, J.-L. M. *J. Org. Chem.* **1996**, *61*, 5485−5491.