Basicity of lactones and cyclic ketones towards I_2 and ICl. An experimental and theoretical study†

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Intermolecular charge transfer (CT) spectra of several complexes between cyclic ketones and lactones and molecular iodine and iodine monochloride were studied in the UV-visible region. Equilibrium constant and free energy changes of the formed complexes were determined in solution. *Ab initio* calculations at HF/LANL2DZ* and MP2(full)/LANL2DZ* were carried out to establish the nature of the complexation site in the case of lactones, to determine the complex structures and to examine the ring size effect. Although range of the basicity towards I_2 and ICl of compounds studied was small, it was found that cyclic ketones are more basic than lactones. This basicity difference decreases from small to large cycles and practically vanishes for six- and seven-membered rings. A comparative analysis between basicities of lactones and aliphatic esters towards I_2 and ICl has also been carried out. Experimental data in solution were found to be linearly correlated with theoretical results in the gas phase. Proton affinities of cyclic ketones and aliphatic carbonyl compounds that do not present any secondary interaction with the Lewis acid correlate very well with their gas-phase basicity towards I_2 and ICl.

Introduction

Over half a century has elapsed since Mulliken's seminal theoretical work on the electronic structure of charge-transfer (CT) complexes.[1,2] These studies were triggered by the experimental determination of the existence, stoichiometry and thermodynamic stability of 1 : 1 complexes between aromatic hydrocarbons and diiodine in solution.[3] Soon after, X-Ray geometries for benzene–dibromine,[4] benzene–dichlorine[5] and acetone–dichlorine[6] complexes in the solid state were reported. The successful interplay between theory and experiment in the field of CT complexes has continued ever since. In 1969 a substantial body of experimental and theoretical data was already available[7] that allowed a rather satisfactory overview and rationalization of structural and energetic aspects of CT complexes involving both π and n-donor bases.[7] Over the following years, the basic concepts of Mulliken's theory have played a key role in the development of Drago's classical quantitative empirical model of reactivity.[8] Most importantly, they have provided quantitative tools and a conceptual framework (*e.g.*, the electron-transfer paradigm)[9,10] for the study of an impressive array of reactions. Very recently, complexes between diiodine or iodine monochloride and n or π bases have been used to probe at the femtosecond time-scale the dynamics and structure of dative bonding in bimolecular electron-transfer reactions.[11,12] Current computational techniques provide extremely accurate geometrical structures and thermodynamic stabilities of charge-transfer adducts.[13] This information is nicely completed by the structural information on CT complexes in their electronic ground state obtained by state-of-the-art rotational spectroscopy.[14,15]

Consider now reactions (1) and (2), corresponding to the CT association between a base B and diiodine or iodine monochloride in solution in an "inert" solvent (*e.g.* saturated hydrocarbon):

$$B + I_2 \rightarrow B\cdots I_2 \quad K_c(I_2)\ \Delta_r G^\circ_m(1,\ \text{soln}) \qquad (1)$$

$$B + ICl \rightarrow B\cdots ICl \quad K_c(ICl)\ \Delta_r G^\circ_m(2,\ \text{soln}) \qquad (2)$$

Some years ago we explored the possibility for these reactions to provide a "solution basicity ranking"[16] based on the standard Gibbs energy changes for these reactions, $\Delta_r G^\circ{}_m$(1, soln) or $\Delta_r G^\circ{}_m$(2, soln), that could be closely related to the "gas-phase basicity ranking"[17] defined through $\Delta_r G^\circ{}_m$(3, g), the standard Gibbs energy change for reaction (3, g), the protonation in the gas phase:

$$B(g) + H^+(g) \rightarrow BH^+(g) \quad \Delta_r G^\circ_m(3, g) \qquad (3)$$

This relationship between $\Delta_r G^\circ{}_m$(1, soln) or $\Delta_r G^\circ{}_m$(2, soln) and $\Delta_r G^\circ{}_m$(3, g) was expected because of the similarity between these processes and the "dative bonding" mesomeric structure A–D postulated in Mulliken's donor(D)–acceptor (A) model (1) of charge-transfer complexes in their electronic ground state. The basic concept was that the quantitative pattern of structural effects on aqueous solution basicity, reaction (4), is often different from that found in the gas phase because the energetics of this reaction heavily depend on the solvation of

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† Electronic supplementary information (ESI) available: total energies (E_h), zero point energies (ZPE) and entropy values (*S*) of ketones and ketones-I_2/ICl and of lactones and lactones-I_2/ICl. See http://www.rsc.org/suppdata/nj/b3/b305387n/

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the onium ions, mostly because of charge-dispersal to water solvent through hydrogen bonding.[18,19] This effect should not be significant in reactions such as (1) and (2) when conducted in "inert" media.

$$\mathrm{B(aq)} + \mathrm{H^+(aq)} \rightarrow \mathrm{BH^+(aq)} \quad \Delta_r G^\circ_m(4, \mathrm{aq}) \qquad (4)$$

These expectations were fulfilled in this and other studies, as we were able to experimentally show in the cases of $N(sp^2)$, $N(sp^3)$, $S(sp^2)$ and $S(sp^3)$ bases[16] that $\Delta_r G^\circ{}_m(1)$ and $\Delta_r G^\circ{}_m(3)$ are linearly related to a very high degree of precision. In some cases, significant exceptions were found.[16a] They could be quantitatively explained on a structural basis. Not surprisingly, whenever comparisons were possible, excellent correlations were found between $\Delta_r G^\circ{}_m(1, \mathrm{soln})$ and the standard Gibbs energy changes for the formation of adducts between some of these bases and Gutmann's donor numbers (DN), defined as the standard enthalpy changes for the formation of 1 : 1 adducts between the relevant bases B and $SbCl_5$.[20] The same applied to correlations between $\Delta_r G^\circ{}_m(1)$ and $\Delta_r H^\circ{}_{m\mathrm{BF3}}$,[21,22] the latter being the standard enthalpy changes for the formation of 1 : 1 adducts between bases and BF_3 in CH_2Cl_2 solution according to Gal and Maria.[21,22] We have recently performed quantum-mechanical calculations of substantial level on a number of these CT complexes, particularly in the case of the 1 : 1 association between acyclic carbonyl compounds and iodine monochloride.[23] We obtained rewarding agreements with the experimental thermodynamics of the associations and detailed insights into the geometries and other important properties of the complexes. We are now extending this study to the basicities of cyclic carbonyl compounds. These species are very important classes of heterocycles. Suffice to mention here lactames, including, *e.g.*, the time-honored β-lactamic antibiotics, vancomycin[24] a small glycoprotein and important antibiotic[25] as well as the cyclic peptide cyclosporin,[26] a valuable immunosupressive drug.[27] Powerful antibiotics such as erythromycin[28] and the immunosupressive rapamycin[29] incorporate both cyclic ketone and lactone functionalities. These species are presently important targets for metabolic engineering.[30] Particular attention has been devoted to analysing ring-size effects and to comparing the intrinsic basicity of these compounds with the corresponding aliphatic esters.[31,32] Here we essentially deal with the experimental and computational study of basicities of cyclic ketones and lactones towards di-iodine and iodine monochloride. To the best of our knowledge, there are no systematic studies about their complexation with molecular iodine and iodine monochloride, and it seemed of interest to see whether the reactivity patterns observed upon protonation where the acid–base interactions are very strong, are similar to those observed for much weaker interactions. For this purpose we have obtained experimental and computed *ab initio* gas phase standard Gibbs energy and enthalpy changes for the reactions (1) and (2) and we have investigated the structure of the complexes formed.

Experimental section

Absorption spectra were obtained on a Cary 219 spectrophotometer. All measurements were carried out using 1 cm matched quartz cells kept at (25.0 ± 0.1) °C. Compounds and solvent (carbon tetrachloride) were used after further purification according to the methods described in the literature.[33,34] The single isobestic point located between the bands of free and associated ICl or I_2 in the UV-visible spectrum (see Fig. 1) indicates that the CT complexes formed obey to a 1 : 1 stoichiometry as shown in the literature.[35] The calculation of the equilibrium constants, K_c, was performed using Drago's method.[36] For systems which present a small value of K_c, the simplification reported in the literature was applied.[16,37]

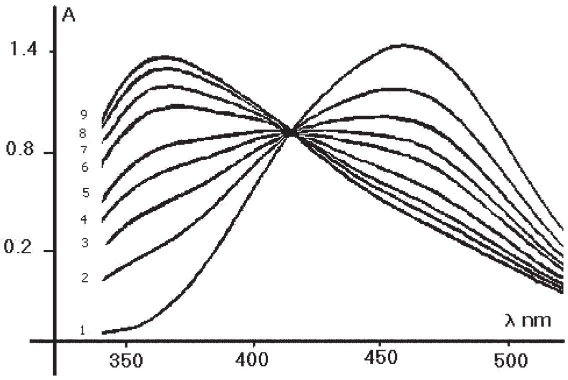

Fig. 1 UV-Visible spectra of the ε-caprolactone–ICl complex (10) the ICl concentration was kept constant at 0.073 mol L^{-1} while the ε-caprolactone concentration was varied: 1, 0; 2, 0.009; 3, 0.012; 4, 0.028; 5, 0.038; 6, 0.059; 7, 0.077; 8, 0.095; 9, 0.113 mol L^{-1}.

Computational details

Ab initio calculations were performed using the Gaussian-98 sets of programs.[38] The structures of the I_2 and ICl complexed as well as those of the isolated bases under study were optimized by using LANL2DZ* basis set which involves an effective core potential (ECP)[39–41] for iodine and chloride but uses the basis set D95V[42] instead of D95[43] for C, N, O and H. This basis set was augmented with a single set of d polarization functions on each heavy atom. The exponents were as follows: $\alpha_C = \alpha_O = \alpha_N = 0.8$ $\alpha_{Cl} = 0.75$, and $\alpha_I = 0.29$.[44–46] This basis set has been recently shown to yield reasonably good results for I_2 complexes with pyridine,[47] thiocarbonyl derivatives[48] and for carbonyl–ICl complexes.[23] For iodine the ECP used incorporates the mass-velocity and Darwin relativistic effects.[40] Spin orbit coupling corrections were not included in our treatment. Nevertheless for the singlet ground state of systems under investigation, first order corrections will be zero. On the other hand, we can assume that these corrections would be similar in the complex as in the isolated atom, because the interactions are very weak, so the effect in the calculated interaction energies should be negligibly small.

Geometry optimizations were initially performed at HF/LANL2DZ* level and then refined at MP2(full)/LANL2DZ* level for the inclusion of electron correlation effects.[49] The harmonic vibrational frequencies were calculated at HF/LANL2DZ* level to characterize the stationary points, to determine the zero point vibrational energy (ZPVE) correction, the thermal correction, the entropy values and to evaluate the frequency shifts due to complex formation with I_2 and ICl. The ZPVE and stretching frequencies were scaled by the empirical factor 0.893. Basis set superposition error (BSSE) was corrected by the counterpoise method of Boys and Bernardi.[50] The information on the charge transfer and electronic charge density was obtained by the natural bond orbital (NBO)[51] method and the atoms in molecule theory of Bader[52–54] respectively.

Results and discussion

The experimental values obtained for the equilibrium constants K_c of different complexes are given in Table 1. The Gibbs energy values of the association between the bases and the acids under consideration as well as those of aliphatic carbonyls,[23,55,56] most of which were taken from the literature, are also listed in this table in order to extend correlations to a large set of compounds and to compare our products with aliphatic systems. The values of $\Delta_r G^\circ{}_m(I_2/\mathrm{ICl})$ (soln) were

Table 1 Experimental values of the equilibrium constants, Gibbs energies and proton affinities for carbonyl–I_2/ICl complexes determined in CCl_4 solvent

	Compound	$K_{c\ I2}$ l mol^{-1}	$\Delta_r G_m{}^\circ{}_{I2}$ kcal mol^{-1}	$K_{c\ ICl}$ l mol^{-1}	$\Delta_r G_m{}^\circ{}_{ICl}$ kcal mol^{-1}	PA kcal mol^{-1}
1	Cyclpropanone	—	—	—	—	—
2	Cyclobutanone	0.44 ± 0.11	0.49 ± 0.15	13.04 ± 0.86	−1.53 ± 0.04	191.8^d
3	Cyclopentanone	0.94 ± 0.18	0.04 ± 0.12	35.00 ± 1.00	−2.12 ± 0.02	196.9^d
4	Cyclohexanone	1.47 ± 0.13	−0.23 ± 0.05	44.00 ± 2.00	−2.26 ± 0.03	201.0^d
5	Cycloheptanone	1.43 ± 0.16	−0.21 ± 0.07	55.17 ± 0.01	−2.39 ± 0.01	202.1^d
6	Oxiran-2-one	—	—	—	—	—
7	β-propiolactone	0.08 ± 0.03	1.50 ± 0.23	0.9 ± 0.1	0.06 ± 0.07	190.2^e
8	γ-butyrolactone	0.52 ± 0.10	0.39 ± 0.12	17.33 ± 0.84	−1.70 ± 0.03	200.8^d
9	δ-valerolactone	0.95 ± 0.02	−0.03 ± 0.02	44.34 ± 0.61	−2.26 ± 0.01	206.2^e
10	ε-caprolactone	1.01 ± 0.12	−0.06 ± 0.07	41.46 ± 0.2	−2.22 ± 0.01	206.7^e
11	$HCON(CH_3)_2$	3.96 ± 0.60^a	−0.82 ± 0.09^a	533 ± 26^c	−3.74 ± 0.03^c	212.1^d
12	$CH_3CON(CH_3)_2$	8.86 ± 0.40^a	−1.30 ± 0.03^a	1696 ± 68^c	−4.43 ± 0.03^c	217.0^d
13	$NCCON(CH_3)_2$	0.27 ± 0.13^b	0.79 ± 0.30^b	2.98 ± 0.2^c	−0.65 ± 0.04^c	198.1^d
14	$(CH_3)_2NCON(CH_3)_2$	8.86 ± 0.90^a	−1.30 ± 0.06^a	1655 ± 84^c	−4.42 ± 0.03^c	222.4^d
15	$(CH_3)_3CCON(CH_3)_2$	4.09 ± 0.40^b	−0.84 ± 0.06^b	384 ± 23^c	−3.55 ± 0.04^c	221.6^d
16	$CH_3OCON(CH_3)_2$	2.31 ± 0.54^b	−0.50 ± 0.14^b	52 ± 4^c	−2.36 ± 0.05^c	209.9^d
17	$H_5C_2OCOCH_3$	0.57 ± 0.03^a	0.34 ± 0.03^a	7.23 ± 0.3^c	−1.18 ± 0.03^c	199.7^d
18	$ClCOCH_3$	0.07 ± 0.07	1.58 ± 0.60^a	0.76 ± 0.4^c	0.16 ± 0.32^c	—
19	H_3CCOCH_3	0.67 ± 0.06^a	0.24 ± 0.05^a	17.8 ± 0.21^c	−1.72 ± 0.01^c	194.0^d
20	$H_5C_2COCH_3$	0.61 ± 0.09^a	0.29 ± 0.09^a	16.8 ± 0.6^c	−1.68 ± 0.02^c	197.7^d
21	$HCOOC_2H_5$	0.29 ± 0.06^a	0.73 ± 0.12^a	1.7 ± 0.25^c	−0.32 ± 0.09^c	191.1^d
22	$CH_3COC(CH_3)_3$	0.68 ± 0.04^a	0.23 ± 0.04^a	23.7 ± 3.5^c	−1.89 ± 0.09^c	200.8^d
23	$HCOCH_3$	—	—	3.34 ± 0.2^c	−0.72 ± 0.03^c	183.7^d
24	$C_6H_5CON(CH_3)_2$	4.92 ± 0.69^a	−0.95 ± 0.08^a	597 ± 32^c	−3.79 ± 0.03^c	222.9^d
25	$CH_3(CH2)_2COOCH_3$	0.60 ± 0.20	0.3 ± 0.2	8.56 ± 2	−1.28 ± 0.20	199.9^d
26	$CH_3(CH_2)_2COOCH_2CH_3$	0.70 ± 0.20	0.21 ± 0.17	9.04 ± 2.5	−1.30 ± 0.20	—

[a] From ref. 55. [b] From ref. 56. [c] From ref. 23. [d] From ref. 17. [e] From ref. 31.

obtained by means of eqn. (5).

$$\Delta_r G^\circ_m(I_2,\ \text{soln}) = -RT\ln K_c \qquad (5)$$

Uncertainties on $\Delta_r G^\circ{}_m(I_2)$ (soln) were estimated by the standard classical method appropriate to these cases. Differentiation of eqn. (5) under constant T (experimental conditions) leads to eqn. (6).

$$\delta(\Delta_r G^\circ_m(I_2,\ \text{soln})) = -RT\delta(\ln K_c) = -RT(\delta\ K_c)/K_c \qquad (6)$$

Because of solubility problems, K_c values are often determined in different media. This problem was solved, as previously,[16,23,55,56] by unifying the data through highly accurate empirical expressions such as eqn. (7)[57]

$$\log K_{XB}(\text{n-}C_7H_{16}) = (0.99 \pm 0.01)\log K_{XB}(CCl_4) + (0.22 \pm 0.01) \quad n = 41,\ r = 0.998,\ \text{s.d.} = 0.11 \qquad (7)$$

We compare first structural effects on the energetics of reactions (1) and (2), simultaneously considering cyclic and acyclic compounds, see Fig. 2. We find that eqn. (8) holds:

$$\log_{10} K_c(\text{ICl}) = (1.62 \pm 0.05)\ \log_{10} K_c(I_2) + (1.62 \pm 0.03) \quad n = 18,\ r = 0.993,\ \text{s. d.} = 0.12 \qquad (8)$$

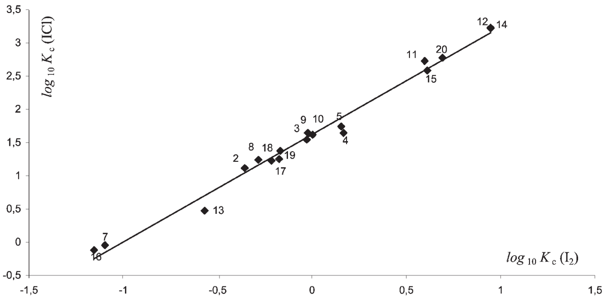

Fig. 2 .Correlation between the experimental values of $\log_{10}K_c$(ICl) and $\log_{10}K_c(I_2)$ for complexed lactones, cyclic and aliphatic ketones.

This result indicates the following:

1. Both cyclic and acyclic carbonyl derivatives examined herein behave as a rather homogeneous family in terms of structural responses to interactions with I_2 and ICl.
2. As already indicated,[23] carbonyl derivatives systematically exhibit a larger basicity towards ICl than towards I_2. Here we confirm that there is a constant effect amounting to *ca.* 2.2 kcal mol^{-1} (as deduced from the intercept in eqn. (8)).
3. Also, the sensitivity to structural effects, as reflected by the slope of eqn. (8) is about 60% larger in the case of the ICl complexes.

The last two results could be anticipated because of the polar nature of iodine monochloride. Its dipole moment, 1.207 ± 0.003 D[58] (in the sense I^+Cl^-)[59] more strongly polarizes the base and, as consequence, behaves as a stronger acceptor than molecular iodine.[14,60]

Concerning the ring size effect, as can be seen from Table 1, the basicity of lactones and cyclic ketones increases progressively with the ring size. This increase vanishes when going to the higher cycles such as ε-caprolactone. The same effect has been found in the case of gas phase protonation[31,32] and in solution towards triethyloxonium ion.[61] As suggested and analyzed in the case of protonation,[31,32] the change in intrinsic basicity of lactones and cyclic ketones, towards I_2 and ICl, with the ring size can be explained by the change of hybridization at the carbonyl carbon atom. In fact, the C(CO)O angle increases on going from β-propiolactone (close to 95°) to ε-caprolactone (close to 120°). This important structural change explains the enhancement of the basicity with the ring size. This increase vanishes in the case of δ-valerolactone and ε-caprolactone that present almost the same hybridization.

Table 1 also shows that, cyclic ketones are more basic towards I_2 and ICl, than their homologue lactones, likely because the inductive effect of the ether oxygen atom reduces the electron donor capacity of the carbonyl oxygen atom of lactones. However, this basicity difference decreases as the size of the cycle increases and becomes practically negligible for six- and seven-membered rings, because the aforementioned effect of the cyclic oxygen decreases as the number of carbon atoms within the cycle increases. Moreover, our theoretical results show that this inductive effect is probably compensated by a specific interaction stabilizing the system (see the next section). This is consistent with the results of the NBO analysis (see Table 2), which shows that cyclic ketones transfer slightly more electronic charge than lactones to the incoming Lewis acid.

The change in the basicity gap with the size of the ring can be envisaged through the use of the isodesmic reactions schematized in Fig. 3. At the MP2(full)/LANL2DZ* level, the corresponding reaction enthalpies are equal to −1.18, −0.70, 0.32 and 0.37 kcal mol^{-1} when the Lewis acid is ICl and −0.81, −0.55, 0.10 and 0.19 kcal mol^{-1} when the Lewis acid is I_2. Thus, for the four and five membered ring, the reaction is exothermic. For the higher homologues, the reaction is slightly endothermic or nearly thermoneutral. Consequently cyclic ketones are more basic than lactones towards I_2 and ICl when four and five membered rings are considered. For six- and seven membered rings, cyclic ketones and lactones have practically the same basicity.

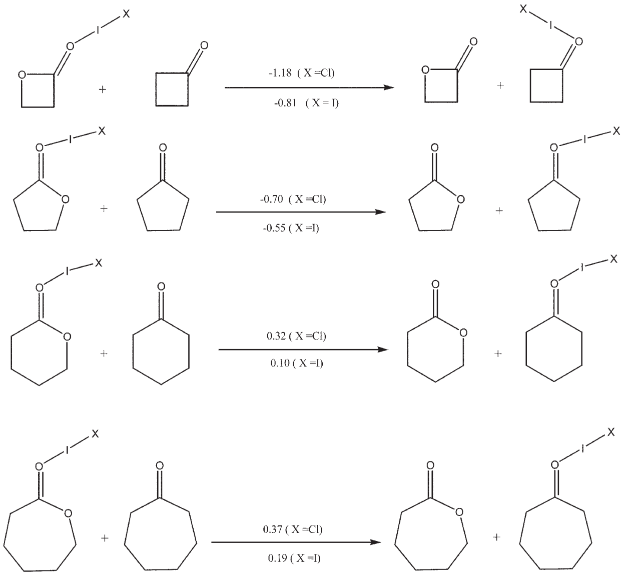

Fig. 3 .Isodesmic reactions for complexation with ICl and I_2, the values are in kcal mol^{-1}.

Structures and relative stabilities

The total energies of the different bases and complexes evaluated at the MP2 (full)/LANL2DZ* level, together with the ZPVE correction and the entropy values are summarized in Tables 1 and 2 of the supporting information. The theoretical complexation enthalpies and free energies of different bases towards I_2 and ICl are summarized in Table 3. The optimized structures of complexes in their most stable conformation, are reported in Fig. 4. For all lactones, the complexation at the oxygen like-ether is the least favorable process. For the larger cycles, the corresponding structure collapses to the carbonyl oxygen attached species.

For five, six and seven membered cycles (with the BSSE and ZPVE corrections taken into account), the *syn*-conformer, defined by the relative position of the I_2 or ICl with respect to the ether-like oxygen, was found to be more stable than the *anti*-conformer. The distance between the incoming halogen and the ring oxygen in the *syn*-conformer decreases progressively on going from five- to seven-membered rings. Hence, the probable explanation is that the *syn*-conformer is stabilized by a weak interaction between the oxygen ether and the first atom of Lewis acid. This interaction may contribute also to the basicity enhancement of larger rings. Finally, the analysis of the basicities of the *anti*-comformers shows that they are lower than those of their homologue cyclic ketones, supporting the inductive effect argument. On the other hand, Wiberg and Waldron,[61] have suggested that the dipole components, at the carbonyl and ether oxygen, are another effect which contributes to the decrease of lactone basicities. They also analyzed and explained why for small rings (oxiran-2-one and β-propiolactone), where the distance between the ether oxygen and the incoming halogen does not favor any interaction, the *syn*-conformation was found to be slightly less stable than the *anti*. Nevertheless, the energy difference between the two conformers is very small, and much smaller than that found in the case of protonation.[31,32] This behavior can be explained by the different nature of the Lewis acids. In

Table 2 Total charge transferred (Q_{CT}) to I_2 and ICl obtained through a NBO analysis

	$-Q_{CT}$(B-I_2)	$-Q_{CT}$ (B-ICl)
1	0.03585	0.05246
2	0.04508	0.05821
3	0.04558	0.06375
4	0.04418	0.06216
5	0.04419	0.06215
6	0.02179	0.03541
7	0.02758	0.04426
8	0.03636	0.05258
9	0.04302	0.06038
10	0.04399	0.06151

Table 3 Calculated values (kcal mol^{-1}) of $\Delta_r\ H^\circ_m(I_2,ICl)$ $T\Delta_rS^\circ_m(I_2,ICl)$ and $\Delta_rG^\circ_m$ (I_2,ICl) at MP2 (full)/LANL2DZ* for all complexes considered

Complex	$\Delta_r\ H^\circ_m(I_2,ICl)$ (calc)[a]	$T\Delta_rS^\circ_m(I_2,ICl)$	$\Delta_rG^\circ_m$ (I_2,ICl) (calc)
1–I_2	−5.73(−7.57)	−6.33	−1.29
1–ICl	−7.37(−9.67)	−7.88	−1.38
2–I_2	−6.31(−8.21)	−6.35	−1.85
2–ICl	−8.15(−10.65)	−8.11	−1.73
3–I_2	−6.55(−8.80)	−6.67	−2.01
3–ICl	−8.57(−11.37)	−8.22	−2.24
4–I_2	−6.64(−8.83)	−6.59	−1.94
4–ICl	−8.70(−11.33)	−8.09	−2.50
5–I_2	−6.66(−8.89)	−6.57	−2.06
5–ICl	−8.75(−11.46)	−8.10	−2.69
6–I_2(*anti*)	−4.85(−6.35)	−5.82	−0.92
6–I_2(*syn*)	−4.74 (−6.09)	−5.63	−1.00
6–ICl (*anti*)	−6.10(−8.00)	−7.35	−0.64
6–ICl (*syn*)	−6.00(−7.76)	−7.23	−0.66
7–I_2(*anti*)	−5.58(−7.41)	−6.22	−1.25
7–I_2(*syn*)	−5.46(−6.99)	−6.11	−1.24
7–ICl (*anti*)	−7.15(−9.47)	−7.75	−1.29
7–ICl (*syn*)	−7.09(−9.06)	−7.43	−1.55
8–I_2(*syn*)	−6.37(−8.15)	−6.48	−1.79
8–I_2(*anti*)	−6.24(−8.26)	−6.43	−1.70
8–ICl (*syn*)	−8.38(−10.64)	−7.95	−2.32
8–ICl (*anti*)	−8.10(10.67)	−8.00	−1.99
9–I_2(*syn*)	−6.63(−8.93)	−6.69	−1.95
9–I_2 (*anti*)	−6.50(−6.62)	−6.47	−1.83
9–ICl (*syn*)	−8.80(−11.65)	−8.17	−2.52
9–ICl (*anti*)	−8.50(−11.19)	−8.02	−2.37
10–I_2(*syn*)	−6.65(−9.08)	−6.73	−1.97
10–I_2(*anti*)	−6.56(−8.68)	−6.45	−1.72
10–ICl (*syn*)	−8.81(−11.84)	−8.21	−2.49
10–I_2(*anti*)	−8.52(−11.19)	−7.99	−2.42

[a] Values in parentheses do not include the BSSE correction.

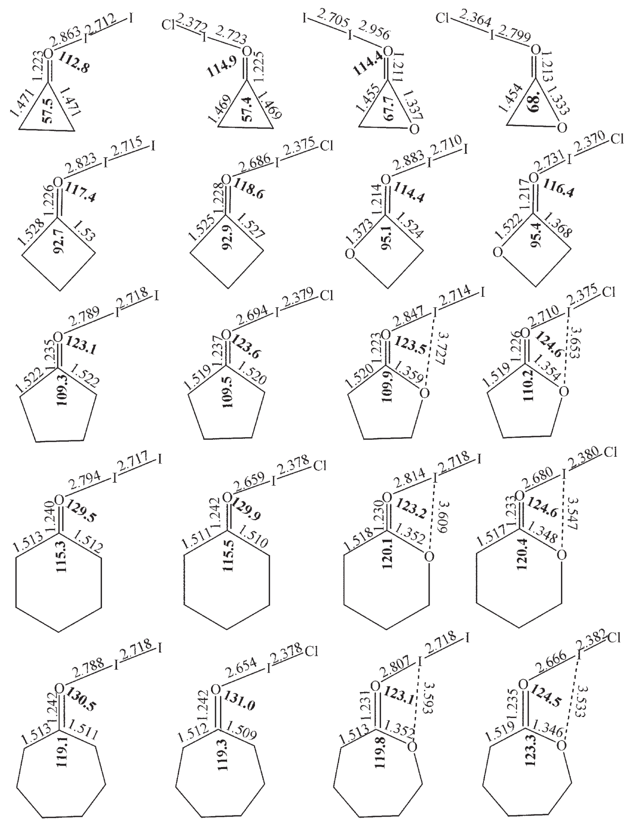

Fig. 4 MP2(full)/LANL2DZ* optimised geometries of the most stable complexes between lactones–I_2/ICl and ketones–I_2/ICl.

the case of I_2 and ICl, the interaction between the incoming halogen and the ether oxygen atom is relatively weaker than that found in the case of proton.

The O–I bond length is on average around 2.714 Å for complexes with ICl and 2.905 Å in the case of complexation with I_2 which are less than the sum of the Van der Waals radii of O and I (3.2 Å). The smaller distance O–I in the case of complexation with ICl indicates that the interaction with I_2 is relatively weaker than with ICl. The angle formed by the ICl (I_2) moiety with the C=O bond is in the range 120–130°. This clearly indicates that, in cyclic ketones–ICl (I_2) and lactones–ICl (I_2), the interaction takes place between iodine and the lone pair of the oxygen atom. In fact, as we have already mentioned, for both lactones and ketones, the complexation at the carbonyl group induces an electron transfer from the lone pairs of the carbonyl oxygen to empty antibonding orbitals of the Lewis acid. This is necessarily reflected in a lengthening of the I–I and I–Cl bond lengths. As a matter of fact, the I–I and I–Cl bond lengths in the complexes are on average about 2.718 Å and 2.378 Å respectively slightly longer than the I–I and I–Cl in the free molecules (2.697 Å and 2.351 Å) calculated at the same level.

On the other hand, the aforementioned charge transfer enhances the electronegativity of the oxygen atom which depopulates slightly the C=O bond. Consistently, upon complexation, the elongation of C=O bond is about 0.009 Å for lactone–I_2 complexes and 0.012 Å for lactone–ICl complexes. These elongations are also reflected in the values of the C=O stretching frequencies. In fact, Table 4 shows that the stretching frequencies appear systematically red-shifted for the two series of compounds, this effect being more important in the case of complexation with ICl. Also interestingly, these frequency shifts, $\Delta\nu_{CO}$, increase with the ring size in all cases (from 12 cm^{-1} to 19 cm^{-1} for ketones–I_2 and from 21 cm^{-1} to 32 cm^{-1} for ketones–ICl and from 16 cm^{-1} to 33 cm^{-1} for lactones–I_2 and from 25 cm^{-1} to 53 cm^{-1} for lactones–ICl), and therefore they follow the same trend as the basicity.

Table 4 Frequencies of the $\nu_{C=O}$ band evaluated at HF/LANL2DZ * and its displacements (in cm^{-1}). The calculated values were scaled by the empirical factor 0.893

	$\nu_{C=O}$(B)	$\nu_{C=O}$(B–I_2)	$\nu_{C=O}$(B–ICl)	$\Delta\nu_{C=O}$(B–I_2)	$\Delta\nu_{C=O}$(B–ICl)
1	1920	1909	1901	11	19
2	1853	1837	1830	16	25
3	1810	1793	1782	17	29
4	1790	1773	1761	17	29
5	1774	1759	1748	15	27
6	1979	1965	1747	15	24
7	1886	1864	1656	22	32
8	1883	1808	1599	25	42
9	1800	1771	1564	29	47
10	1795	1767	1563	28	45

It can also be noted that, for all compounds, the harmonic CO stretching decreases on going from three to seven membered rings for the uncomplexed cyclic ketones and lactones, in good agreement with the variation of the hybridization with the ring size. In all cases the Cl–I···O or I–I···O arrangements are essentially linear with an angle close to 180°.

Experimental *vs.* theoretical values

To compare experimental and calculated free energies it must be taken into account that both sets of values refer to different standard states and therefore a direct comparison is not possible. Indeed, the experimental measurements are carried out in solution and the standard state is 1 mol l^{-1}. Calculated values correspond to 1 atm and 298.15 K gas-phase conditions which, assuming an ideal behavior, correspond to a concentration of 4.09 10^{-2} mol l^{-1}. Taking the standard state as 1 mol l^{-1}, the chemical potential of an ideal species X at T (K) is simply $\mu_X = \mu_X^0$. The chemical potential for the same species at the same temperature, but at a concentration c is given by:

$$\mu_X = \mu_X^0 + RT\ln c \tag{9}$$[62]

Thus, taking into account that $c = 4.09\ 10^{-2}$ mol l^{-1}, one gets:

$$\mu_X - \mu_X^0 = -1.89 \text{ kcal mol}^{-1} \tag{10}$$

and therefore

$$\Delta_r G_m^0(1, \text{gas}) = \Delta_r G_m^0(1, \text{soln}) + 1.89 \tag{11}$$

In other words, one should expect the calculated values (corrected by the 1.89 $kcal.mol^{-1}$ factor), reported in Table 3, to be more or less close to the experimental ones. The agreement is not very good because the calculated free energies will be affected by a much larger error, because of the sizable error affecting the calculated entropy for these weakly bound complexes. Indeed, these complexes present several low-frequency vibrational modes which are highly anharmonic and therefore are not adequately described by using the harmonic approximation routinely employed in *ab initio* calculations.[49] Hence, we have decided to use calculated enthalpies instead for our comparison with experimental free energies. Both magnitudes as linearly correlated as shown in Figs. 5 and 6. These correlations fulfill the equations:

$$\Delta_r G_m^\circ(\text{ICl}, \text{soln}) = (1.34 \pm 0.05)\ \Delta_r H_m^\circ(\text{ICl}, \text{calc}) + (9.5 \pm 0.4) \tag{12}$$

with $n = 17$, $r = 0.992$ and s.d. = 0.16 kcal mol^{-1}.

$$\Delta_r G_m^\circ(I_2, \text{soln}) = (1.5 \pm 0.1)\ \Delta_r H_m^\circ(I_2, \text{calc}) + (10.2 \pm 0.5) \tag{13}$$

with $n = 8$, $r = 0.991$ and s.d. = 0.04 kcal mol^{-1}.

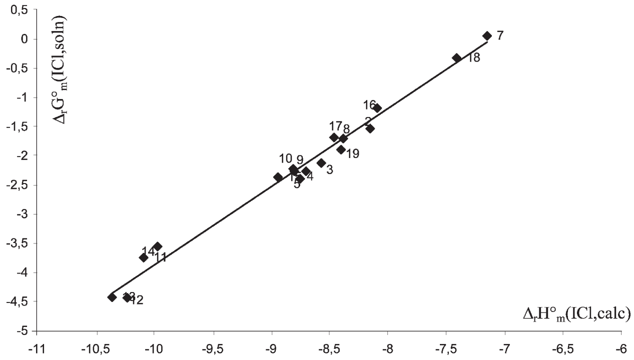

Fig. 5 Correlation between $\Delta_r G^\circ_m$(ICl,soln) and $\Delta_r H^\circ_m$(ICl) calculated at MP2/LANL2DZ* level. Values of points 11–19 are taken from ref. 23.

Comparison with gas-phase protonation

The aim of this section is to investigate whether the basicity of lactones and ketones with respect to I_2 and ICl in solution follows similar trends as the gas-phase basicity of these compounds with regards to H^+ in the gas phase.[17] Table 1 seems to indicate that in the case of protonation, cyclic ketones are less basic than their counterpart lactones, except for β-propiolactone/cyclobutanone couple. In fact, protonated lactones take part of their stability from the favorable interaction between the incoming proton and the ether oxygen atom.[31,32,61,63] This interaction is much stronger than the possible one between the incoming halogen and the ether-like oxygen. As a matter of fact, our theoretical calculations show that the distance between iodine and the ether-like oxygen is relatively longer than the sum of the van der Waals radii of I and O, indicating that the interaction between I and O in the case of complexation with I_2 or ICl is very weak, although it increases with the size of the ring, as reflected in shorter O–I distances. Therefore, to compare gas-phase basicities and $\Delta_r G^\circ_m(I_2,\mathrm{soln})$ or $\Delta_r G^\circ_m(\mathrm{ICl,soln})$ it is necessary to exclude those compounds that present specific interaction between the ether oxygen atom and the Lewis acid, as well as those where the incoming I_2 or ICl approaches perpendicularly to the carbonyl group.[23] For this reduced set of compounds the correlations found are reasonably good (see Figs. 7 and 8).

$$\Delta_r G^\circ_m(\mathrm{ICl,soln}) = (-0.11 \pm 0.01)\ \mathrm{PA} + (19.84 \pm 1.54) \tag{14}$$

with $n = 9$, $r = 0.983$, s.d $= 0.22$ kcal mol^{-1}

$$\Delta_r G^\circ_m(I_2,\mathrm{soln}) = (-0.07 \pm 0.01)\ \mathrm{PA} + (13.56 \pm 1.07) \tag{15}$$

with $n = 8$, $r = 0.993$, s.d. $= 0.09$ kcal mol^{-1}.

It is also worth noting that lactones are more basic than their homologue aliphatic esters with respect to I_2, ICl and H^+ except for β-propiolactone, although as far as the basicity towards I_2 is concerned the differences are quite small.

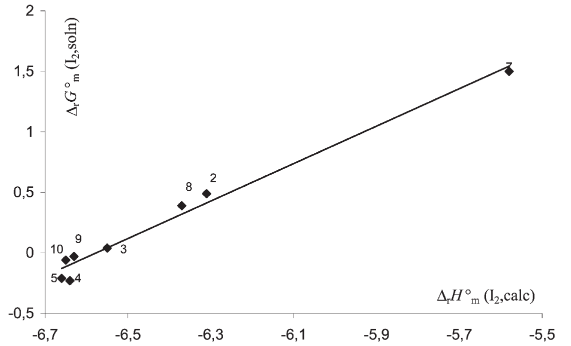

Fig. 6 Correlation between $\Delta_r G^\circ_m$ (I_2,soln) and $\Delta_r H^\circ_m$ (I_2) calculated at MP2/LANL2DZ* level.

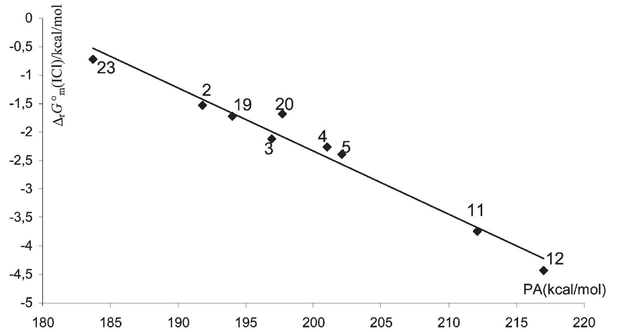

Fig. 7 Correlation between the experimental values of $\Delta_r G^\circ_m$ (ICl, soln) and PA.

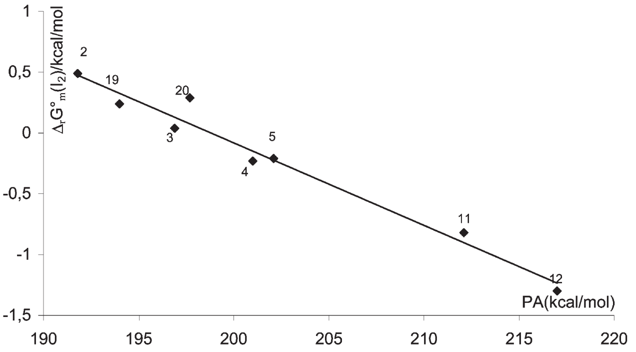

Fig. 8 Correlation between the experimental values of $\Delta_r G^\circ_m$ (I_2,soln) and PA.

Conclusions

The first significant result is that, as was found previously for aliphatic carbonyl compounds,[23] the basicity of cyclic ketones and lactones with respect to ICl is systematically larger than that exhibited with respect to I_2.

Cyclic ketones are more basic than lactones with respect to molecular iodine and iodine monochloride for three, four and five membered cycles. For the larger cycles, they have almost the same basicity.

The good correlation between the experimental proton affinities and experimental free energies in solution with respect to I_2 and ICl seems to indicate that the size effect discussed in the literature[31] for the gas-phase protonation processes also applies to the behavior of these bases in solution.

In the case of lactones, it was found that for five, six and seven membered rings the *syn*-conformer is lightly more stable than the *anti*-one because of favorable dipole interactions and the weak interaction between the ether oxygen and the incoming halogen.

There is a good agreement between *ab initio* calculated values, which will typically correspond to the gas-phase situation and the experimental values obtained in solution. This good agreement can be taken as an indirect indication of the reliability of the theoretical scheme used and, by extension, of the optimized structures employed to obtain these energies. It also indicates that the effects on the basicity of lactones and cyclic ketones with respect to I_2 and ICl are similar in the gas phase and in solution.

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