Small cyclophanes: the bent benzene business

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Abstract - Short bridges in [n]meta- and [n]paracyclophanes with n < 8 bend the benzene ring into a boat shape. Naturally, this influences the properties of the aromatic system, but two seemingly contradictory trends are observed. On the one hand, the chemical reactivity of such small cyclophanes is strongly increased and quite unusual; sometimes, it is remindful of that of a localized cyclohexatriene-like system. On the other hand, aromatic delocalization appears to be hardly impaired as judged from aromatic carbon-carbon bond lengths and from ring current effects such as proton chemical shifts and anisotropy of the magnetic susceptibility. In line with recent theoretical analyses implying that delocalization in benzene is due to the σ-system and counteracted by the π-system, it is proposed that bent benzene rings are full-fledged aromatics, and that their sometimes olefin-like reactivity is predominantly a consequence of their high strain energy being dramatically relieved in the initial stages of such reactions.

INTRODUCTION

Benzene is one of the most important molecules. This proposition holds for the entire and rather broad range stretching from economic value and industrial volume via experimental organic chemistry to structural chemistry and theory of chemical bonding.

It is therefore not surprising that benzene is one of the most intensively investigated compounds, and its properties and structure appear to be well-known indeed: the benzene molecule C_6H_6 is a rather compact disk which is about twice as long as thick, a regular hexagon with D_{6h} symmetry involving six equal carbon-carbon (1.396 Å) and carbon-hydrogen bonds (1.085 Å [1]) and a bulky cloud of π-electrons above and below the six-membered ring (Fig. 1).
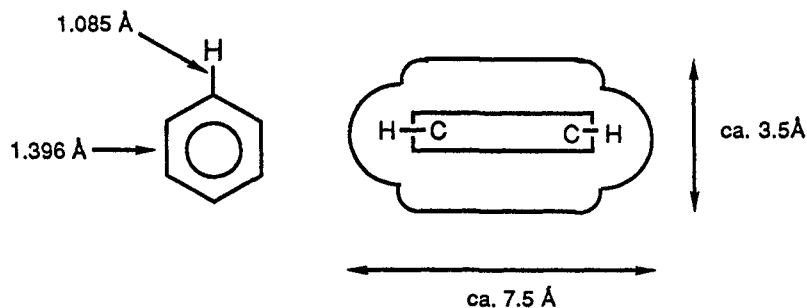

Fig. 1. Experimentally determined structure of benzene (IR, X-ray, electron diffraction)

The theoretical description of the electronic structure of benzene, too, seems to be well established[2]: in Valence Bond theory, resonance between two Kekulé, three Dewar-type and several higher energy structures is used to describe the total electron distribution (Fig. 2a); in (Hückel) Molecular Orbital theory, a σ-skeleton connects the six sp^2 hybridized carbon atoms and the six hydrogens, while a π-system is obtained by lateral overlap of the remaining six carbon p_z-orbitals (Fig. 2b).
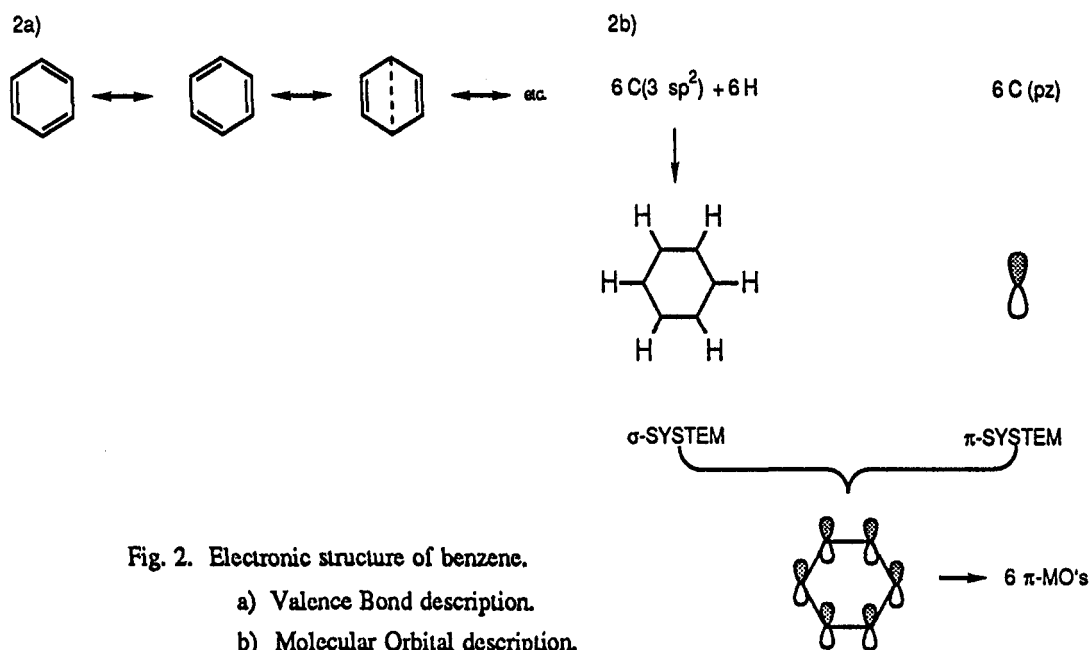
2a)   2b)

$6 C(3 sp^2) + 6 H$ $6 C (pz)$

σ-SYSTEM   π-SYSTEM

Fig. 2. Electronic structure of benzene.

a) Valence Bond description.

b) Molecular Orbital description.

⟶ 6 π-MO's

This relatively simple, satisfactory and peaceful picture has recently become increasingly perturbed by controversial discussions arising primarily from the camp of theoreticians. Rather than trying to be comprehensive, three examples will be used to illustrate different kinds of challenges which have been brought forth. In the first place, provocative and potentially the most far-reaching: the simple and generally adhered idea has been questioned[3] that benzene is regular and delocalized *because* this is beneficial for and required by the π-electrons (although this traditional concept of Erich Hückel still has contemporary support[4]). Rather, a detailed analysis of the effects of bond length variations on the σ- and the π-system[3c] has been interpreted as indicating that it is the σ-system which enforces delocalization, i.e. equality of bond lengths, whereas the π-system would tend to localize and favour alternating short and long C-C bonds, in other words, a cyclohexatriene structure. Atrocious as this may sound at first to the practising organic chemist, indoctrinated as he is by traditional teaching, the new paradigm appears more credible to the simple mind when one realizes that the π-component of a C=C bond indeed experiences increasingly better overlap between the two p_z orbitals on shortening of the C-C distance; even in ethene, the actual bond length of 1.34 Å is a compromise between this tendency and that of the σ-component to be longer, because in pure single bonds between two sp²-hybridized carbon atoms, the preferred bond length is about 1.47 Å. This is illustrated in a qualitative fashion in Fig. 3. The final solution of this controversy must, of course, come from high quality theoretical calculations. Independent from the outcome of this analysis, it should be pointed out that there is no doubt on the π-system in benzene being in fact delocalized[3c]; the question is rather whether π-delocalization is the driving force for delocalization, as commonly assumed, or a "byproduct" resulting from the delocalization of the σ-system[3c].

π-BOND :

sp²-σ-BOND :

1.47 Å   ⟶   1.34 Å

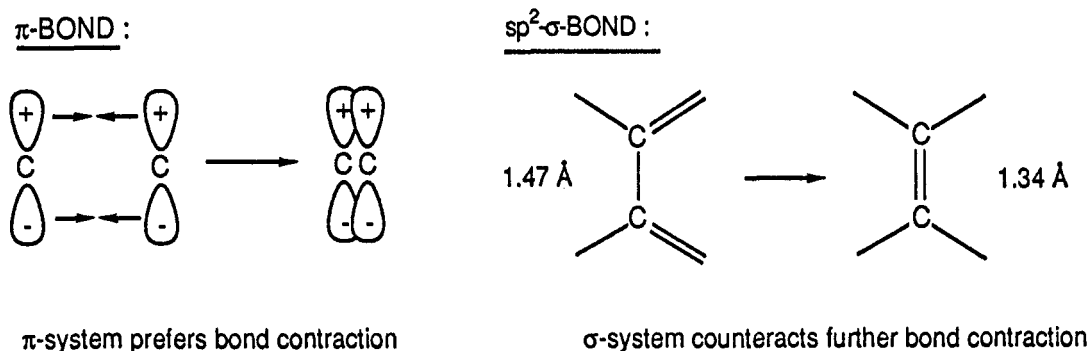π-system prefers bond contraction   σ-system counteracts further bond contraction

Fig. 3. Influence of the π- and σ-system on the length of a C=C bond

Without further comment, we mention as a second indication of commotion around the structure of benzene a theoretical paper which questions the reality of π-bonding altogether and reports a greater stability for benzene when banana bonds are introduced[5] (Fig. 4).

A third and completely different point had been raised by Ermer[6] who showed that contrary to general belief, the experimental proof for the D_{6h} structure of benzene as apposed to a localized Kekulé or cyclohexatriene structure with D_{3h} symmetry *cannot* be derived from X-ray crystal structure data because in most cases, they are insufficiently accurate to distinguish between the two. The most reliable *experimental* evidence for the D_{6h} symmetry comes from the vibrational analysis of (deuterated) benzene by Ingold many years ago.[6,8] Against the background of all these question marks, it is reassuring that the benzene molecule has recently been "seen" by scanning tunnel microscopy,[9] even though this visible proof of its existence does not help us in distinguishing between the localized and the delocalized structure due to insufficient resolution and - more seriously - to the fact that the observed benzene molecules were absorbed on a rhodium surface together with carbon monoxide in such a way that the benzene geometry is, by its environment, distorted towards D_{3h} symmetry.
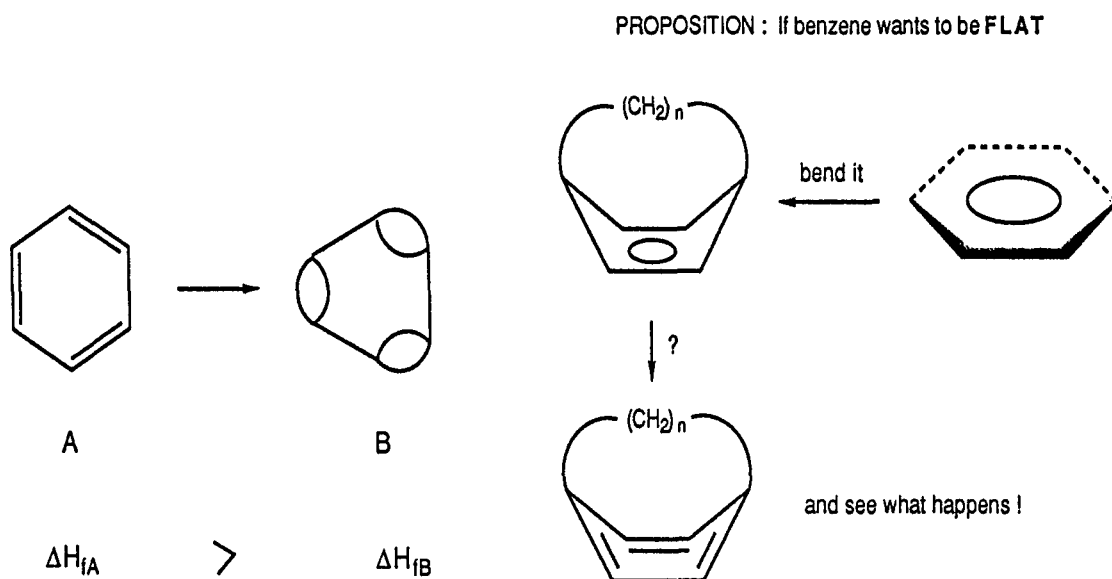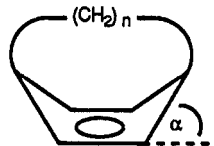

Fig. 4. Banana bonds in benzene

Fig. 5. Small [n]paracyclophanes as probes for the delocalization in benzene

SMALL CYCLOPHANES AS PROBES FOR BENZENE

We have been investigating many aspects of the chemistry of small [n]meta- and [n]paracyclophanes (n < 8) for several years.[10] In the course of these investigations, it occurred to us, that such small cyclophanes might be probes which give us experimental information on the matter of delocalization in benzene. Our considerations are briefly outlined in Fig. 5: if benzene is flat because of the π-system, one may try to bend the ring by a short oligomethylene bridge across the *para*-positions. This must cause weakening of the π-system, and *if* the π-system is really causing delocalization, the latter will be weakened, too, and the structure may collaps to that of a cyclohexatriene. The theoretical data of Table 1 indicate that considerable bending and strain may be expected; however, a *caveat* must be derived from the last column, as the variation of carbon-carbon bond lengths in the benzene ring is not straightforward and, moreover, not indicative of alternating bonds.

Table 1. Some theoretical predictions on [n]paracyclophanes

n	α^a [°]	STRAIN ENERGY ^{a} [kcal/mol]	$\Delta r(C-C)^b$ [Å]
8	8	-	0.007
7	13	32	0.011
6	17	44	0.012
5	23	63	0.047
4	28	88	0.027

^{a} MNDO. b STO-3G ; Δr indicates the maximum variation in aromatic bond lengths.

Of course, in order to study small cyclophanes, one first must make them ; this is not a trivial task, especially for the smaller members. As indicated in Fig. 6, the normal strategy of closing the bridge over a preexisting benzene ring will fail because the ends will not meet above the planar ring. Therefore, an alternative approach must be chosen of preparing a precursor which contains the intact bridge crossing an entity A which can be transformed afterwards to the desired benzene ring.
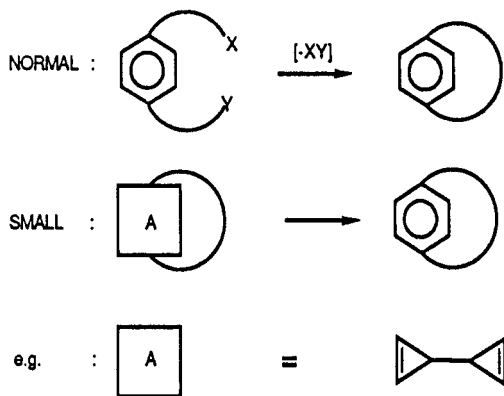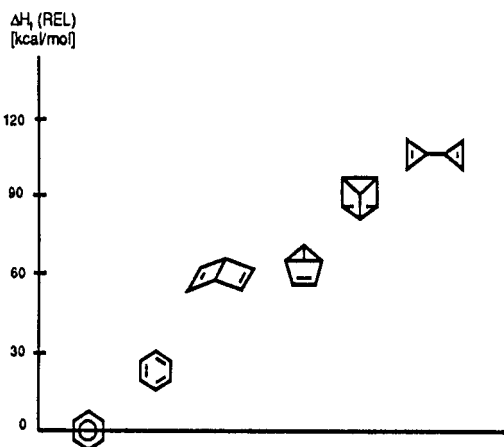
Fig. 6. Strategies for small cyclophanes

Fig. 7. Relative energies of valence isomers of benzene

Several groups have realized this approach using different entities A.[13] Our choice[10] for A has been 3,3'-bicyclopropenyl for the following reasons. In the first place, bicyclopropenyl is not only one of the 217 structural isomers of the composition C_6H_6 but, like benzene itself, it belongs also to the much smaller group of 6 isomers with a 1:1 bonding relation between one carbon and one hydrogen $(CH)_6$, the so-called valence isomers of benzene. Moreover, as illustrated in Fig. 7, bicyclopropenyl, while possessing the least geometrical similarity to benzene, is the valence isomer with the highest energy content and may therefore be expected to furnish easy downhill pathways to other highly energetic (valence) isomers or to strained, shortbridged cyclophanes.

The synthesis of the bridged bicyclopropenyls 2 from 1,2-bismethylenecycloalkanes 1 has been described elsewhere, as has the conversion of 2 to the 1,2-bridged and 1,4-bridged Dewar benzenes 3 and 4[10] (eq. 1).
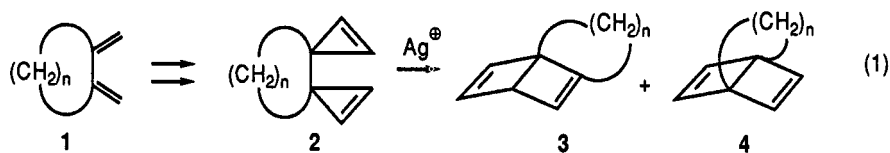
Not surprisingly, the ease of conversion of the Dewar benzenes **4** to the corresponding [n]paracyclophanes **5** depends on the length n of the bridge. Thus, **4a** (n = 6) is completely converted to **5a** under conditions which are typical for a Dewar benzene, i.e. warming to 60°C (eq. 2); **5a** is cleanly reconverted to **4a** by irradiation, which is not typical for a benzene derivative (cf. xylene !) and signals strain in **5a**.[14] In contrast, **4b** is stable in solution up to 150°C and pyrolyses to **6**, presumably via **5b**, on flash vacuum thermolysis only at 280°C (eq. 3a).[10] Irradiation of **4b** below -20°C leads to a photostationary equilibrium involving **4b** and **5b** in a ratio of 93 : 7.[15] (eq. 3b). Contrary to **5a**, **5b** is not stable at room temperature.
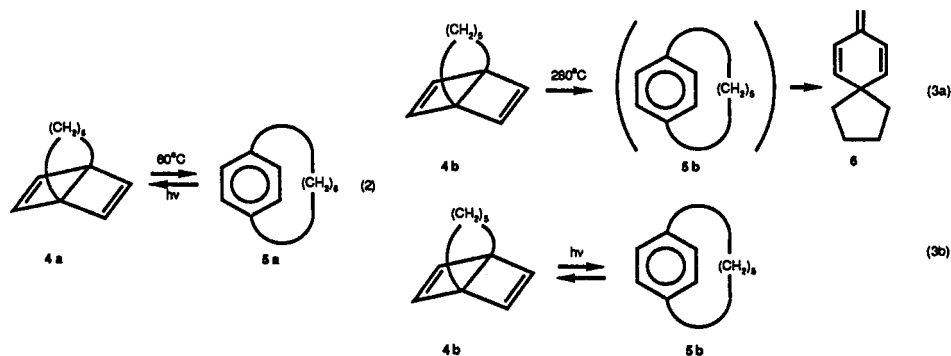
Even more elusive is [4]paracyclophane (**5c**) (Fig. 8). It needs harsher conditions for its thermal formation from **4c** (300°C), and fragments directly to ethene and *para*-xylylene (**7**).[10] Formed from **4c** on irradiation in solution, **5c** polymerizes even at -60°C. It has been identified by matrix UV-spectroscopy,[16] and was intercepted by protonation and addition of a nucleophile to give **9**.[17a] It has been estimated that the formal "basicity" of **5c** is so high that its conjugate acid **8** has a pK$_A$ of 37[17b] (cf. the pK$_A$ of normal protonated aromatics analogous to **8** of about -6[18]). Of course, this unusual basicity is not so much due to any unusual electronic properties of **5c**, but rather to the tremendous relief of strain in its transformation to **8**.
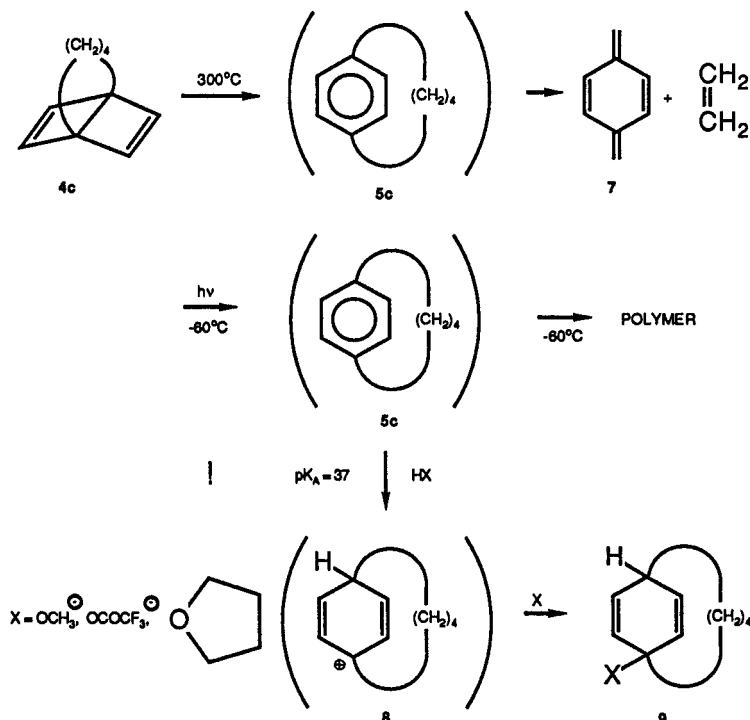

Fig. 8. Formation and reactions of [4]paracyclophane (**5c**).

In the course of our investigations on small [n]paracyclophanes, we also obtained - initially to our surprise[19], later in a systematic fashion[10] - a number of (halogenated) [5]metacyclophanes 10 and the 2,6-tetramethylene Dewar benzene 11; the latter turned out to be a precursor of [4]metacyclophanes (12). These strained cyclophanes also played an important role in shedding light on the aromaticity of bent benzene rings.[10]

X = Y = H
X = Y = Cl
X = Y = Br
X = Br, Y = Cl

10   **12**   **11**
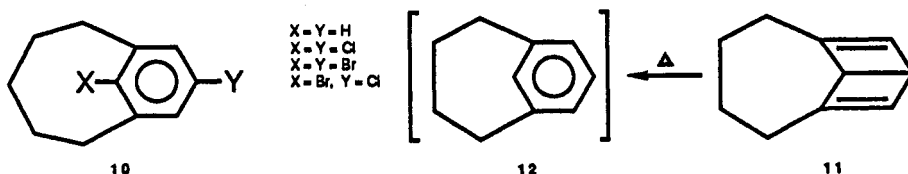
ARGUMENTS AGAINST AROMATICITY

Small cyclophanes, whether of the *meta*- or the *para*-series, have boat-shape bent benzene rings and a high reactivity which may be quite atypical for aromatic compounds.[10,13] As an example, and without a detailed mechanistic analysis,[10] we present the reaction of 10 with *tert*-butyllithium as shown in Fig. 9. Compound 10b leads to what is formally an S_N2Ar substitution of the *most hindered* chlorine to furnish 13; note that the other chlorine behaves normal by doing nothing under these conditions. Substitution of one chlorine by bromine (10c) leads to a completely different reactivity pattern yielding 14. The apparently innocent transition from 10c to 10d results in "normal" halogen-metal exchange. Incidentally, the D_2O-quench of 15 gave 16; the latter proved to be of great value as a witness for aromaticity (*vide infra*).

10b   t-BuLi   **13**   **10d**   t-BuLi   **15**

D₂O

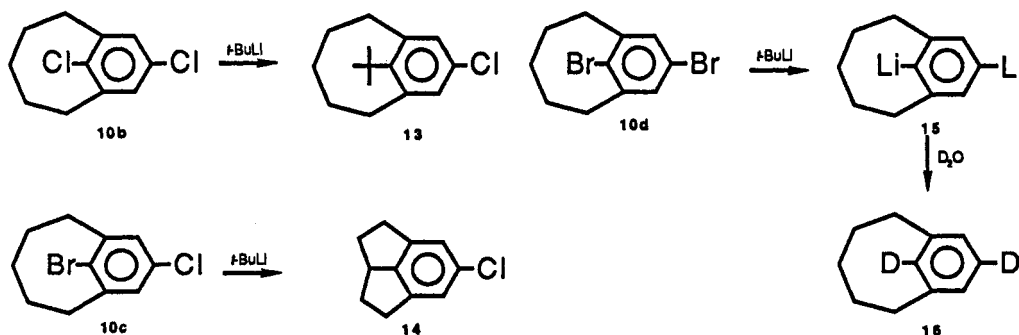**10c**   t-BuLi   **14**   **16**

Fig. 9. Reactivity of dihalo[5]metacyclophanes.

While the reactions of Fig. 9 are "non-aromatic" in an unspecific way, the Diels Alder reaction under mild conditions is usually considered to be highly indicative of olefinic behaviour. Thus, the mild conversion of 10a to 17 by the weak dienophile dimethyl acetylenedicarboxylate seems to signal cyclohexatriene reactivity of 10a (eq. 4); maleic anhydride gives the Diels Alder adduct instantaneously.[10]

COOMe
MeOOC

+ MeOOC-C≡C-COOMe   20°C / 5 h   100%   (4)

10 a   **17**
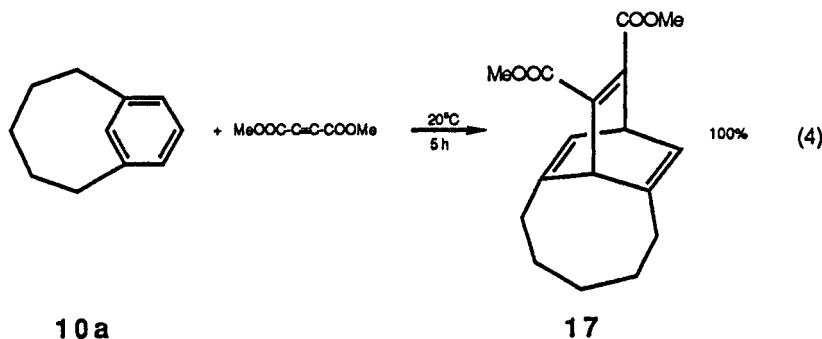
Although it may not be obvious at first sight, the amazing dimerization of 11 to 18 and 19 (eq. 5) starts off with a Diels Alder reaction between 11 and 12 (eq. 6);[10,20a] the primary adducts 20 and 21 were isolated, characterized and shown to convert to 18 and 19 (X,Y = H,D).[20b] Again, the high reactivity of 12 towards the *unactivated* double bond in 11 is without precedent in aromatic chemistry. Thus we, and others, began considering the possibility of Kekulé-type bond fixation in cyclophanes as illustrated in Fig. 5. However, there were always doubts, too.
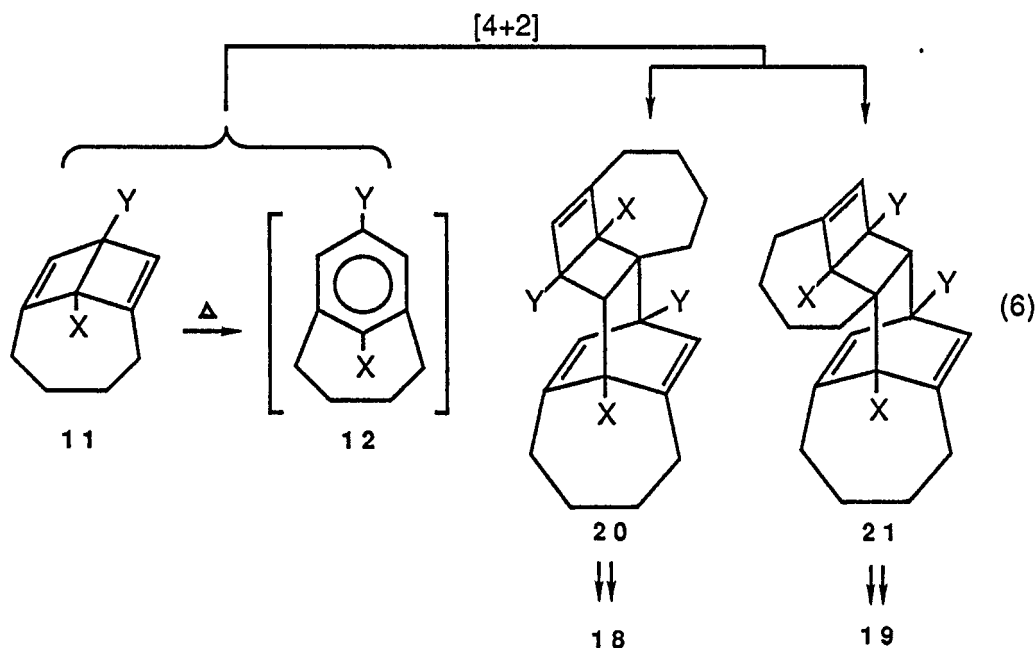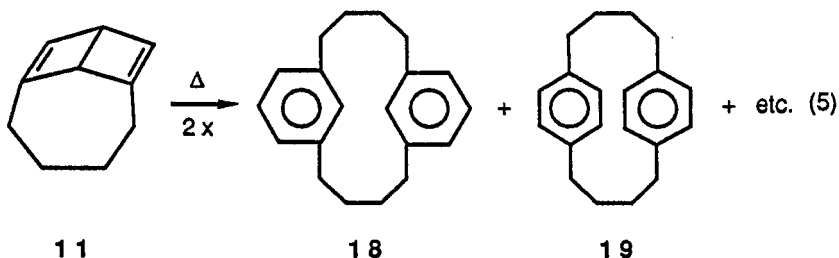
ARGUMENTS FOR AROMATICITY

In the first place, there were the ^{1}H-NMR spectra, e.g. of those of 5b and of 10, which had two characteristic features. The aromatic protons are found in the typical deshielded aromatic region of δ = 6.75 - 7.85 ppm, which does not fit very well with an olefinic character. More compelling evidence comes from the strongly shielded shifts of those bridge protons which lie in the shielding cone of the aromatic ring (δ close to 0 ppm), which testifies to a considerable aromatic ring current.[10] As one may argue that chemical shifts are complex entities which may be influenced by factors difficult to analyze, we have successfully tried to assess the ring current in an unambiguous and quantitative fashion, namely by determining the anisotropy of the magnetic susceptibility $\Delta\chi$ by means of the quadrupolar deuterium couplings at very high field (14.1 Tesla): compared to a suitable planar aromatic analogue ([2-D]-1,3-diisopropylbenzene, $10^{28}\Delta\chi$ = -1.35 cm^3), compound 16 ($10^{28}\Delta\chi$ = -1.5 cm^3) turned out to be at least equally "aromatic". Although the limit of error of the experimental values is of the order of 10-20%, they are clearly above those of a localized model by a factor 3 to 4.[21] If one takes the value for 16 as reliable, one might even - with appropriate caution ! - come to believe that the bent benzene ring of 16 is slightly *more aromatic* than a planar one.

Equality of carbon-carbon bond lengths near the typical intermediate value of benzene (d = 1.39 ± 0.01 Å) is another strong criterium for aromatic delocalization. One of the most strained cyclophanes for which an X-ray crystal structure is available, is 10b. Its bond distances (Fig. 10) are practically identical to those of benzene itself; the small differences may be largely caused by substituent effects and are, moreover, not in line with a cyclohexatriene-type bond fixation, as the molecule has C_s symmetry.[22] Structural information available on other, less strained cyclophanes seems to confirm this result. [10]

$\alpha_1 = 26.8°$

$\alpha_2 = 12°$

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a = 1.391 Å
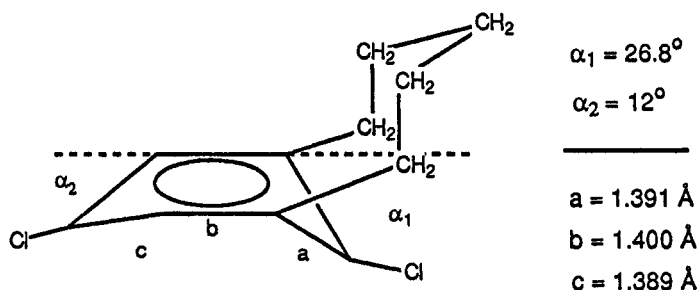
b = 1.400 Å

c = 1.389 Å

Fig. 10. Some X-ray crystal structure data of 10b.

DISCUSSION

When one compares the chemical reactivity of the bent benzene rings of small cyclophanes on the one hand and the physical data on the other, there appears to be a contradiction: reactivity seems to suggest that these compounds are polyolefins, while ring current effects, bond delocalization and other criteria (e.g. a thorough theoretical vibrational analysis by H.F. Schaefer, III, et al. [12]) unambiguously identify them as being fully aromatic.

In our opinion, this apparent dilemma can be clearly and simply resolved as follows. As more often in chemistry, reactivity is deceiving because it is not necessarily a reliable criterium for ground state characteristics. The high reactivity of the compounds discussed here is dominated by their high strain which is release in the transition state of the initial reaction step. All these reactions involve a rate determining conversion of one or two (aromatic) trigonal carbon atoms into (aliphatic) tetragonal ones; this allows a considerably less strained attachment of the bridge (Bredt's rule). Other factors may also contribute to the increased reactivity, for instance the unsymmetrical shape of the π-cloud on the two sides of the aromatic ring (see Fig. 11), and rehybridization of the aromatic carbon atoms.[23]

If then one accepts the message from physical measurements that the bent benzene rings remain aromatic, the next question is: why is this so inspite of the fact that bending must be detrimental to the π-system ? In the first place, simple qualitative reasoning (confirmed by MNDO calculations[11a,24]) tells us that while π-overlap decreases on the convex (outer) side of the bent ring, it improves on the inner side (Fig. 11); so the damage may be much smaller than one might intuitively anticipate.[11b]
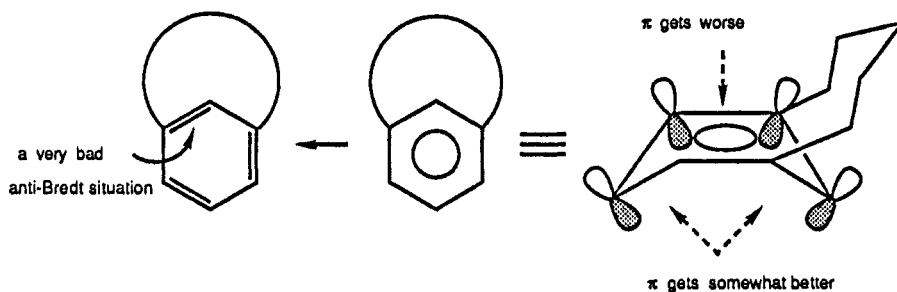

Fig. 11. Consequences of ring bending on the aromatic π-cloud.

Furthermore, if the aromatic ring would respond to bending by bond fixation, the strain situation would not improve, because two bridgehead double bonds are retained as illustrated in Fig. 11. This qualitative argument is supported by a quantitative MNDO analysis as first performed by Baird and Dewar for planar benzene itself.[25] One determines the heat of formation ΔH^{o}_{f} of a delocalized benzene ring and that of an artificially frozen cyclohexatriene ring with alternating bonds (C=C : 1.337 Å, C-C : 1.483 Å[25]). The difference $\Delta\Delta H^{o}_{f}$ is a measure for the additional delocalization energy in benzene; it is smaller than the total resonance energy of benzene, because the cyclohexatriene shows considerable overlap between the π-bonds. We have extended this approach to benzene rings (C_6H_6) which were distorted to the bent geometry of [5]- and [4]metacyclophane and of [4]paracyclophane (Table 2).[20b,26]

Table 2. MNDO heats of formation ΔH^{o}_{f} (C_6H_6)

geometry imposed	ΔH^{o}_{f} [kcal·mol⁻¹]		$\Delta\Delta H^{o}_{f}$
(flat)	21.2 (D_{6h})	30.8 (D_{3h})	9.6
	54.5	63.3	8.8
	79.1	89.2	10.1
	93.0	103.6	10.6

It turns out that the energy difference between the delocalized and the localized geometry is independent of the kind and degree of bending. This means that the strain in a localized and delocalized cyclophane is about equal; not gaining anything in terms of strain, but losing about 10 kcal/mol of additional delocalization energy, the bent benzene ring has no incentive to become a cyclohexatriene.

CONCLUSION

A provocative contemplation may be appropriate at the end of this paper. Although the experimental results concerning the anisotropy of the magnetic susceptibility need to be confirmed with higher accuracy, they furnish some evidence that a bent benzene ring may be more aromatic than a planar one (cf. the arguments for aromaticity).[21] This strange conclusion would appear to be reasonable if indeed, as suggested by theoretical calculations,[3] it is the σ-system which is responsible for delocalization. After all, the σ-system may be expected to suffer less from bending than the π-system does; (some) weakening of the latter must therefore be favourable for delocalization.

REFERENCES

1. K. Tamagawa, T. Iijima, M. Kimura, *J. Mol. Struct. 30*, 243 (1976).
2. P.J. Garrat, *Aromaticity*, Wiley-Interscience, New York (1986).
3. a) R.S. Berry, *J. Phys. Chem. 35*, 2253 (1961). b) N.D. Epiotis, *Nouv. J. Chim. 8*, 11 (1984). c) S.S. Shaik, P.C. Hiberty, G. Onessian and J.M. Lefour, *J. Phys. Chem. 92*, 5086 (1988), and references cited therein. d) A. Stanger and K.P.C. Vollhardt, *J. Org. Chem. 53*, 4889 (1988).
4. N.C. Baird, *J. Org. Chem. 51*, 3907 (1986).
5. P.A. Schultz and R.P. Messmer, *Phys. Rev. Lett. 58*, 2416 (1987).
6. O. Ermer, *Angew. Chem. 99*, 791 (1987), and references cited therein; see also ref. 7.
7. R. Janoschek, *Angew. Chem. 99*, 1332 (1987).
8. W.R. Angus, C.R. Bailey, J.B. Hale, C.K. Ingold, A.H. Leckie, C.G. Raisin, J.W. Thompson and C.L. Wilson, *J. Chem. Soc.* 966, 971 (1936).
9. Cited in : *Bild der Wissenschaft, 10*, 34 (1988).
10. F. Bickelhaupt and W.H. de Wolf, *Recl. Trav. Chim. Pays-Bas 107*, 459 (1988), and references cited therein.
11. a) L.W. Jenneskens, Thesis, Vrije Universiteit, Amsterdam, 1986. b) L.W. Jenneskens, to be published.
12. a) J.E. Rice, T.J. Lee, R.B. Remington, W.D. Allen, D.A. Clabo, Jr., and H.F. Schaefer, III, *J. Am. Chem. Soc. 109*, 2902 (1987). b) T.J. Lee, J.E. Rice, R.B. Remington and H.F. Schaefer, III, *Chem. Phys. Lett. 150*, 63 (1988). c) T.J. Lee, J.E. Rice, W.D. Allen, R.B. Remington and H.F. Schaefer, III, *Chem. Phys. 123*, 1 (1988).
13. P.M. Keehn and S.M. Rosenfeld, Eds., *Cyclophanes*, Academic Press, New. York, 1983. Vol. I and II, and references cited therein.
14. S.L. Kammula, L.D. Iroff, M. Jones, Jr., J.W. van Straten, W.H. de Wolf and F. Bickelhaupt, *J. Am. Chem. Soc. 99*, 5815 (1977).
15. L.W. Jenneskens, F.J.J. de Kanter, P.A. Kraakman, L.A.M. Turkenburg, W.E. Koolhaas, W.H. de Wolf, F. Bickelhaupt, Y. Tobe, K. Kakiuchi and Y. Odaira, *J. Am. Chem. Soc. 107*, 3716 (1985).
16. a) T. Tsuji and S. Nishida, *J. Chem. Soc., Chem. Commun.*, 1189 (1987). b) T. Tsuji and S. Nishida, *J. Am. Chem. Soc. 110*, 2157 (1988).
17. a) G.B.M. Kostermans, M. Bobeldijk, W.H. de Wolf and F. Bickelhaupt, *J. Am. Chem. Soc. 109*, 2471 (1987). b) G.B.M. Kostermans, P.J. Kwakman, P.J.W. Pouwels, G. Somsen, W.H. de Wolf and F. Bickelhaupt, *J. Phys. Org. Chem. 2*, 331 (1989).
18. E.L. Mackor, A. Hofstra, and J.H. van der Waals, *Trans. Far. Soc. 53*, 186 (1957).
19. J.W. van Straten, W.H. de Wolf and F. Bickelhaupt, *Tetrahedron Lett.*, 4667 (1977).
20. a) G.B.M. Kostermans, P. van Dansik, W.H. de Wolf and F. Bickelhaupt, *J. Am. Chem. Soc. 109*, 7887 (1987). b) G.B.M. Kostermans, Thesis, Vrije Universiteit, Amsterdam, 1989.
21. P.C.M. van Zijl, L.W. Jenneskens, E.W. Bastiaan, C. MacLean, W.H. de Wolf and F. Bickelhaupt, *J. Am. Chem. Soc. 108*, 1415 (1986).
22. L.W. Jenneskens, J.C. Klamer, H.J.R. de Boer, W.H. de Wolf, F. Bickelhaupt amd C.H. Stam, *Angew. Chem. 96*, 236 (1984).
23. R.C. Haddon, *J. Am. Chem. Soc. 109*, 1676 (1987).
24. L.A.M. Turkenburg, Thesis, Vrije Universiteit, Amsterdam, 1982.
25. N.C. Baird and M.J.S. Dewar, *J. Am. Chem. Soc. 91*, 352 (1969).
26. L.W. Jenneskens, F.J.J. de Kanter, W.H. de Wolf and F. Bickelhaupt, *J. Comput. Chem. 8*, 1154 (1987).