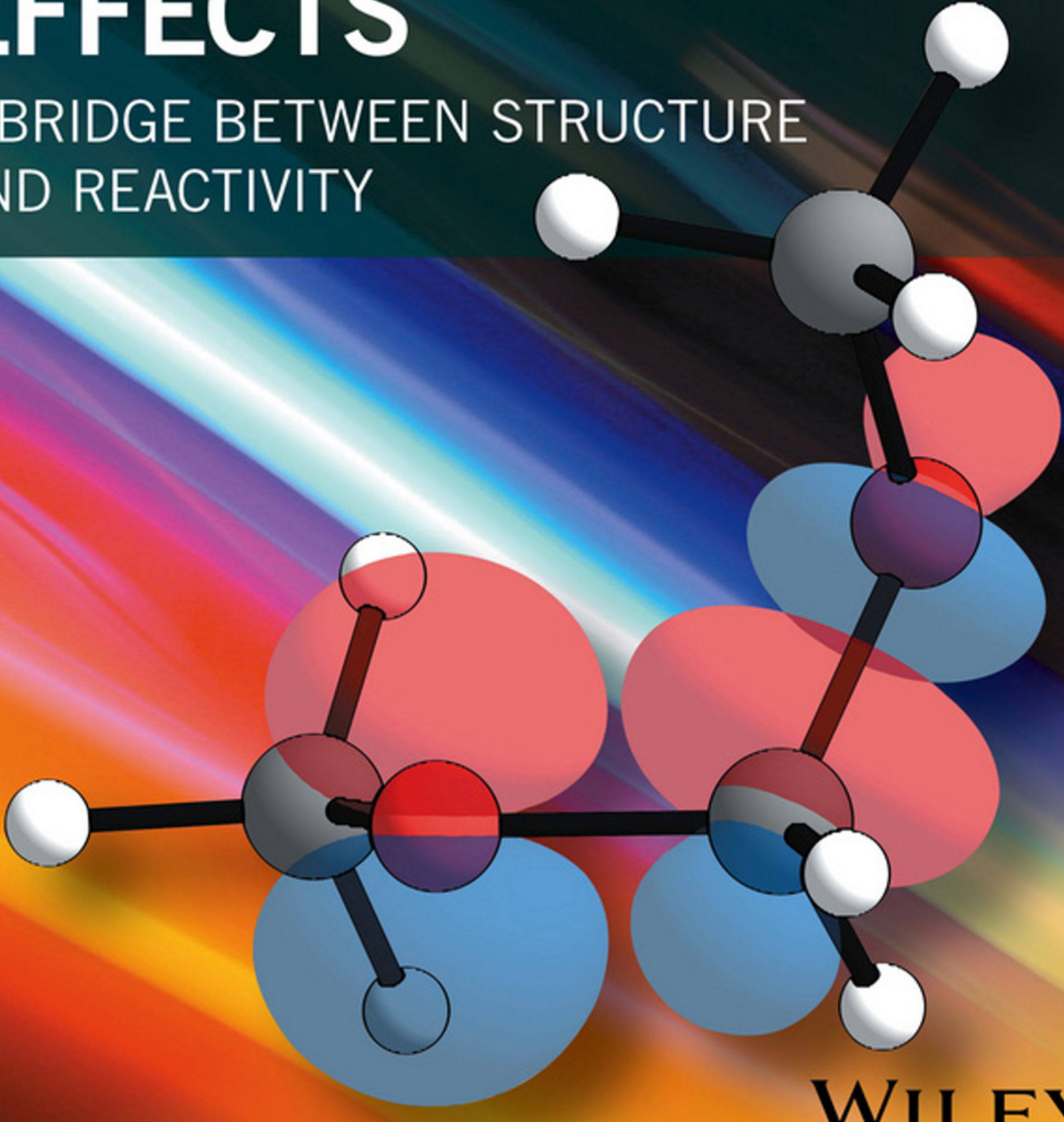
IGOR V. ALABUGIN

STEREOELECTRONIC EFFECTS

A BRIDGE BETWEEN STRUCTURE AND REACTIVITY



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A Bridge Between Structure and Reactivity

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Contents

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This book reviews and summarizes the hard work of several generations of chemists who uncovered hidden controlling factors bringing order to the seemingly bewildering diversity of chemical reactivity. My mentors, Nikolai Zefirov, Howard Zimmerman, and Frank Weinhold, stoked my early interest in this topic and gave me the tools necessary for understanding the role of orbital interactions in chemistry. Discussions with Wes Borden, Eusebio Juaristi, Joseph Lambert, Hans Reich, and Peter Schreiner provided valuable insights into the broader implications of stereoelectronic concepts. I also appreciate the feedback and comments from my colleagues at FSU: Greg Dudley, Jim Frederich, and Jack Saltiel.

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Supplementary Material

Instructors can access PowerPoint files of the illustrations presented within this text, for teaching, at: http://booksupport.wiley.com.

1

Introduction

> When people thought the earth was flat, they were wrong. When people thought the earth was spherical, they were wrong. But if you think that thinking the earth is spherical is just as wrong as thinking the earth is flat, then your view is wronger than both of them put together. I. Asimov

1.1 Stereoelectronic effects – orbital interactions in control of structure and reactivity

It is easy to believe that the Earth is flat when driving through the Great Plains. Furthermore, the “flat Earth” approximation works quite well in many other aspects of everyday life. Because the small deviation from planarity – only 8 inches per mile – does not make a difference for everyday activities, we can order a cup of coffee or play a game of golf without worrying about the fine details of planetary shapes. However, once one prepares to launch a satellite instead of a golf ball or to navigate “around the globe”, the planet’s curvature becomes crucial. But ***is*** Earth a globe? A closer look from space finds that Earth is not a sphere but an “oblate spheroid” that bulges at the equator. Another revision! When should refinements stop and why should a chemist care?

The story of the flat Earth, borrowed from Isaac Asimov,[1] reflects the common evolution of scientific models. Sometimes, models are discarded completely (e.g. phlogiston) but, more often, they are refined and taken to the next level of applicability (such as Newton’s theory of gravity paving the way for Einstein’s theory of relativity). How does it apply to organic chemistry? How adequate are the undergraduate organic foundations for the broad understanding of structure and reactivity? Do we really need to go deeper?

The importance of continuous improvement of models is illustrated by the following “diagnostic quiz” given to first-year graduate students at the Florida State University. Take a minute and test yourself.

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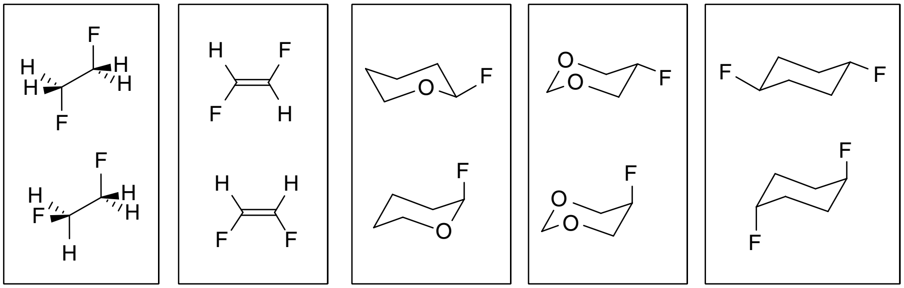

Figure 1.1 *Circle the more stable structure in each of the above pairs.*

The answers may or may not be surprising, depending on how far the reader is separated from the undergraduate organic class. For each pair in Figure 1.1, the *bottom structure is more stable* than the top structure. In particular, the gauche conformation of 1,2-difluoroethane is more stable than the anti conformations; *cis*-difluoroethene is more stable than the *trans*-isomer; the equatorial conformers of the two fluoro-substituted oxacyclohexanes are less stable than their axial counterparts; and the *di*axial 1,4-difluorocyclohexane is ~1 kcal/mol *more* stable than the diequatorial conformer. The answer in each case is opposite to expectations based on the steric repulsion – the "flat Earth" models that have served reasonably well as a foundation of undergraduate organic chemistry.

It is not surprising that it is a rare undergraduate student who gives correct answers to all of the above problems. Almost invariably, the correct answers come as a surprise, even to a student with a good mastery of undergraduate organic chemistry. Clearly, a new set of concepts is needed to refine the initial model of organic structure and reactivity. This book aims to introduce these concepts in a way that will provide a logical ascension from a simplified discussion of an undergraduate textbook to a level appropriate for a practicing organic chemist.

Undergraduate organic chemistry lays the foundation of chemical knowledge – a reasonable approximation and a useful and often sufficient way to describe molecules as Lewis structures augmented, as needed, by resonance. However, once one realizes that organic molecules are quantum objects delocalized in space, far from the flat two-dimensional drawings on a sheet of paper or a blackboard, it may not be a complete surprise that this simple concept has its limitations.

The way to get to the next step in understanding molecular structure is to move from the flat Lewis structures on a flat sheet of paper to the 3rd dimension. The elements of stereochemistry are introduced, of course, in undergraduate courses. However, this important step is not enough – when one needs to design, understand, and control new reactions, it is crucial to start thinking about organic molecules as intrinsically delocalized and spatially anisotropic quantum objects. This book focuses on the importance of delocalization – the deviation of real molecules, quantum objects par excellence, from idealized Lewis structures.

The laws of chemical attraction in the world of atoms and molecules are defined by quantum mechanics. Constructive interference of electronic wavefunctions is the quantum essence of chemical bonding that "glues" smaller fragments into larger molecular assemblies. As a result, the chemical world at the molecular level is defined by interactions between atomic and molecular orbitals. Because orbitals and molecules are three-dimensional, such interactions depend on the relative atomic arrangements in space. The modulations of electronic interactions by changes in molecular geometry are generally referred to as *stereoelectronic effects*. In organic chemistry, stereoelectronic effects can be defined as *stabilizing electronic interactions maximized by a particular geometric arrangement which can be traced to a favorable orbital overlap*. Stereoelectronic interactions are omnipresent in chemistry, as only a small subgroup of electronic effects, i.e. the long-range[2] electrostatic effects, can be considered, with a degree of approximation, as not having a substantial stereoelectronic component.

There is one common misunderstanding that needs to be addressed early: “stereoelectronic” is not the same as “steric + electronic”! By definition, stereoelectronic effects are *always* stabilizing, reflecting increased delocalization at favorable conformations. Repulsive steric interactions also depend on the arrangement of orbitals in space but, historically, are not included under the umbrella of stereoelectronic effects.

Stereoelectronic factors control interactions between different atoms or molecules and interactions between different parts of a single molecule. Although our focus will be on the latter, we will also briefly illustrate the fundamentals of intermolecular interactions, because they broaden the conceptual foundation for subsequent discussion and illustrate the key patterns for orbital overlap without intramolecular constraints being imposed on the geometries.

Understanding the role of orbital interactions can be beneficial from the practical perspective. For example, the symmetry of frontier molecular orbitals can explain why thermal [2 + 2] cycloaddition fails, whereas the analogous reaction of transition metal alkylidenes, compounds that can be described as having a metal–carbon double bond, proceeds efficiently under mild conditions (Figure 1.2). In this case, *an extra orbital node is the difference between a failed reaction and a Nobel Prize!*

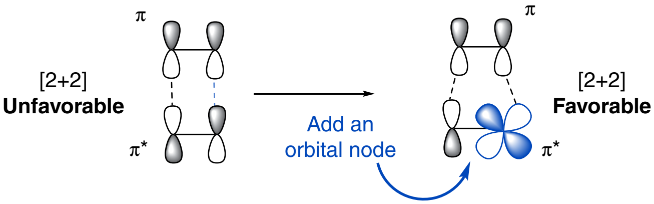

Figure 1.2 *The striking effect of orbital symmetry on [2 + 2] cycloadditions.*

1.2 Orbital interactions in theoretical chemistry

The concept of stereoelectronic effects resulted from the cross-pollination of quantum-mechanical ideas (both valence bond, VB and molecular orbital, MO) with the three-dimensional thinking of organic chemists. The involvement of orbitals evolved over the 20th century from the qualitative ideas of Lewis and Pauling through the approximations of Hückel and semi-empirical treatments to the sophisticated accuracy of modern multiconfigurational approaches. However, even the most complex wavefunctions can still be analyzed in terms of individual orbitals using such methods as natural bond orbital (NBO) analysis (introduced in Chapter 4). Such dissection allows one to recover the basic Lewis concepts that seem to be lost in the mathematical jungle and to use them as a foundation for developing the deeper understanding of electronic structure.

In parallel, experimental organic chemistry grew in scope and sophistication. A large body of information was acquired allowing precise measurements of molecular geometries, spectroscopic parameters, and reaction kinetics to provide the necessary basis for the fruitful application of stereoelectronic ideas on a quantitative basis.

The accuracy of computational methods has started to rival experimental measurements, but finding the optimal compromise between computational accuracy and cost is an ever-moving target. Time-resolved experimental techniques allow understanding reactivity on the fly, accessing increasingly exotic and increasingly unstable species with even transition states[3] and, more recently, hilltops on potential energy surfaces[4] succumbing to experimental scrutiny. This is a productive interplay. Experiments are important for benchmarking and testing theory,[5] whereas theory is useful in guiding and streamlining experiments.

1.3 The birth of stereoelectronic concepts in organic chemistry

Initially, even the simple 3D description of molecules was a controversial idea. In fact, Van't Hoff's 1874 book *La chimie dans l'espace* was ridiculed by such eminent chemists as Adolph Kolbe, the editor of the Journal für Praktische Chemie, who stated:

> A Dr. H. van't Hoff of the Veterinary School at Utrecht has no liking, apparently, for exact chemical investigation. He has considered it more comfortable to mount Pegasus (apparently borrowed from the Veterinary School) and to proclaim in his "La chimie dans l'espace" how the atoms appear to him to be arranged in space, when he is on the chemical Mt. Parnassus which he has reached by bold flight.[6]

However, the situation had already changed drastically before the early 1950s when important stereochemical concepts had already permeated the fabric of organic chemistry. In 1954, the term "stereoelectronic" was born in a paper by Hirschmann et al.[7] who disclosed a remarkable coordinated ring contraction/expansion in rockogenin (Figure 1.3).[8] The authors stated that "the stereoelectronic requirements are fulfilled only in the case of the natural C_{12}-β-configuration. The significance of this geometrical factor is reflected in the extraordinary ease with which this rearrangement occurs." The unprecedented rearrangement to a new ring system took place instead of the more mundane methyl migration or elimination without rearrangement.

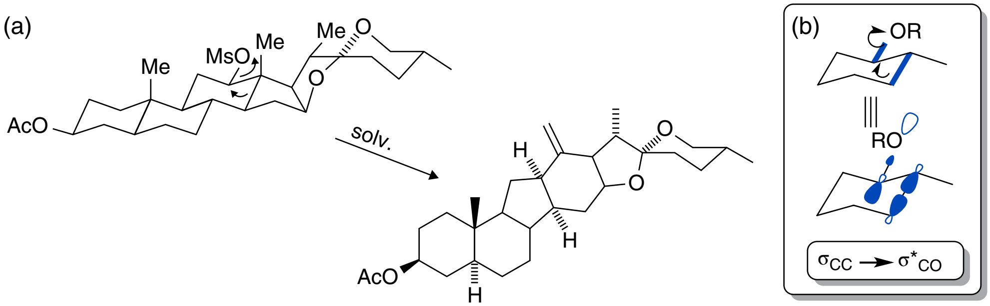

Figure 1.3 *(a) Rearrangement of rockogenin as reported by Hirshmann (Source: Hirschmann 1954 (7). Reproduced with permission of American Chemical Society). (b) Orbital interactions involved in the bond reorganization.*

Two years later, in 1956, E. J. Corey, a young professor at the University of Illinois used "stereoelectronic" in the title of a paper ("Stereoelectronic Control in Enolization-Ketonization Reactions").[9] In this paper, he associated the faster loss of axial hydrogen in enolization and the faster gain of axial hydrogens in ketonization with the more favorable orbital overlap of the carbonyl π-system with the axial C-H bonds relative to the equatorial C-H bonds (Figure 1.4).

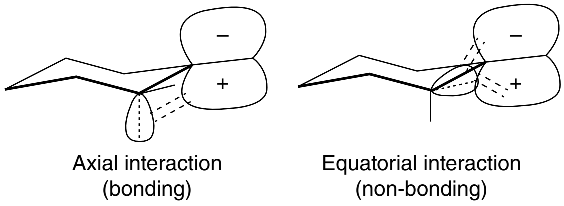

Figure 1.4 *Early comparison of the carbonyl π-system overlap with the axial and equatorial C-H bonds. (Source: Corey 1956 (9). Reproduced with permission of American Chemical Society.)*

The evolution of stereoelectronic concepts was further catalyzed by steroid synthesis and rapid development of conformational analysis recognized by the 1969 Nobel Prize to Barton and Hassel. However, it was not until 1983, that an organized treatise dedicated to stereoelectronics was published (the important books by Deslongchamps and Kirby).[10]

What does the future hold, or "Are we living on an oblate spheroid"? To take the Earth analogy even further, one can illustrate that the basic stereoelectronic concepts are likely to have their own limitations as well. Further refinements of our understanding of chemical structure are unavoidable. For example, stereoelectronic concepts discussed in the following sections are still just an approximation of the exuberant variety of bonding patterns created by the chemical cornucopia known as the periodic table. There are systems so delocalized that starting with a Lewis structure is simply too far off for arriving to a useful description. For such highly delocalized structures, the Lewis approximation is just too crude, and the perturbative approach, which we refer to as resonance, is not able to correct this deficiency. In such cases, it is more productive to describe a molecular system from an MO perspective. Striving to delocalization, transition states and unstable reactive intermediates defy the limitations imposed by the classic two-center two-electron bond: the Lewis structure's line between atoms. Odd-electron systems are incapable of perfect electron-pairing by their nature. Aromatic and antiaromatic molecules, inorganic clusters, and multicentered bonding in reactive intermediates are examples that further emphasize the primary importance of electronic delocalization.

Quantum tunneling Furthermore, the assumption that nuclear motion is slow enough to be separated from the motion of electrons (the Born–Oppenheimer approximation) and the expectation, that one can always assign distinct connectivity to a molecule, are also only approximations. In the world of quantum phenomena, the whole system of electrons and nuclei can take advantage of Heisenberg's uncertainty principle and "miraculously" morph into a different molecule with different connectivity even under conditions approaching absolute zero, as long as the barrier separating the two molecular structures is relatively narrow ("quantum tunneling")[11] – Figure 1.5.

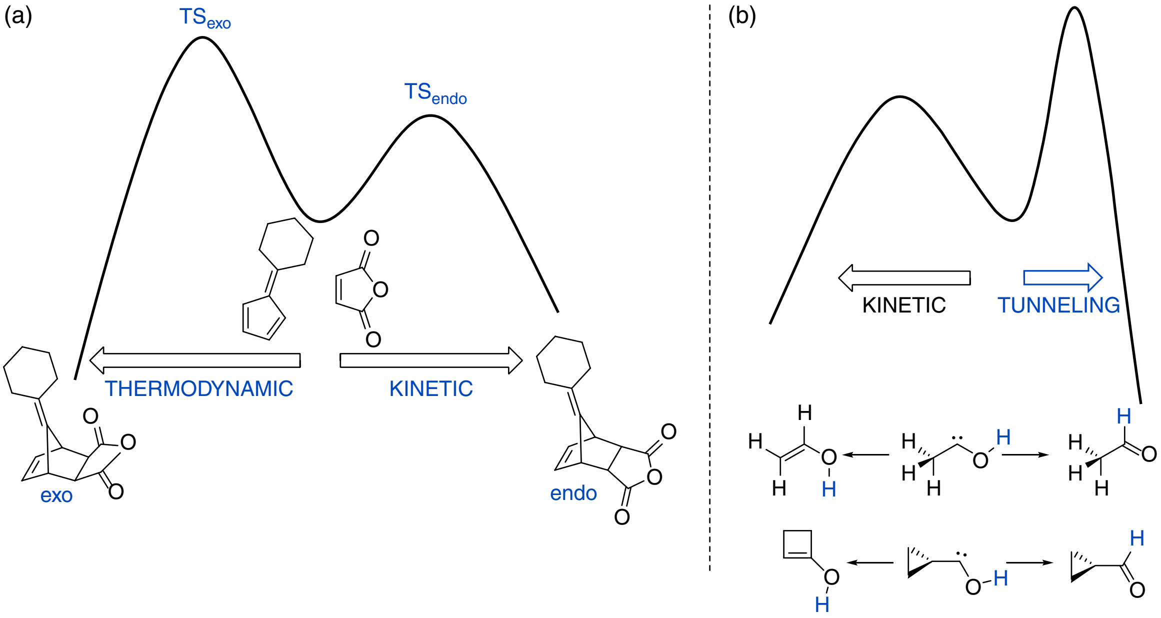

Figure 1.5 *Three regimes of reaction control. (a) kinetic vs. thermodynamic control*[12] *(b) kinetic vs. tunneling control.*[11]

Molecular trajectories Further conceptual limitations of our understanding of chemical reactivity are illustrated by the simple notion that even the exact knowledge of energies and structures of every stationary point at the potential energy surface for a chemical system is not sufficient for accurately predicting the distribution of products for a given set of starting materials. One has to know the shape of the TS region in the $3N-6$ dimensional space and the forces that affect a N-atom molecular system that traverses this region on its route from reactants to products.[13]

"Shapeshifting molecules" Not just the position of atoms but also molecular connectivity can be dynamic in the most unusual ways. In so-called fluxional molecules, the whole concept of a single Lewis structure fails at a different level. In these systems, *nuclear* structural reorganization and bond breaking/bond reforming are fast on the chemical timescale.[14] For example, the 10 carbon atoms of bullvalene have identical bonding environment at 140 °C. Both the proton and the carbon NMR spectra show single signals (at 4.2 and 86.4 ppm, respectively), indicating that every carbon atom experiences the identical surroundings and that 10!/3 or 1,209,600 contributing Lewis structures interconvert in this unique "molecule". There are no permanent C-C bonds in bullvalene, but every atom is equally connected to any other atom! As stated by Doering: "all ten carbon atoms [must] inevitably wander over the surface of a sphere in ever changing relationship to each other".[15] In the presence of several substituents, each bullvalene molecule becomes a "dynamic library" of compounds[16] – Figure 1.6.

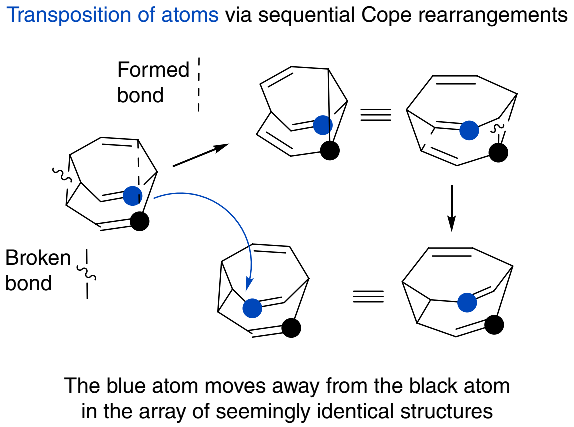

Figure 1.6 *Part of the extended reaction network connecting multiple isomers of bullvalene via degenerate Cope rearrangements. Although the structure seems to remain unchanged, note that the blue carbon atom moves away from the black atom.*

The future of chemistry is full of surprises and, as the boundary with the unknown parts of the chemical universe continues to expand, we need to refine our models as we move deeper into the rich world of fuzzy objects at the subnanoscale.

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2

Direct Effects of Orbital Overlap on Reactivity

Stabilizing orbital interactions come in a variety of patterns. For example, in intramolecular scenarios, they can either involve formation of covalent bonds from two non-bonding orbitals (e.g. two p-orbitals in a π-bond, or a lone pair and an empty p-orbital in oxycarbenium ions, heteroatom-substituted singlet carbenes etc.), or be responsible for a plethora of "second order interactions". The latter include interactions between π-bonds (conjugation), between non-bonding orbitals and σ-bonds (classic negative or positive hyperconjugation), or between two σ-bonds (σ-conjugation). The intermolecular scenarios can involve supramolecular contacts with n→σ* or n→π* components (Figure 2.1). The list of such interactions rapidly expands from the familiar hydrogen bonding to halogen, pnictogen, chalcogen and tetrel bonding (vide infra).

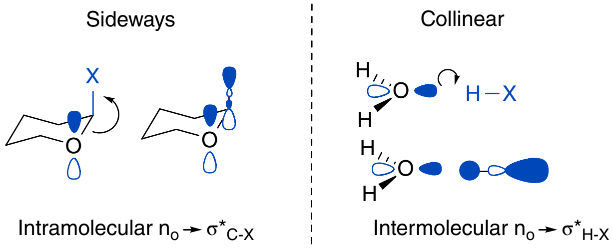

Figure 2.1 *Comparison of intramolecular and intermolecular overlap patterns for interaction between lone pairs and antibonding orbitals.*

We will start with the simplest case – interaction of two non-bonding orbitals with an overall population of two electrons. This case corresponds to the classic formation of a two-center/two-electron (2c,2e) chemical bond. However, even this familiar situation allows for a number of interesting modifications. For example,

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even within a narrow class of bonds, e.g. C-C and C=C bonds, remarkable variations in the apparent bond strength can be found (Figure 2.2).[1]

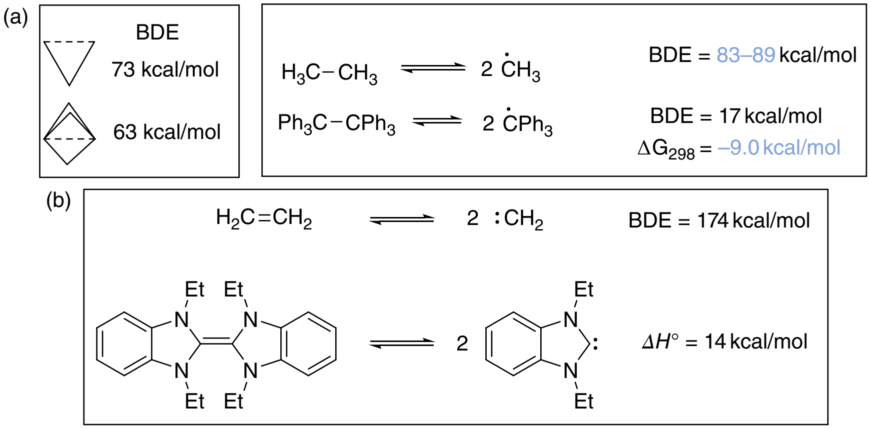

Figure 2.2 *Variations in apparent bond strength as evaluated by selected bond dissociation energies (BDEs) and enthalpies. (a) C-C bonds, (b) C=C bonds.*

2.1 Bond formation without bond breaking: types of overlap in two-orbital interactions

In the language of molecular orbital theory, the 2c,2e chemical bond is described via the formation of two new orbitals: the low energy filled bonding orbital and the high energy empty antibonding orbital, (Figure 2.3). In such systems, bond formation is not complicated by simultaneous bond breaking. Furthermore, one does need to consider the effect of four-electron repulsion. Nevertheless, this process is still controlled by stereoelectronic effects, and many interesting variations are possible.

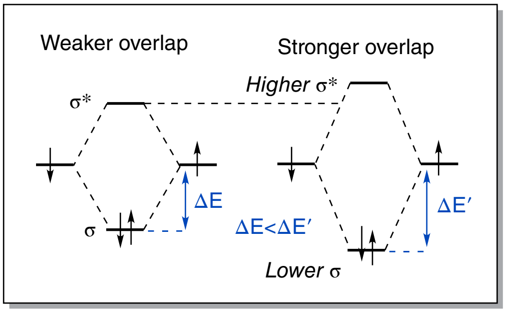

Figure 2.3 *Formation of two-center two-electron chemical bonds and role of overlap in bond strength. The stronger overlap also leads to lower energy bonding orbitals and higher energy antibonding orbitals.*

Interactions of s-, p-, and d-orbitals are usually classified within the three main types of orbital overlap: σ, π, and δ (Figure 2.4).

For the intermolecular formation of a single bond between two interacting fragments, the direct approach where interacting orbitals overlap along the line connecting the two atomic centers is preferred, leading to the textbook description of σ-bond formation.

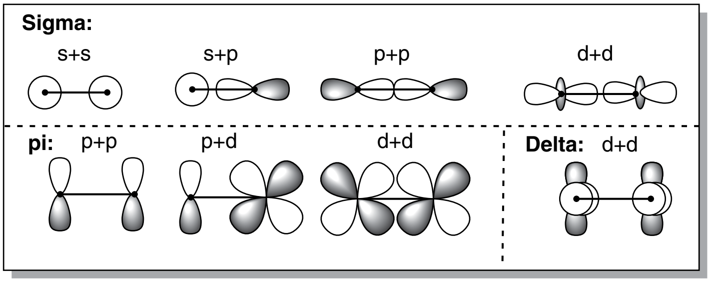

Figure 2.4 Making bonds out of atomic orbitals: σ, π, and δ-overlaps of s-, p-, and d-orbitals.

The second type of overlap, the π-type, is characteristic for molecules that already possess a σ-bond. In this approach, the two orbitals are parallel rather than collinear. Not only does this overlap pattern describe such important functional groups as alkenes, alkynes, aromatics, and carbonyl derivatives, but the π-type overlap often plays a key stereoelectronic role even in molecules *without* a double bond. For example, the π-overlap is important in vicinal hyperconjugative interactions (Figure 2.5), providing a stereoelectronic basis to such phenomena as the anomeric effect, gauche effect, and *cis*-effect (vide infra).

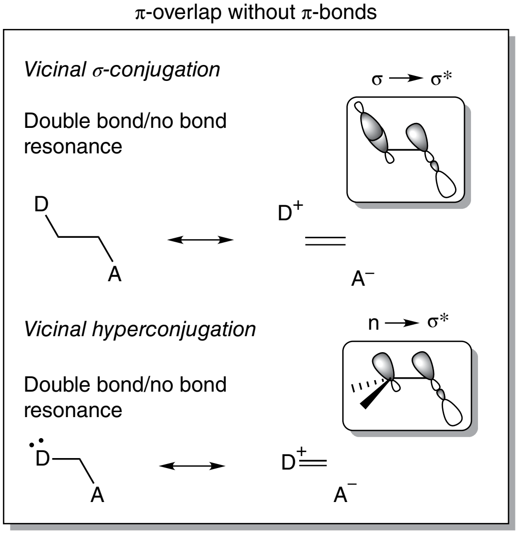

Figure 2.5 Examples of interactions using π-overlap in systems lacking formal double bonds in the main Lewis structure.

Finally, metal–metal interactions may include δ-bonding, where *four* lobes of one atomic orbital overlap with four lobes of the other atomic orbital (Figure 2.4). The δ-bonds have two nodal planes which intersect at the internuclear axis (for the first compound with a δ-bond and for the first example of d-aromaticity, see references 2 and 3, respectively).

The σ,π,δ-overlaps can combine to make compounds with bond orders exceeding those available to organic molecules (e.g. 1 σ, 2 π, and 2 δ orbitals for a quintuple bond order) and, in combination with structural constraints, can lead to very short M-M bonds (Figure 2.6).[4] The 1.706 Å metal–metal distance observed in a quintuply bonded Cr-Cr bimetallic complex is the same as the longest C-C bond in stable alkanes.[5]

The efficiency of the overlap is generally reflected in the strength of the chemical bond formed by this overlap: σ>π>δ. As a result, it is common to have σ-bonds without π and δ bonds in a molecule, but

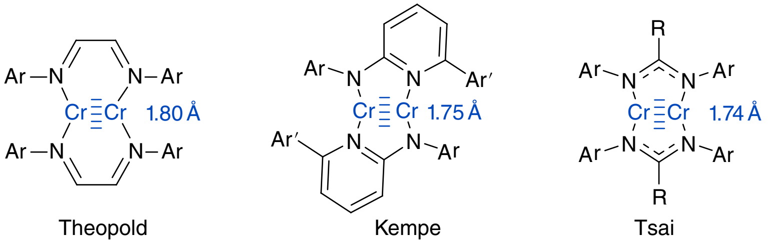

Figure 2.6 *Selected molecules with very short metal–metal bonds.*

a bonding situation where, for example, a π-bond is formed without a single bond is uncommon; although, curiously, it is not impossible. For example, C2 can be considered as a molecule held together by two "levitating π-bonds" without a single bond.[6] Furthermore, four and even five atoms in the Mg_3^-, $NaMg_3^-$, and Na_2Mg_3 species, respectively, were suggested to be held together by only a single π-bond without involving σ-bonds.[7]

A useful, but relatively rare, alternative description for the systems with both σ- and π-bonds between two atoms is the bent bond model. In this model, the double bond is described as a combination of two equivalent "banana bonds" formed from sp^5 hybrid orbitals (Figure 2.7a). Such orbitals correspond to the linear combinations of the classic sp^2 and p-orbitals of the σ,π-description.[8] The two descriptions are complementary because the linear combinations of two orbitals correspond to the same overall electron density.[9] We will show in a later chapter that there are cases when such an unconventional description of alkenes can be helpful in understanding conformational effects. In a few cases, when a double bond is connected to a σ-acceptor group that draws additional p-character from the central atom to satisfy Bent's rule (the classic correlation between hybridization and electronegativity introduced by H. Bent),[10] there is not enough p-character left for formation of normal π-bond and the banana bond description becomes the *only* choice for making a double bond.[11] Furthermore, the dichotomy between σ/π vs. "mixed hybrids" descriptions of a pair of orbitals at a given atom also displays itself in systems with two non-bonding orbitals, e.g. CH_2 (singlet carbene) and H_2O. The two systems have the same set of molecular orbitals (MOs), albeit populated with a different number of electrons. In both cases, two of the MOs can be considered non-bonding. It is curious that whereas the non-bonding MOs (NBMOs) of carbene are generally considered different and assigned as σ (for the occupied MO) and π (for the empty MO), the choice between the two different descriptions for the lone pairs of water is often made in a seemingly sporadic fashion. In a physical chemistry textbook, the lone pairs can be different (σ and π) and look very similar to the non-bonding MOs of carbene. On the other hand, an organic chemistry

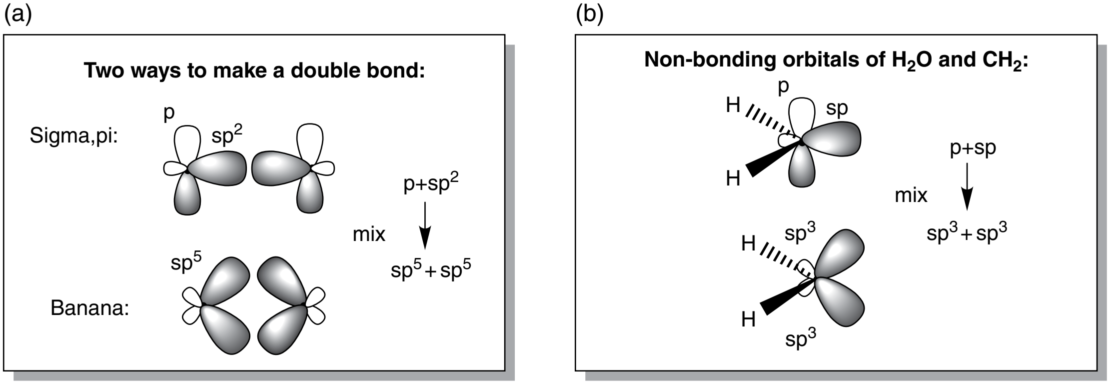

Figure 2.7 *Two descriptions of double bonds and two descriptions of lone pairs in water.*

paper will often utilize equivalent sp^3 hybridized “rabbit ears” (Figure 2.7b). An excellent discussion of this and other “orbital anachronisms” can be found in a recent educational review of Weinhold and coworkers.[12] We will provide a detailed discussion of lone pairs of oxygen and other heteroatoms in Chapter 5.

Another example of “bent” bonds is provided by the chemical bonding in heavier analogues of carbon, where hybridization is hampered by the cost of electron promotion. The double bonds in distannane may be regarded as two banana donor–acceptor (dative) bonds as opposed to the common description of double bond model of one σ-bond and one π-bond. This description explains why, instead of the “usual” alkene-like geometry, these species are “trans-bent” with a weak Sn=Sn double bond.[13] The heavier triple bond analogues, such as disilyne,[14] also have the “trans-bent” structure. In the latter case, bonding involves two donor–acceptor (dative) banana bonds augmented by one π-bond (Figure 2.8).

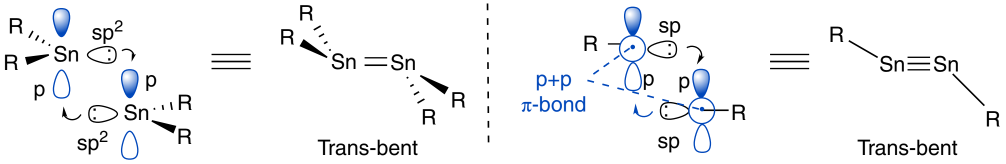

Figure 2.8 Banana double and triple bonds in heavier elements.

2.1.1 Factors controlling σ-bond overlap

Hybridization As we saw above, unusual hybridizations can lead to unusual bonding patterns and geometries. Such effects are not limited to “exotic” species made out of heavier atoms. Carbon also has its surprises.

It is well-known that σ-overlap of two p-orbitals or two s-orbitals does not take full advantage of the available orbital density. In order to maximize σ-overlap, the interacting atoms change their orbital shapes in a non-symmetric way (rehybridize). Because hybridization is associated with changes in orbital overlap, it can be considered as one of the most basic stereoelectronic effects that can impose significant modulations on other stereoelectronic interactions.

Even for a σ-bond between the same pair of atoms, hybridization strongly affects the bond strength as illustrated by the differences in BDEs for sp(C-H) > sp^2(C-H) and sp^3(C-H) (Figure 2.9). Both the greater overlap and the increased polarity contribute to this BDE increase. In bonds with increased s-character, carbon behaves as a more electronegative element. From the sp-hybridized carbon point of view, the homolytic C-H bond cleavage is an oxidation reaction that goes against the natural C-H bond polarization in this system! On the other hand, deprotonation at an sp-hybridized carbon is, of course, more favorable in comparison to the C-H bonds with lower s-characters since it takes advantage of the increased electronegativity of sp-hybridized carbon. Such textbook observations reflect the strong correlation between hybridization and electronegativity.

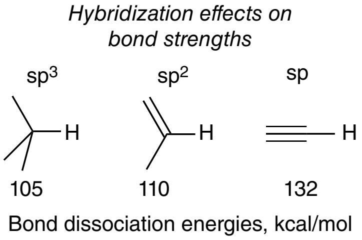

Figure 2.9 Hybridization effects on bond strengths in C-H bonds.

Hybridization is commonly applied to carbon-based chemistry since all σ-bonds formed by carbon atoms are hybridized.[15] However, this concept extends to a variety of other bonds across the periodic table, with electronegativity and orbital size effects leading to dramatic variations in hybridization efficiency for the different bond types.[16] On occasions, other elements can form sigma bonds with little or no help from hybridization (e.g. the orbitals forming the F-F bond in F_2 have >90% of p-character, corresponding to ~ sp^9 hybridization). In general, s/p mixing becomes progressively less important as the nuclear charge increases from left to right in the periodic table because the energy of s-electrons decreases faster than the energy of p-electrons (Figure 2.10). In the case of F_2 and similar cases with large s,p energy separation, the gain in overlap does not compensate for the cost of electron promotion (i.e. the involvement of the low energy s-electrons in chemical bonding). When mixing of s and p-orbitals becomes unfavorable, unusual reactivity is often observed.[17]

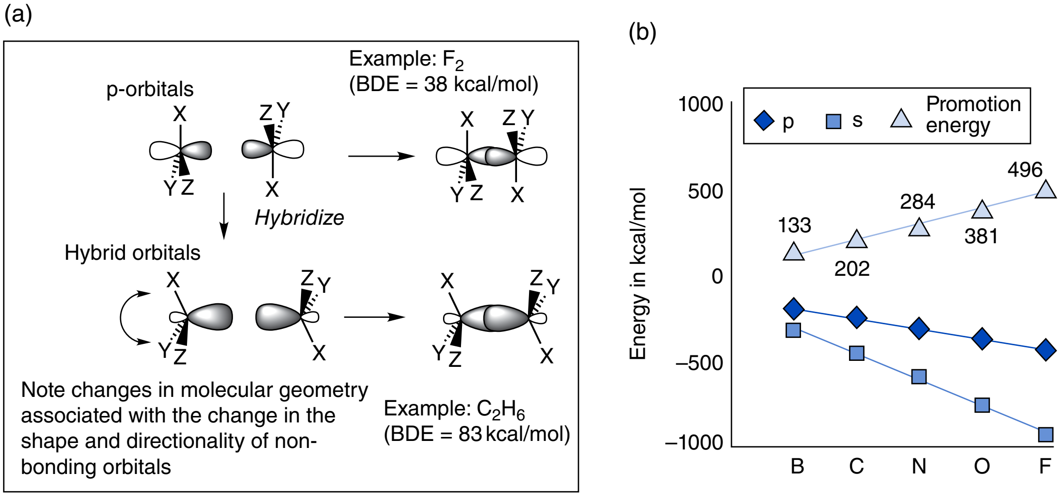

Figure 2.10 *Optimization of orbital overlap in bond formation provided by hybridization of atomic orbitals: energy of the p-orbitals (diamonds), s-orbitals (squares) and the promotion energy (triangles) for B, C, N, O, and F. As the promotion energy rises, the importance of hybridization is expected to decrease. Values from ref. 18.*

Although hybridization is more often used in VB theory, this concept is introduced naturally in MO theory via mixing of s and p-orbitals. Modern computational techniques (such as NBO analysis discussed in Chapter 4) can find the "optimal" hybridization for localized orbitals constituting a particular wavefunction, providing a convenient approach to quantifying hybridization trends. In addition to polarity, hybridization is related to bond strength and can be probed via isotope effects and spectroscopic methods. Furthermore, it manifests itself in numerous effects on structure and reactivity. An expanded analysis of such effects with the particular emphasis on a very useful correlation between hybridization and electronegativity (Bent's rule) can be found in the recent literature (ref. 10,11) and will not be repeated here.

Orbital size mismatch Orbital size differences play a role in determining the strength of bonds between different partners.[19] For example, the relatively strong C-H bond, one of the most stable structural units of organic chemistry, starts to weaken considerably as carbon is changed to its heavier cousins (Si, Ge, Sn, Pb in Figure 2.11).[20] In particular, the enormous utility of organostannanes ("lovingly" referred to as "the Tyranny of Tin" by radical chemists) for the initiation of radical transformations stems from the weakness of the Sn-H bond originating from the large difference in size between tin and hydrogen (Figure 2.11). This bond can be broken relatively easily with carbon-centered radicals, and the generated tin radicals can attack weaker carbon-halogen bonds, i.e. the C-Br and C-I bonds to form stronger Sn-Br and Sn-I bonds.

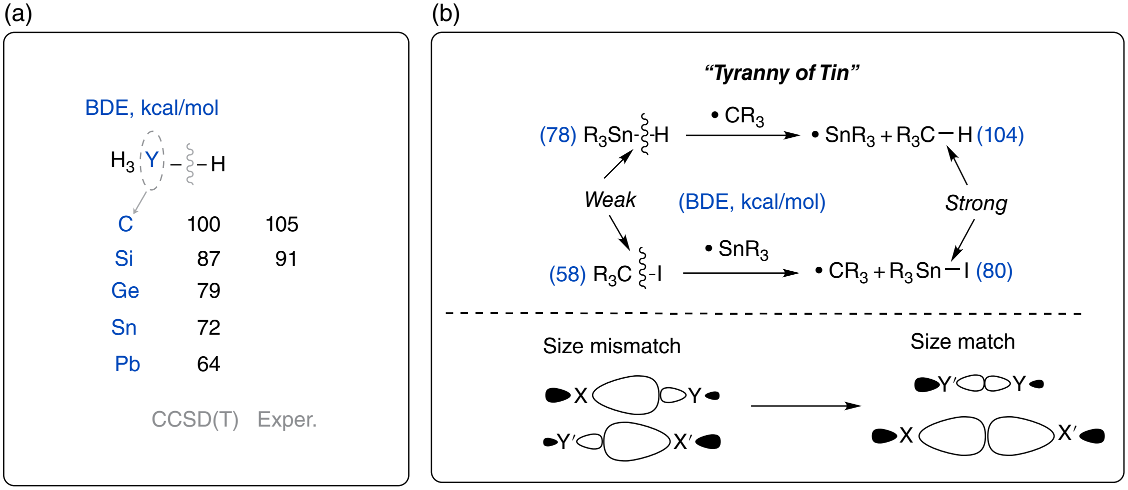

Figure 2.11 (a) Relative orbital size affects bond strength.[21] (b) Progression from weaker bonds to stronger bonds drives the radical chain process. BDEs for bonds involved in this transformation are shown in parenthesis. X, X′ are larger atoms, Y, Y′ are smaller atoms.

Steric effects Geometric restrictions to the σ-overlap inspired the elegant concept of frustrated Lewis pairs (FLPs).[22] The FLP concept takes advantage of steric effects to weaken chemical bonds, rendering such systems structurally "unsaturated" and catalytically active. The structural implications of steric "frustration" are shown in Figure 2.12.[23] Interestingly, even though the P…B distance is too long for the formation of a dative covalent bond (Figure 2.12), the combination of multiple C-H···F hydrogen bonds and dispersion interactions leads to an association energy of −11.5 kcal/mol (SCS-MP2). FLPs show enormous potential in activating small unreactive molecules such as H_2 and CO_2.

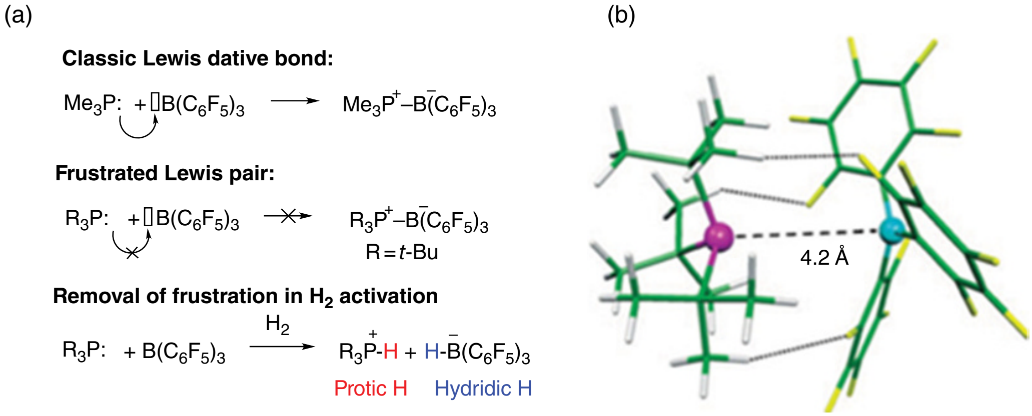

Figure 2.12 (a) Transition from dative bonds to frustrated Lewis pairs (FLPs) upon increase in the size of substituents at the donor and acceptor sites and utility of FLPs in H_2 activation. (b) Decreased overlap as a result of steric congestion in FLPs. The calculated structure of the $[(tBu)_3P]\cdots[B(C_6F_5)_3]$ complex (SCS-MP2 curve). C–H···F type hydrogen bonds (with d(H–F) < 2.4 Å) are indicated with dotted lines. The dashed line indicates distance between the "frustrated" atoms (Source: Rokob 2008 (23). Reproduced with permissions of John Wiley and Sons.)

An important insight into the nature of binding energies in sterically crowded molecules (including but not limited to FLPs) is provided by the work of Schreiner and coworkers on the role of dispersion effects. Such non-covalent interactions can significantly increase apparent bond strength in seemingly strained and unstable structures. [24] In an apparent paradox, surprisingly strong *and* long C-C bonds were found in a family of sterically congested alkanes (Figure 2.13).

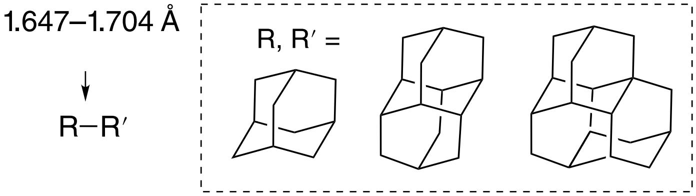

Figure 2.13 *Long, yet strong C-C bonds in sterically congested alkanes.*

These compounds are stable (up to 300 °C) despite having C-C bonds longer than 1.7 Å. A large part of the apparent bond strength is drawn not from the two atoms in the formal C-C bond but from numerous dispersive interactions. These results suggest that similar interactions can contribute significantly to the bonding energy in FLPs.

Directionality mismatch The importance of directionality in chemical bonding is illustrated by "inverted bonds" such as the central bond of [1.1.1]propellane (Figure 2.14). In such systems, strain and hybridization combine to weaken the bond.[25] The bonds are also weakened when the stereoelectronic requirement of collinearity is violated and orbitals forming a single bond are not directed along the shortest distance between atoms (i.e. "banana bonds" in small cycles).[8,26] Angle strain can be considered a negative stereoelectronic effect originating from suboptimal overlap of orbitals forming a σ-bond.

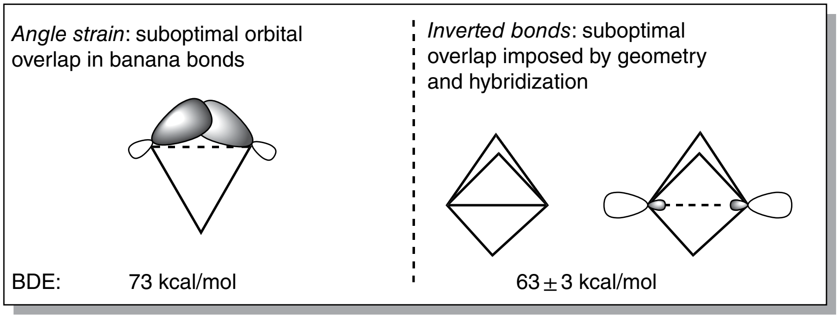

Figure 2.14 *Geometrical constraints leading to reduced overlap and weaker bonds in strained systems.*

The application of variable orbital overlaps can expand in unexpected directions. For example, it has been creatively utilized by Michl and coworkers in engineering excited state energies for singlet fission (transformation of an excited state singlet into two lower energy triplet states).[27] The application of this phenomenon towards the design of solar cells has the promise of significant increase in their maximum theoretical efficiency. The key energetic requirement for singlet fission is that the singlet excitation energy (S0→S1) should be approximately twice the first triplet excitation energy but lower than the energy of the second triplet state. It was shown that real chromophores satisfying these stringent photophysical conditions can be designed based on understanding of the evolution of biradicaloid energy states within a simple two-electron, two-orbital model. Two such models for the low energy states of H_2 and ethylene are shown in Figure 2.15. They illustrate how variations in the internuclear separation and in the double-bond twist angle control the relative energies of multiple excited states. This data illustrates that stereoelectronic effects in excited states can provide a new tool for scientists interested in utilizing solar energy for practical applications.

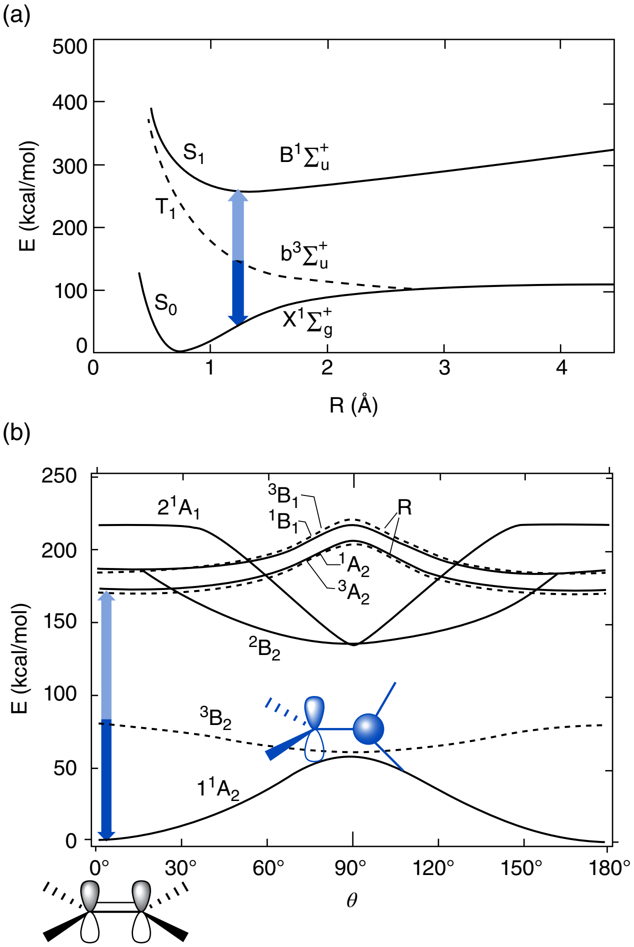

Figure 2.15 *Computed potential energy curves of low-energy states of H_2 (a) and ethylene (b), as a function of internuclear separation R and of twist angle θ, respectively. Rydberg states of ethylene are indicated by letter R. (Source: Wen 2015 (27). Reproduced with permission of American Chemical Society.)*

Ionic bonds In extreme cases, when electronegativity differences between two atoms (groups) are large, the covalent term becomes unimportant in comparison to the Coulombic attraction between ions of opposite charge formed by electron transfer from the more electropositive atom to the more electronegative partner. In addition to this general textbook scenario, it was suggested recently that ionic terms can play a significant role even in the formation of chemical bonds between two atoms of similar (or even identical) electronegativity. Such bonds, referred to as "charge-shift bonds" were suggested to occur when covalent overlap is inefficient but classic ionic bond is impossible (vide infra).

***Charge-shift bonds*:** Shaik and coworkers introduced the concept of charge-shift bonds for a variety of bonds, from the relatively weak F-F bond in F_2 to very strong C-F and Si-F bonds.[28] These bonds were suggested to originate from the superposition of two ionic resonances without significant contribution of covalent resonance where the electron pair is shared *between* the atoms. This idea is illustrated by comparison of the electron localization function (ELF) domains for F_2 and ethane that shows little electron density localization in the bonding region of F_2.

In the formalism of valence bond (VB) theory, the overall energy of an A-B bond can be represented by the three terms, one covalent and two ionic:

$$\Psi(\mathrm{VB}) = c_1\phi_{\mathrm{cov}}(\mathrm{A-B}) + c_2\phi_{\mathrm{ion}}(\mathrm{A^+B^-}) + c_3\phi_{\mathrm{ion}}(\mathrm{A^-B^+}) \tag{2.1}$$

The first term dominates for a covalent bond and one of the ionic terms plays the largest role for an ionic bond. For example, term two of Eq. (2.1) dominates when B is much more electronegative. In the case of charge-shift bonds, there is no single dominant term. In these systems, the covalent term is often weak whereas *both* of the ionic terms contribute at the same time, so the charge distribution is relatively balanced. If both ionic terms contribute equally, the system can be non-polar (i.e. F_2). Shaik and coworkers suggested that such bonds are ubiquitous and not limited to the exotic inverted bonds of propellanes discussed above. Many classic systems and bond types (i.e. C-F, H-F, Si-O, and others) are dominated by charge-shift bonding.

Some of the above conclusions are provocative: for example, the F-F molecule, commonly defined as having a covalent bond, would not be a bound molecule if covalent VB structure were included alone (Figure 2.16). This is, of course, a counterintuitive suggestion because F-F is a homonuclear bond, where static ionicity

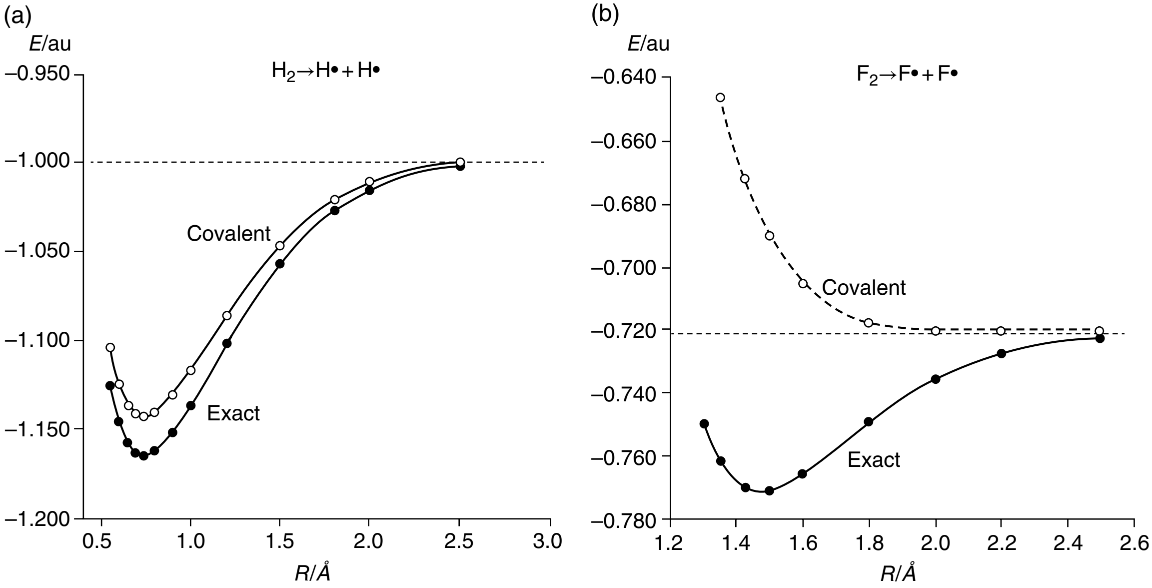

Figure 2.16 *Dissociation energy curves for (a) H_2 and (b) F_2. Open circles: the purely covalent VB structure. Bold lines, filled circles: the optimized covalent + ionic "exact" ground state. (Source: Shaik 2005 (28). Reproduced with permission of John Wiley and Sons.)*

should not matter. However, Shaik suggests that F_2 is sustained by the very large charge-shift resonance energy, where the bonding mostly originates from dynamic ionicity ("the ionic-covalent fluctuation of the electron pair density"). This idea is consistent with the abovementioned unusual hybridization in F-F bond (~>90% p-character), that is not optimized for efficient covalent bonding.

Control of π-overlap Although π-overlap allows binding interactions in two regions of space, the combined orbital overlap in a π-bond is weaker than σ-overlap. The weaker overlap accounts for the higher reactivity of alkenes. C-C π-bonds are so commonly used that it is easy to forget that they are highly strained. The bent bond description of alkenes reminds us that a double bond can be considered as the smallest cycle – with much more strain *per carbon* than cyclopropane or cyclobutane.

The energy cost for the formation of a π-bond can be described by the reaction energies in Figure 2.17. Redistribution of chemical bonds in two propane molecules to give $CH_2=CH_2$ and two ethanes increases the overall energy by 28 kcal/mol. The double bond is a truly high energy functionality! Similarly, the "formation" of the triple bond of ethyne from bond metathesis of 2-methylpropane (Figure 2.17, bottom) comes with the energy penalty of >65 kcal/mol. The high energy of alkynes accounts for many interesting features of this functional group.[29]

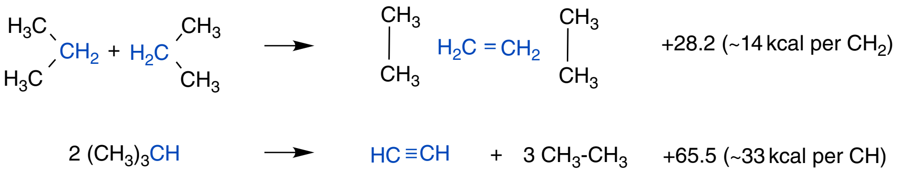

Figure 2.17 *Energy cost for formation of π-bonds. Calculated at the MP2/6–311++G(d,p) level of theory.*

Another illustration of the strain associated with π-bonds can be provided by thermodynamics of the dimerization of ethylene into cyclobutane (Figure 2.18). This process is enthalpically favored by 18.2 kcal/mol, suggesting that ethylene is a highly strained molecule even when compared to cyclobutane![30] Due to hybridization effects, the difference increases even further for fluorinated alkenes, accounting for their facile [2+2] reactions.[31]

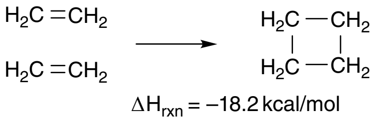

Figure 2.18 *The favorable thermodynamics of the dimerization of ethylene to cyclobutane illustrates high strain in alkene π-bonds.*

Variations in orbital overlap can make π-bonds stronger or weaker. For example, the shorter C-C distance in alkynes increases π-overlap and renders the π-bonds of alkynes stronger than the π-bonds of alkenes (Figure 2.19).[32,29]