

Topic III

Introduction to Structure and Bonding

III.1 Module 11: Quantum Chemistry 101

1/27:

- We will have a normal class on Friday and hold review sessions at different times where we can ask questions.
- Suggested readings: Nocera Lecture 6, Nocera Lecture 7, MIT OCW quantum mechanics^[1].
- In chemistry, most problems are solved with the time-independent Schrödinger equation $\hat{H}\Psi = E\Psi$.
 - Ψ is the wavefunction; it contains information on movement of the electron and its position.
 - $|\Psi(x, y, z)|^2 \propto P(x, y, z)$.
 - E is an eigenvalue of \hat{H} .
- If we are working with the time-dependent Schrödinger equation, we have another variable besides x, y, z , namely t . This allows us to calculate the probability that an electron is in a certain position at a given time.
- The Hamiltonian operator $\hat{H} = \hat{T} + \hat{V}$ describes the total energy.
 - \hat{T} is the kinetic energy operator. $\hat{T} = \frac{\hat{p}^2}{2m}$, where $\hat{p}_x = -i\hbar \frac{d}{dx}$ is the momentum operator.
 - \hat{V} is the potential energy operator. It typically describes the Coulombic attraction between the nucleus and the electron, which is approximately $\frac{1}{r}$ where r is the distance from the nucleus to the electron.
- For a free electron in one dimension, the Schrödinger equation reduces to

$$\begin{aligned}-\frac{\hbar^2}{2m} \frac{d^2\Psi}{dx^2} &= E\Psi \\ \frac{d^2\Psi}{dx^2} &= -\frac{2mE}{\hbar^2}\Psi\end{aligned}$$

- Dirac's bra-ket notation $\langle \Psi | A | \Psi \rangle \equiv \int_V \Psi^* \hat{H} \Psi \, dx \, dy \, dz$.
 - The **bra vector** (the first term inside the brackets) and **ket vector** (the last term inside the brackets) correspond to complex conjugates of the wave function.

¹Many chemistry courses go too deep into the math and physics of quantum mechanics, which obfuscates the chemistry and confuses us in Dr. Talapin's opinion.

- **LCAO method:** A way of finding the wavefunction of a molecule; of solving the Schrödinger equation after applying simplifications. Short for linear combination of atomic wavefunctions, i.e., the atomic orbitals ϕ .

$$\Psi = \sum_i c_i \phi_i$$

- Each c_i is a coefficient, and the atomic orbitals form the basis set.
- Basically, we think of the wave function of a molecule as a linear combination of its atomic orbitals.
- ϕ is normalized, thus $\int \phi_i^2 d\tau = 1$ where $d\tau = dx dy dz$.
- Continuing, we can calculate the expected value of \hat{H} :

$$E = \frac{\int \Psi \hat{H} \Psi d\tau}{\int \Psi^2 d\tau}$$

- Shortcomings: Does not count for electron correlation and a few other things.

- Electronic structure of H_2 molecule.

- H_2 's structure is H–H.
- $\Psi = a\phi_1 + b\phi_2$, where $\phi_{1,2}$ are two atomic hydrogen $1s$ orbitals.
- The electron density function is: $\phi^2 = a^2\phi_1^2 + b^2\phi_2^2 + 2ab\phi_1\phi_2$.
- By symmetry of H–H molecule, $a = \pm b$.
 - Symmetry of the coefficients should reflect symmetry of the atoms.
 - Hydrogen atoms are indistinguishable, so since the electron can't identify which atom it corresponds to, the math shouldn't either.
- If $S = \int_{\tau} \phi_1 \phi_2 d\tau$ or $\langle \phi_1 | \phi_2 \rangle$ is the overlap integral between two hydrogen $1s$ orbitals, we have bonding and antibonding orbitals:

$$\Psi_b = \frac{1}{\sqrt{2(1+s)}}(\phi_1 + \phi_2) \quad \Psi_a = \frac{1}{\sqrt{2(1-s)}}(\phi_1 - \phi_2)$$

- The first orbital is σ_g bonding.
- The second orbital is σ_u^* antibonding.
- Introducing the normalizing requirement gives us the above coefficients.
- In the Hückel theory:

$$\alpha = \langle \phi_1 | \hat{H}_{\text{eff}} | \phi_1 \rangle = \langle \phi_2 | \hat{H}_{\text{eff}} | \phi_2 \rangle \quad \beta = \langle \phi_1 | \hat{H}_{\text{eff}} | \phi_2 \rangle$$

- If we calculate the expectation integrals, we will arrive at the above.
- \hat{H}_{eff} is some effective Hamiltonian.
- The α integral is the **Coulomb integral**.
- The β integral is the **interaction integral**.
- In the **Hückel approximation** (the simplest approximation of quantum mechanics), we define integrals as parameters that we can extract from empirical data:
 - $H_{ii} = \alpha$.
 - $H_{ij} = 0$ for ϕ_i not adjacent to ϕ_j .
 - $H_{ij} = \beta$ for ϕ_i adjacent to ϕ_j .
 - $S_{ii} = 1$.
 - $S_{ij} = 0$.

- Expectation values for energy are

$$E_{a,b} = \frac{\langle \Psi_{a,b} | \hat{H}_{\text{eff}} | \Psi_{a,b} \rangle}{\langle \Psi_{a,b} | \Psi_{a,b} \rangle}$$

so

$$E_a = \frac{\alpha - \beta}{1 - s} \quad E_b = \frac{\alpha + \beta}{1 + s}$$

- Note that $\beta < 0$ for atomic s -orbitals and $\beta > 0$ for p -orbitals in σ -bonds.
- Also, in the Hückel one-electron model, the integrals α and β remain unsolved.
- Note: As always, the bonding orbitals are less stabilized than the antibonding orbitals are destabilized.
 - This is a consequence of overlap, e.g., for a dimer, the $1 \pm S$ term in $E_{+/-} = \frac{\alpha \pm \beta}{1 \pm S}$.
 - This is why He_2 does not exist.

- **Overlap integral:** An integral proportional to the degree of spatial overlap between two orbitals. It is the product of wave functions centered on different lattice sites. Varies from 0 (no overlap) to 1 (perfect overlap). *Also known as S .*
- **Coulomb integral:** An integral giving the kinetic and potential energy of an electron in an atomic orbital experiencing interactions with all the other electrons and all the positive nuclei. *Also known as α .*
- **Interaction integral** (on two orbitals 1,2): An integral giving the energy of an electron in the region of space where orbitals 1 and 2 overlap. The value is finite for orbitals on adjacent atoms, and assumed to be zero otherwise. *Also known as β_{12} , resonance integral, exchange integral.*
- Symmetry and quantum mechanics:

- Say we have $\hat{H}\Psi = E\Psi$ where \hat{H} is the Hamiltonian and R is a symmetry operator (e.g., C_2 or σ_v).
- Note that the Hamiltonian commutes with the symmetry operator: $R\hat{H} = \hat{H}R$.
- Since a symmetry operation does not change the energy of a molecule (it just moves it), $\hat{H}R\Psi_i = E_i R\Psi_i$.
- It follows that R does not change the form of the wave function, i.e., $R\Psi_i = \pm 1\Psi_i$. This reflects the fact that R cannot change the probability $P[e(x, y, z)] = |\Psi(x, y, z)|^2$ of finding an electron somewhere.
- Thus, the eigenfunctions of the Schrödinger equation generate a representation of the group.
- Non-degenerate wave functions are A or B type.
- Double-degenerate wave functions are E type.
- Triple-degenerate wave functions are T type.

- Back to the LCAO method:

$$E_i = \frac{\int \Psi_i^* H \Psi_i \, dV}{\int \Psi_i^* \Psi_i \, dV} \quad \Psi_i = \sum_i c_i \phi_i$$

- If we have a sizeable molecule with a couple dozen atoms, every molecular orbital (wave function) will be the sum of a couple dozen atomic orbitals.
- This generates a set of i linear homogenous equations, numbering in the hundreds or thousands that need to be solved.
- This is clearly too computationally expensive, so we need a trick.

- An example where symmetry arguments help a lot:
 - If f is odd ($f(x) = -f(-x)$), then we know that $\int_{-\infty}^{\infty} f(x) dx = 0$.
- Group theory allows us to generalize this method to broader symmetry operations.
- Three important theorems:
 1. The characters of the representation of a direct product are equal to the products of the characters of the representations based on the individual sets of functions.
 - For example, in the T_d point group, $T_1 = (3, 0, -1, 1, -1)$, and $T_2 = (3, 0, -1, -1, 1)$. By the theorem, $T_1 \times T_2 = (9, 0, 1, -1, -1)$.
 2. A representation of a direct product, $\Gamma_c = \Gamma_a \times \Gamma_b$, will contain the totally symmetric representation only if the irreducible representations of a and b contain at least one common irreducible representation.
 - Continuing with the above example, $T_1 \times T_2$ can be decomposed into $A_2 + E + T_1 + T_2$. Thus, by this theorem, if we take the product $\Gamma_c = E \times T_1 \times T_2$, the representation will contain the totally symmetric representation A_1 (since $\Gamma_b = E$ and $\Gamma_a = T_1 \times T_2$ contains E). Indeed, $E \times T_1 \times T_2 = (18, 0, 2, 0, 0)$.
 3. The value of any integral relating to a molecule $\int_V \Psi d\tau$ will be zero unless the integrand is invariant under all operations of the symmetry point group to which the molecule belongs. That is Γ_Ψ must contain the totally symmetric irreducible representation.
 - This example will concern the D_{4d} point group. We want to evaluate the integral $\int_V \Psi_a \mu_z \Psi_b d\tau$ where $\Gamma_{\Psi_a} = A_1$, $\Gamma_{\mu_z} = B_2$, and $\Gamma_{\Psi_b} = E_1$.
 - By Theorem 1, we can easily determine the representation $\Psi_a \times \mu_z$. We can then decompose it.
 - Noting that it does not contain the E_1 irreducible representation (the only representation in Ψ_b), we can learn from Theorem 2 that $\Psi_a \mu_z \Psi_b$ does not contain the A_1 irreducible representation.
 - Therefore, by Theorem 3, $\int_V \Psi d\tau = \int_V \Psi_a \mu_z \Psi_b d\tau = 0$.
- We use these three theorems to tell us what integrals will be zero in a much less computationally intensive fashion. We can then evaluate the remaining nonzero integrals.
- We can take direct products by hand, but there are also tables of direct products of irreducible representations.

III.2 Module 12: IR and Raman Active Vibrations (part 2)

- **Fermi's golden rule:** The rate of an optical transition from a single initial state to a final state is given by the transition rate for a single state.

$$\Gamma_{i \rightarrow f} = \frac{2\pi}{\hbar} E_0^2 | \langle f | H' | i \rangle |^2 \delta(E_f - E_i - h\nu)$$

- By state, we typically mean energy level.
- The transition rate is the probability of a transition happening.
- If it's an optical transition, conservation of energy implies that the energy difference between the initial and final state will equal the energy of the photon that the molecule absorbs or emits.
- E_0^2 is the light intensity.
- $h\nu$ is the photon energy.
- $\langle f | H' | i \rangle = \int \Psi_f^* H' \Psi_i d\tau$ is the square of the matrix element (the strength of the coupling between the states).

- $\delta(E_f - E_i - h\nu)$ is the resonance condition (energy conservation).
- In the dipole approximation, $H' = -e\vec{r} \cdot \vec{E}$.
- This is derived with time-dependent perturbation theory.
 - The matrix element $M = \langle f | H' | i \rangle = \int \Psi_f^*(\vec{r}) H' \Psi_i(\vec{r}) d^3\vec{r}$.
 - Perturbation: $H' = -\vec{p}_e \cdot \vec{E}_{\text{photon}}$. Dipole moment: $\vec{p}_e = -e\vec{r}$. Light wave: $\vec{E}_{\text{photon}}(r) = \vec{E}_0 e^{\pm i\vec{k} \cdot \vec{r}}$, $H'(\vec{r}) = e\vec{E}_0 \cdot \vec{r} e^{\pm i\vec{k} \cdot \vec{r}}$.
 - This implies that in one dimension, $|M| \propto \int \Psi_f^*(\vec{r}) x \Psi_i(\vec{r}) d^3\vec{r}$.
 - We include other variables in higher dimensions.
- For IR absorption, the intensity I satisfies $I \propto \int \Psi_{\text{e.s.}} \hat{\mu}_e \Psi_{\text{g.s.}} d\tau$.
 - $\Psi_{\text{e.s.}}$ is the excited state wavefunction, $\Psi_{\text{g.s.}}$ is the ground state wavefunction, and $\hat{\mu}_e$ is the dipole operator.
 - We now apply the three theorems:
 - It is always true in vibration spectroscopy that $\Gamma_{\text{g.s.}} = A_1$. This is because in the ground state, the molecule is completely relaxed (nothing is perturbed).
 - Thus, we can already reduce to $\Gamma_{\text{e.s.}} \cdot \Gamma_\mu \cdot \Gamma_{\text{g.s.}} = \Gamma_{\text{e.s.}} \cdot \Gamma_\mu \cdot 1$.
 - Now Γ_μ transforms as x, y, z unit vectors. In D_{3h} , this implies that $\Gamma_\mu = E' + A''_2$.
 - Therefore, $I \propto \Gamma_{\text{vibs}} \cdot (E' + A''_2)$.
 - For PF_5 , since $\Gamma_{\text{vibs}} = 2A'_1 + 3E' + 2A''_2 + E''$ has E' and A''_2 in common with Γ_μ , only $3E'$ and $2A''_2$ are IR active.
 - Additionally, with elements in common, $\Gamma_{\text{vibs}} \cdot \Gamma_\mu$ will contain A_1 by Theorem 2, and thus, the integrals $\int \Psi_{\text{e.s.}} x \Psi_{\text{g.s.}} d\tau$, $\int \Psi_{\text{e.s.}} y \Psi_{\text{g.s.}} d\tau$, and $\int \Psi_{\text{e.s.}} z \Psi_{\text{g.s.}} d\tau$ are all nonzero. Some linear combination of them will be proportional to I .
- The exam will include material from today's class, but not Friday's class.
- PSets 1 and 2 will cover all material on the exam?

III.3 Nocera Lecture 6

From Nocera (2008).

1/29: • Solving the Schrödinger equation with the LCAO method for the k th molecular orbital Ψ_k :

$$\begin{aligned} \hat{H}\Psi_k &= E\Psi_k \\ |\hat{H} - E| \Psi_k \rangle &= 0 \\ |\hat{H} - E| c_a \phi_a + c_b \phi_b + \dots + c_i \phi_i \rangle &= 0 \end{aligned}$$

- Left-multiplying the above by each ϕ_i yields a set of i linear homogenous equations.

$$\begin{aligned} c_a \langle \phi_a | \hat{H} - E | \phi_a \rangle + c_b \langle \phi_a | \hat{H} - E | \phi_b \rangle + \dots + c_i \langle \phi_a | \hat{H} - E | \phi_i \rangle &= 0 \\ c_a \langle \phi_b | \hat{H} - E | \phi_a \rangle + c_b \langle \phi_b | \hat{H} - E | \phi_b \rangle + \dots + c_i \langle \phi_b | \hat{H} - E | \phi_i \rangle &= 0 \\ &\vdots \\ c_a \langle \phi_i | \hat{H} - E | \phi_a \rangle + c_b \langle \phi_i | \hat{H} - E | \phi_b \rangle + \dots + c_i \langle \phi_i | \hat{H} - E | \phi_i \rangle &= 0 \end{aligned}$$

- We can then solve the **secular determinant**,

$$\begin{vmatrix} H_{aa} - ES_{aa} & H_{ab} - ES_{ab} & \cdots & H_{ai} - ES_{ai} \\ H_{ba} - ES_{ba} & H_{bb} - ES_{bb} & \cdots & H_{bi} - ES_{bi} \\ \vdots & \vdots & \ddots & \vdots \\ H_{ia} - ES_{ia} & H_{ib} - ES_{ib} & \cdots & H_{ii} - ES_{ii} \end{vmatrix} = 0$$

where $H_{ij} = \int \phi_i \hat{H} \phi_j d\tau$ and $S_{ij} = \int \phi_i \phi_j d\tau$.

- To evaluate these integrals, see the notes in Module 11 concerning the Hückel approximation.

- **Extended Hückel theory:** An alternate integral approximation method that includes all valence orbitals in the basis (as opposed to just the highest energy atomic orbitals), calculates all S_{ij} s, estimates the H_{ii} s from spectroscopic data (as opposed to a constant α), and estimates H_{ij} s from a simple function of S_{ii} , H_{ii} , and H_{ij} . This is a zero differential overlap approximation. *Also known as EHT.*

- A **semi-empirical** method.
- **Semi-empirical** (method): A method that relies on experimental data for the quantification of parameters.
- Other semi-empirical methods include CNDO, MINDO, and INDO.
- Hückel's method and LCAO example: Examine the frontier orbitals and their associated energies (i.e., determine eigenfunctions and eigenvalues, respectively) of benzene.

- We assume that the frontier MO's will be composed of LCAO of the $2p\pi$ orbitals.
- Using orbitals as our basis and noting that benzene is of the D_{6h} point group, we can determine that $\Gamma_{p\pi} = (6, 0, 0, 0, -2, 0, 0, 0, 0, -6, 2, 0)$.
- Using the decomposition formula, we can reduce $\Gamma_{p\pi}$ into $\Gamma_{p\pi} = A_{2u} + B_{2g} + E_{1g} + E_{2u}$. These are the symmetries of the MO's formed by the LCAO of $p\pi$ orbitals in benzene.
- With symmetries established, LCAOs may be constructed by “projecting out” the appropriate linear combination with the following projection operator, which determines the linear combination of the i th irreducible representation.

$$P^{(i)} = \frac{\ell_i}{h} \sum_R [\chi^{(i)}(R)] \cdot R$$

- ℓ_i is the dimension of Γ_i .
- h is the order.
- $\chi^{(i)}(R)$ is the character of Γ_i under operation R .
- R is the corresponding operator.

- To actually apply the above projection operator, we will drop to the C_6 subgroup of D_{6h} to simplify calculations. The full extent of mixing among ϕ_1 - ϕ_6 is maintained within this subgroup, but the inversion centers are lost, meaning that in the final analysis, the Γ_i s in C_6 will have to be correlated to those in D_{6h} .
- In C_6 , we have $\Gamma_{p\pi} = (6, 0, 0, 0, 0, 0) = A + B + E_1 + E_2$.
- The projection of the Symmetry Adapted Linear Combination (SALC) that from ϕ_1 transforms as A is

$$\begin{aligned} P^{(A)} \phi_1 &= \frac{1}{6} [1E + 1C_6 + 1C_6^2 + 1C_6^3 + 1C_6^4 + 1C_6^5] \phi_1 \\ &= \frac{1}{6} [E\phi_1 + C_6\phi_1 + C_6^2\phi_1 + C_6^3\phi_1 + C_6^4\phi_1 + C_6^5\phi_1] \\ &= \frac{1}{6} [\phi_1 + \phi_2 + \phi_3 + \phi_4 + \phi_5 + \phi_6] \\ &\cong \phi_1 + \phi_2 + \phi_3 + \phi_4 + \phi_5 + \phi_6 \end{aligned}$$

where we make the last congruency (dropping the constant) because the LCAO will be normalized, which will change the constant, regardless.

- With a similar process, we can find that

$$\begin{aligned} P^{(B)}\phi_1 &= \phi_1 - \phi_2 + \phi_3 - \phi_4 + \phi_5 - \phi_6 \\ P^{(E_{1a})}\phi_1 &= \phi_1 + \varepsilon\phi_2 - \varepsilon^*\phi_3 - \phi_4 - \varepsilon\phi_5 + \varepsilon^*\phi_6 \\ P^{(E_{1b})}\phi_1 &= \phi_1 + \varepsilon^*\phi_2 - \varepsilon\phi_3 - \phi_4 - \varepsilon^*\phi_5 + \varepsilon\phi_6 \\ P^{(E_{2a})}\phi_1 &= \phi_1 - \varepsilon^*\phi_2 - \varepsilon\phi_3 + \phi_4 - \varepsilon^*\phi_5 + \varepsilon\phi_6 \\ P^{(E_{2b})}\phi_1 &= \phi_1 - \varepsilon^*\phi_2 - \varepsilon^*\phi_3 + \phi_4 - \varepsilon\phi_5 - \varepsilon^*\phi_6 \end{aligned}$$

- Since some of the projections contain imaginary components, we can obtain real components by taking \pm linear combinations and noting that $\varepsilon = e^{2\pi i/6}$ in the C_6 point group.

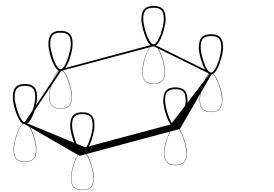
$$\begin{aligned} \Psi_3(E_1) &= \Psi'_3(E_{1a}) + \Psi'_4(E_{1b}) = 2\phi_1 + \phi_2 - \phi_3 - 2\phi_4 - \phi_5 + \phi_6 \\ \Psi_4(E_1) &= \Psi'_3(E_{1a}) - \Psi'_4(E_{1b}) = \phi_2 + \phi_3 - \phi_5 - \phi_6 \\ \Psi_5(E_2) &= \Psi'_5(E_{2a}) + \Psi'_6(E_{2b}) = 2\phi_1 - \phi_2 - \phi_3 + 2\phi_4 - \phi_5 + \phi_6 \\ \Psi_6(E_2) &= \Psi'_5(E_{2a}) - \Psi'_6(E_{2b}) = \phi_2 - \phi_3 + \phi_5 - \phi_6 \end{aligned}$$

- We can now normalize: If $\Psi_i = \sum_j c_j \phi_j$ where $\text{gcd}(c_1, \dots, c_n) = 1$, the normalizing constant is

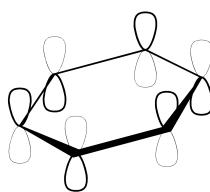
$$N = \frac{1}{\sqrt{\sum_j c_j^2}}$$

meaning that

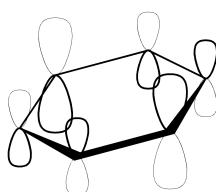
$$\begin{array}{ll} \Psi_1(A) = \frac{1}{\sqrt{6}} (\phi_1 + \phi_2 + \phi_3 + \phi_4 + \phi_5 + \phi_6) & \Psi_2(B) = \frac{1}{\sqrt{6}} (\phi_1 - \phi_2 + \phi_3 - \phi_4 + \phi_5 - \phi_6) \\ \Psi_3(E_1) = \frac{1}{\sqrt{12}} (2\phi_1 + \phi_2 - \phi_3 - 2\phi_4 - \phi_5 + \phi_6) & \Psi_4(E_1) = \frac{1}{2} (\phi_2 + \phi_3 - \phi_5 - \phi_6) \\ \Psi_5(E_2) = \frac{1}{\sqrt{12}} (2\phi_1 - \phi_2 - \phi_3 + 2\phi_4 - \phi_5 + \phi_6) & \Psi_6(E_2) = \frac{1}{2} (\phi_2 - \phi_3 + \phi_5 - \phi_6) \end{array}$$



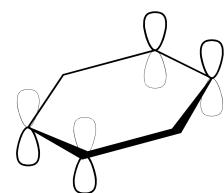
(a) $\Psi_1(A) \sim \Psi(A_{2u})$.



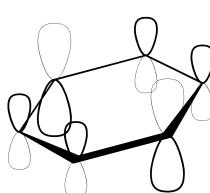
(b) $\Psi_2(B) \sim \Psi(B_{2g})$.



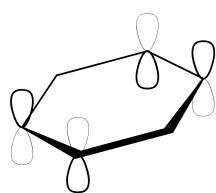
(c) $\Psi_3(E_1) \sim \Psi(E_{1g}^a)$.



(d) $\Psi_4(E_1) \sim \Psi(E_{1g}^b)$.



(e) $\Psi_5(E_2) \sim \Psi(E_{2u}^a)$.



(f) $\Psi_6(E_2) \sim \Psi(E_{2u}^b)$.

Figure III.1: Molecular orbitals of benzene.

- Figure III.1 shows pictorial representations of the SALCs.

III.4 Nocera Lecture 7

From Nocera (2008).

- This lecture continues with the benzene example from Nocera Lecture 6.
- Finding the total energy of benzene:
 - The energies (eigenvalues of the individual wavefunctions) may be determined using the Hückel approximation as follows.

$$\begin{aligned}
 E(\Psi_{A_{1g}}) &= \int \Psi_{A_{1g}} \hat{H} \Psi_{A_{1g}} d\tau \\
 &= \langle \Psi_{A_{1g}} | \hat{H} | \Psi_{A_{1g}} \rangle \\
 &= \left\langle \frac{1}{\sqrt{6}} (\phi_1 + \phi_2 + \phi_3 + \phi_4 + \phi_5 + \phi_6) \middle| \hat{H} \middle| \frac{1}{\sqrt{6}} (\phi_1 + \phi_2 + \phi_3 + \phi_4 + \phi_5 + \phi_6) \right\rangle \\
 &= \frac{1}{6} \left((H_{11} + H_{12} + H_{13} + H_{14} + H_{15} + H_{16}) + (H_{21} + H_{22} + H_{23} + H_{24} + H_{25} + H_{26}) + \sum_{i=3}^6 \sum_{j=1}^6 H_{ij} \right) \\
 &= \frac{1}{6} \left((\alpha + \beta + 0 + 0 + 0 + \beta) + (\beta + \alpha + \beta + 0 + 0 + 0) + \sum_{i=3}^6 (\alpha + 2\beta) \right) \\
 &= \frac{1}{6}(6)(\alpha + 2\beta) \\
 &= \alpha + 2\beta
 \end{aligned}$$

- Similarly, we can determine that

$$\begin{aligned}
 E(\Psi_{B_{2g}}) &= \alpha - 2\beta \\
 E(\Psi_{E_{1g}^a}) &= E(\Psi_{E_{1g}^b}) = \alpha + \beta \\
 E(\Psi_{E_{2u}^a}) &= E(\Psi_{E_{2u}^b}) = \alpha - \beta
 \end{aligned}$$

- We can now construct an energy level diagram (Figure III.2). We set $\alpha = 0$ and let β be the energy parameter (a negative quantity; thus, a MO whose energy is positive in units of β has an absolute energy that is negative).

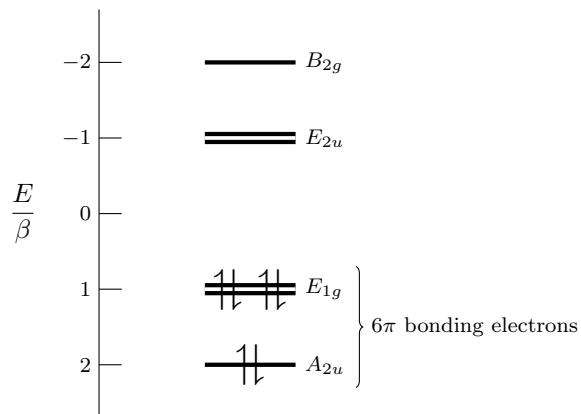


Figure III.2: Energy level diagram of benzene.

- From Figure III.2, we can determine that the energy of benzene based on the Hückel approximation is

$$E_{\text{total}} = 2(2\beta) + 4(\beta) = 8\beta$$

- **Delocalization energy:** The difference in energy between a molecule that delocalizes electron density in a delocalized state versus a localized state. *Also known as resonance energy.*

- Finding the delocalization energy of benzene:

- Consider cyclohexatriene, a molecule equivalent to benzene except that it has 3 *localized* π bonds. Cyclohexatriene is the product of three condensed ethene molecules.

- Ethene has 2 π bonds ϕ_1 and ϕ_2 .

- Following the procedure of Nocera Lecture 6, we can determine that

$$\Psi_1(A) = \frac{1}{\sqrt{2}}(\phi_1 + \phi_2) \quad \Psi_2(B) = \frac{1}{\sqrt{2}}(\phi_1 - \phi_2)$$

- Thus,

$$E(\Psi_1) = \left\langle \frac{1}{\sqrt{2}}(\phi_1 + \phi_2) \middle| \hat{H} \middle| \frac{1}{\sqrt{2}}(\phi_1 + \phi_2) \right\rangle = \frac{1}{2}(2\alpha + 2\beta) = \beta$$

$$E(\Psi_2) = \left\langle \frac{1}{\sqrt{2}}(\phi_1 - \phi_2) \middle| \hat{H} \middle| \frac{1}{\sqrt{2}}(\phi_1 - \phi_2) \right\rangle = \frac{1}{2}(2\alpha - 2\beta) = -\beta$$

- Correlating the above calculations (performed within $C_2 \subset D_{2h}$) to the D_{2h} point group gives $A \rightarrow B_{1u}$ and $B \rightarrow B_{2g}$.

- We can now construct an energy level diagram.

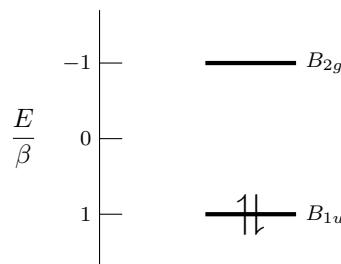


Figure III.3: Energy level diagram of ethene.

- Figure III.3 tells us that $E_{\text{total}} = 2(\beta) = 2\beta$. Consequently, the total energy of cyclohexatriene is $3(2\beta) = 6\beta$.
- Therefore, the resonance energy of benzene based on the Hückel approximation is

$$E_{\text{res}} = 8\beta - 6\beta = 2\beta$$

- **Bond order:** A quantity defined for a given bond as

$$\text{B.O.} = \sum_{i,j} n_e c_i c_j$$

where n_e is the orbital e^- occupancy and $c_{i,j}$ are the coefficients of the electrons i, j in a given bond.

- Finding the bond order of benzene between carbons 1 and 2:

- Just apply the formula:

$$\begin{aligned} \text{B.O.} &= [\Psi_1(A_2)] + [\Psi_3(E_{1g}^a)] + [\Psi_4(E_{1g}^b)] \\ &= (2) \left(\frac{1}{\sqrt{6}} \right) \left(\frac{1}{\sqrt{6}} \right) + (2) \left(\frac{2}{\sqrt{12}} \right) \left(\frac{1}{\sqrt{12}} \right) + (2)(0) \left(\frac{1}{2} \right) \\ &= \frac{1}{3} + \frac{1}{3} + 0 \\ &= \frac{2}{3} \end{aligned}$$

III.5 Module 13: Why Molecular Orbitals?

- Note that in the point group flow chart, there is no D_{nv} ; only D_{nd} .
- We will not need LCAO for the exam tomorrow.
- The exam covers Modules 1-12.
- Lewis (1916) first proposed that bonds came from interpenetrability of electron density.
- The next step came from Linus Pauling, who proposed valence bond theory.
 - In order to account for polarity of the bond, he created a term that described the probability that both electrons bond to one atom.
- A new approach then emerged from Robert Mulliken, Friedrich Hund, and Clemens C.J. Roothaan. All three men worked at UChicago!
 - Mulliken is mainly credited for the development of MO theory.
 - Roothaan retired, found retirement boring, moved to Palo Alto and was key in the development of computer processors.
- Molecular orbital theory:
 - Atomic orbitals of different atoms combine to create molecular orbitals.
 - The number of atomic orbitals equals the number of molecular orbitals.
 - Electrons in these molecular orbitals are shared by the molecule as a whole.
 - Molecular orbitals can be constructed from LCAO.
 - For diatomic molecules: $\Psi = c_a \Psi_a + c_b \Psi_b$.
- There is no such thing as a chemical bond (this model is only intuitively helpful), only molecular orbitals!
- **Bonding** (orbital): An orbital that has most of the electron density between the two nuclei.
- **Anti-bonding** (orbital): An orbital that has a node between the two nuclei.
- **Nonbonding** (orbital): An orbital that is essentially the same as if it was only one nucleus.
- We find the energy of electronic states using theoretical calculations that we test with photoelectron spectroscopy^[2].
- **Photoelectron spectroscopy**: A photo-ionization and energy-dispersive analysis of the emitted photoelectrons to study the composition and electronic state of a sample. *Also known as PES.*
 - A sample (solid, liquid, or gas) is impinged upon by a focused beam of X-rays (say of 1.5 kV).
 - When the sample is exposed to the X-rays, electrons fly out of the sample. The KE of these electrons can be measured.
 - Essentially, $h\nu$ takes an electron from the core level to above the vacuum level. We know $h\nu$ and we measure KE_{electron} , allowing us to calculate the bonding energy of the electron: $h\nu = I_{\text{BE}} + E_{\text{kinetic}}$ (see Figure III.4).
- **X-ray photoelectron spectroscopy**: Using soft (200-2000 eV) x-ray excitation (photons in the x-ray energy range) to examine core levels. *Also known as XPS.*
- **Ultraviolet photoelectron spectroscopy**: Using vacuum UV (10-45 eV) radiation (photons in the UV energy range) from discharge lamps to examine valence levels. *Also known as UPS.*

²Refer to Labalme (2020), specifically Figure 7.20 and the accompanying discussion.

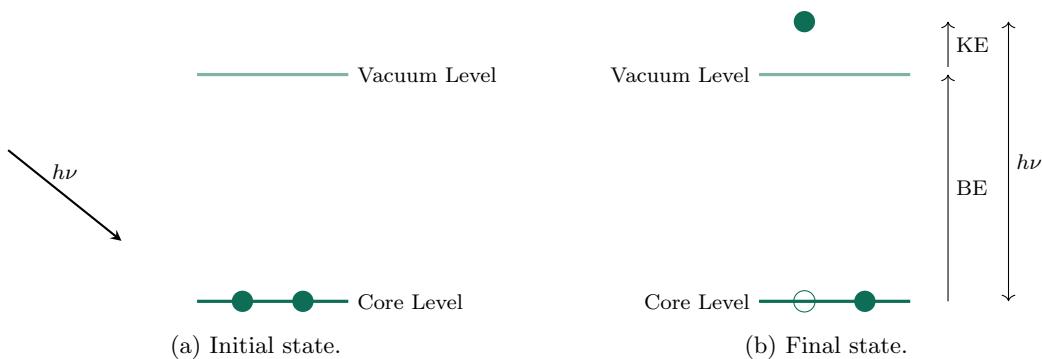


Figure III.4: Photoelectron spectroscopy at an atomic level.

- If we apply PES to O₂, we get counts that correspond to molecular orbitals π_g^* , π_u , σ_g , and σ_u^* .
- Photoelectron spectrum of H₂O:
 - Pauling's theory suggest that the lone pairs should have equal energy greater than the equal energy of the bonds.
 - However, PES reveals that the lone pairs have two different energies. This is a nail in the coffin of Pauling's valence bond theory.
- PES of CH₄:
 - There are two states; one with degeneracy 3 and one with degeneracy 1.
 - We have 3 bonds of one energy and 1 with another.
 - Another nail in the coffin.

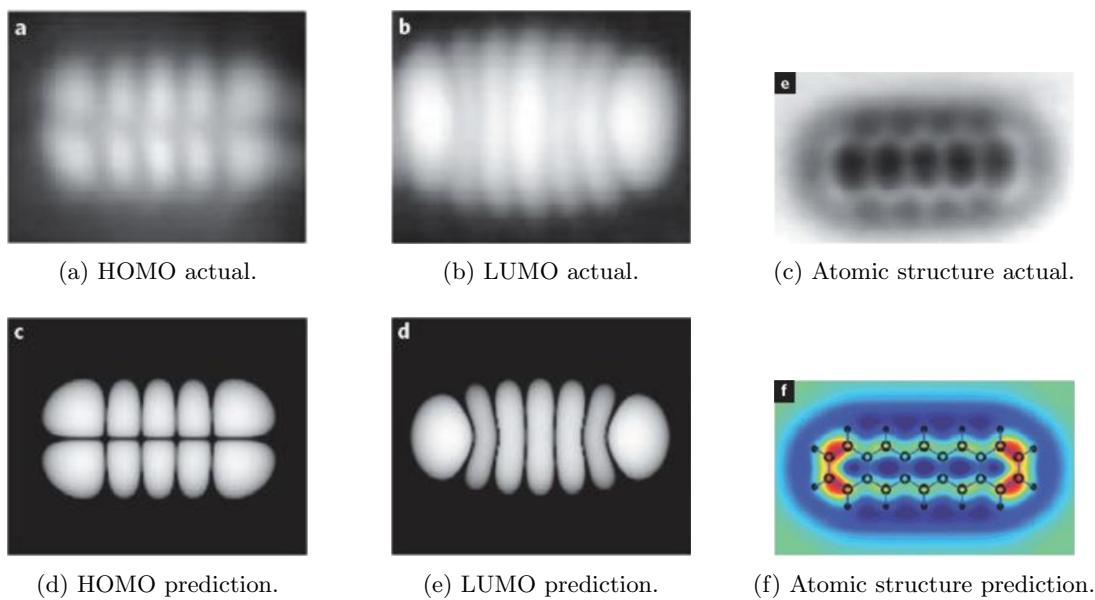


Figure III.5: Correspondence between MO predictions and scanning tunnelling microscopy.

- Can one “see” molecular orbitals? With a scanning tunneling microscope, we can “see” pentatene (5 linearly fused benzene rings). The correspondence between the pictures and MO theory’s predictions is impressive (see Figure III.5).

III.6 Module 14: Constructing Molecular Orbitals (Part 1)

- Bonding: $\Psi_\sigma = \Psi_+ = \frac{1}{\sqrt{2}}(\psi_{1s_a} + \psi_{1s_b})$.
- Anti-bonding: $\Psi_{\sigma^*} = \Psi_- = \frac{1}{\sqrt{2}}(\psi_{1s_a} - \psi_{1s_b})$.
 - Addition doesn't necessarily correlate to bonding and subtraction to anti-bonding.
- With simple orbitals, we can combine orbitals by inspection.
 - However, we will learn to build molecular orbitals for much more complicated molecular orbitals, such as those of ferrocene ($\text{Fe}(\text{C}_5\text{H}_5)_2$).
- Degree of orbital overlap/mixing depends on:
 1. Energy of the orbitals (the closer the energy, the more mixing; when the energies differ greatly, the reduction energy due to bonding is insignificant).
 2. Spatial proximity (the atoms must be close enough that there is *reasonable* orbital overlap, but not so close that repulsive forces interfere).
 3. Symmetry (atomic orbitals mix if they have similar symmetries; regions with the same sign of Ψ overlap).
- The strength of the bond depends upon the degree of orbital overlap.
- For heteronuclear molecules:

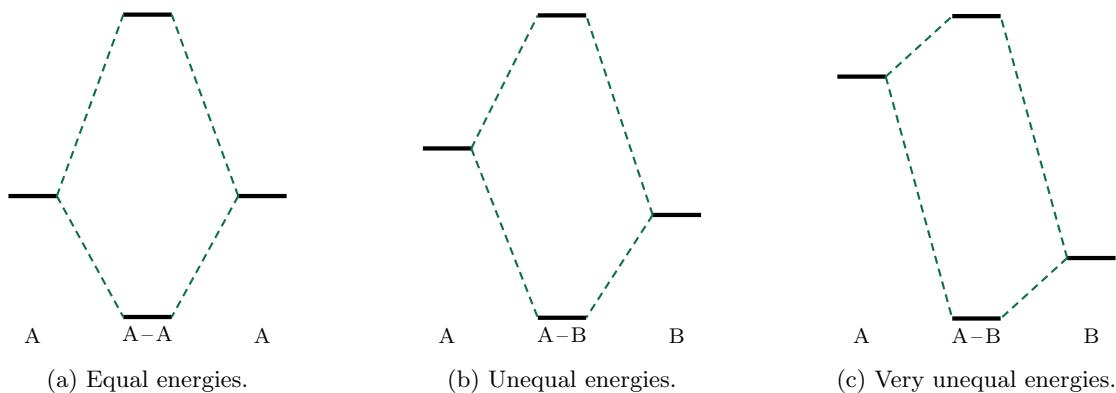


Figure III.6: Combining orbitals of varying energies.

- The bonding orbital(s) will reside predominantly on the atom of lower orbital energy (the more electronegative atom).
- The anti-bonding orbital(s) will reside predominantly on the atom with greater orbital energy (the less electronegative atom).
- The energies of atomic orbitals (measured by PES) have been tabulated (see Figure III.7).
- If you want to measure orbital energies in the range of -10 eV (i.e., upper valence orbitals; see Table 5.2 in Miessler et al. (2014)), use UPS. If you want to look at energy states that are very deep, very core (i.e., 1s in Fe), use XPS.
- Symmetry and orbital diagrams (suggested reading Cass and Hollingsworth (2004)):
 - **State conservation principle:** The number of molecular orbitals is equal to the number of incipient (atomic) orbitals.

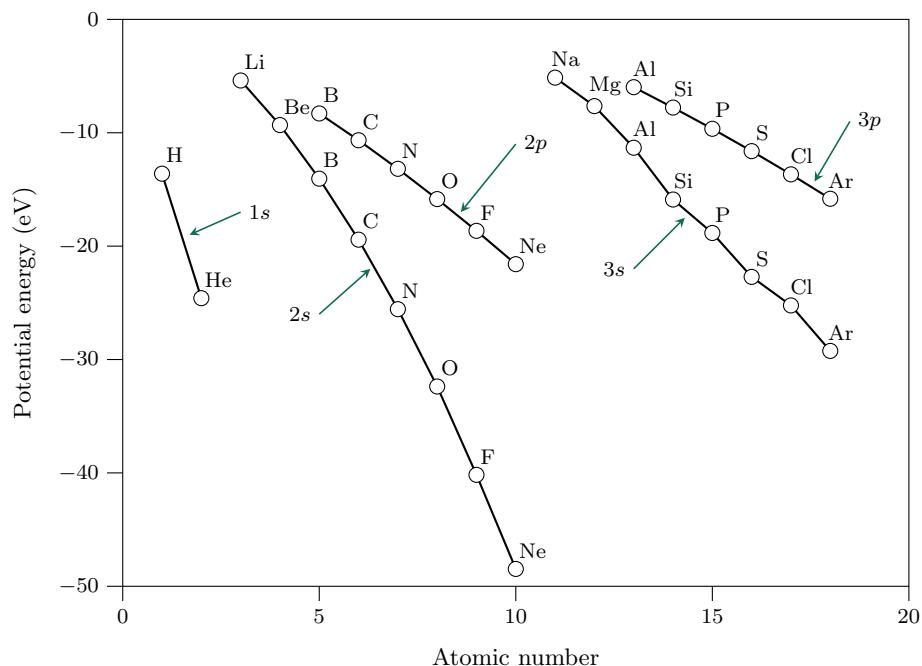
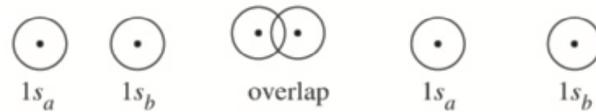
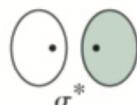


Figure III.7: Orbital potential energies.

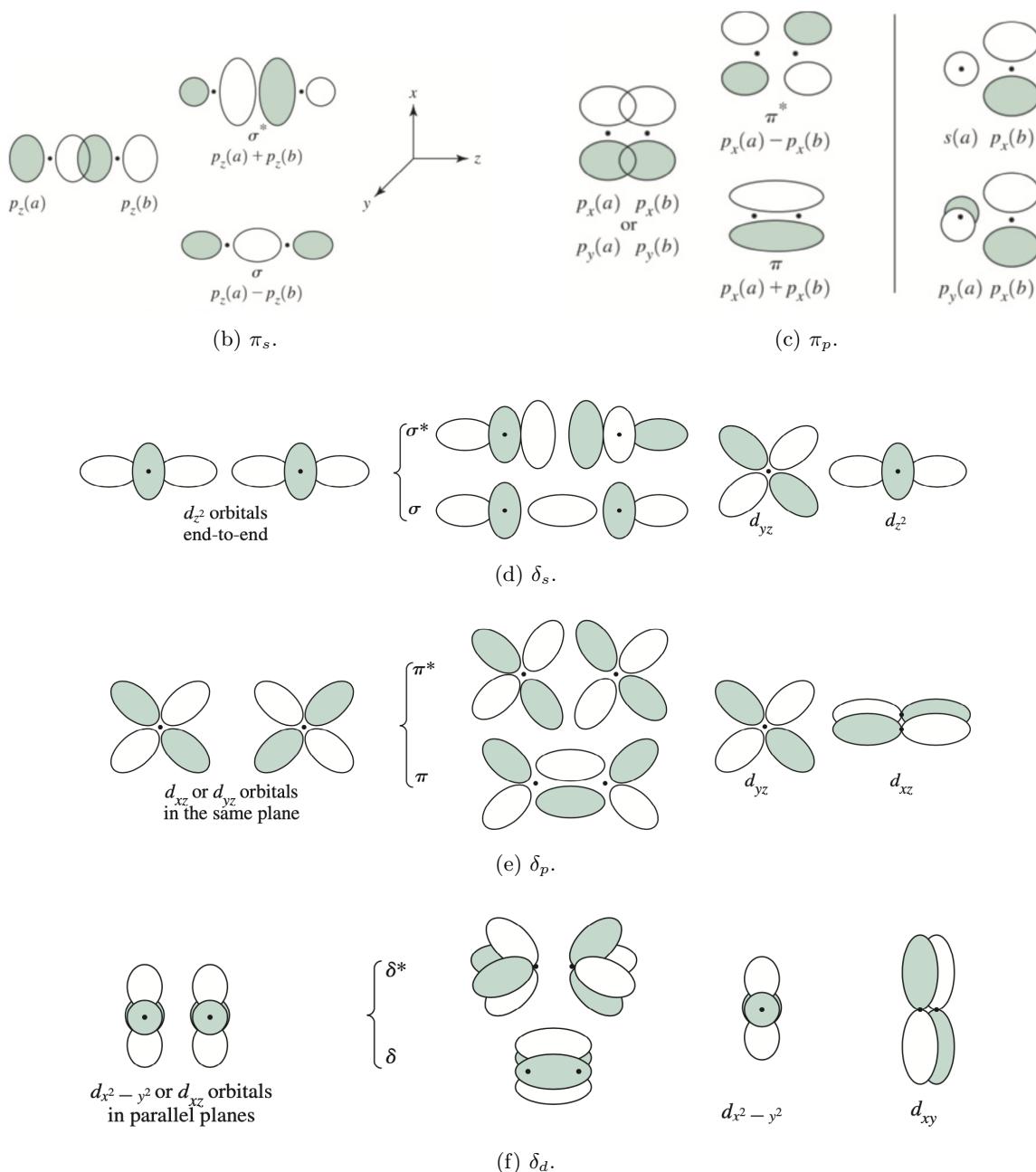
- Orbitals of the same symmetry mix.
- Orbital interactions can be bonding, nonbonding, or antibonding.
- There are three basic types of orbital overlap: σ (end-on interaction), π (side-by-side approach) and δ (off-axis approach).
 - σ orbitals are symmetric to rotation about the line connecting nuclei.
 - π orbitals change sign of the wave function with C_2 rotation about the bond axis.
 - Orbitals also denoted g are symmetric to inversion.
 - Orbitals also denoted u are antisymmetric to inversion.
- Orbitals with the correct symmetry and most similar energy mix to the greatest extent.

$$\sigma^* = \frac{1}{\sqrt{2}} [\psi(1s_a) - \psi(1s_b)]$$



$$\sigma = \frac{1}{\sqrt{2}} [\psi(1s_a) + \psi(1s_b)]$$

(a) σ_s .

Figure III.8: Constructing s , p , and d molecular orbitals.

- There are six MO constructions to be aware of: $s-s$, $p-p$ (σ and π), and $d-d$ (σ , π , and δ). See Figure III.8.
 - There are similar orbitals in simple diatomic molecules.
- As the mixing of σ_g orbitals gets stronger, the $2p$ state drops in energy faster than the $2s$ state, causing the $2p$ state to have lower energy than the $2s$ state after a while.

III.7 Office Hours (Wang)

- Quantum mechanics will not be included on tomorrow's exam.
- Hybridization was developed (1931 by Pauling) earlier than VSEPR theory (1940 by Sidgwick and Powell).
- What *is* a direct product of representations?
 - The direct product generates an $n \times m$ matrix?
 - Di has no idea what's going on, but he's gonna give me a link to his Advanced Inorganic Chemistry textbook.
- Can you go over how to do problems IV and VI with direct product analysis?
 - The IR absorption way from Module 12?
 - $\Gamma_{\text{g.s.}} = A_1$ always.
 - Γ_μ is the sum of the x, y, z linear irreducible representations.
 - $\Gamma_{\text{e.s.}} = \Gamma_{\text{vibs}}$.
 - We need to calculate the direct product of every representation in Γ_{vibs} with every term in Γ_μ . Products that contain A_1 are IR active?
 - Example:

$$A_1 \times (A_1 + B_1 + B_1) = A_1 \times A_1 + A_1 \times B_1 + A_1 \times B_2$$
 - Since $A_1 \times A_1 = A_1$, we already know after taking only this product that A_1 is IR active.
 - Things aren't IR active because they're linear. Things are IR active because their direct product with the sum of the linear groups contains the totally symmetric representation.
- You can only take the direct product of irreducible representations?

III.8 Office Hours (Talapin)

- To what extent do we need to know term symbols (they appeared briefly in Module 1)?
- How do we identify the T_h point group?

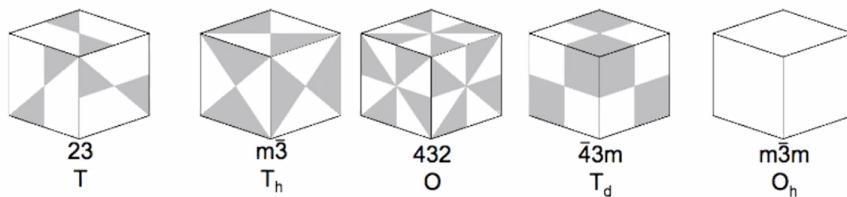


Figure III.9: Tetrahedral point groups.

- It's a very rare point group. It has no dihedral planes.
- What is the point of conjugate elements?
 - Conjugate elements allow us to group symmetry elements into classes — conjugate elements are in the same class!
- Will we need to be able to work with infinite character tables?
- What was that whole thing you did with a symmetry operation R and the Schrödinger equation?

- What level of familiarity do we need with Dirac's bra-ket notation?
 - Just know that it represents an integral.
- Do we need to evaluate/what do we need to know about those integrals from Wednesday's class?
 - It will be sufficient to write down representations.
- Why does Γ_μ , the representation of the dipole moment, transform with x, y, z ?
 - n -degree perturbation theory.
- The problems will be similar to homework problems.

III.9 Module 15: Constructing Molecular Orbitals (Part 2, HF Molecule)

2/1:

- MO Diagrams from Group Theory:
 1. Assign a point group.
 2. Choose basis functions (orbitals).
 3. Apply operations.
 4. Generate a reducible representations and SALCs (the latter if applicable).
 5. Reduce to irreducible representations.
 6. Combine central and peripheral orbitals by their symmetry.
 7. Fill MOs with e^- 's. Draw orbitals.
 8. Generate SALCs of peripheral atoms (if applicable).
 9. Draw peripheral atom SALC with central atom orbital to generate bonding/antibonding MOs (if applicable).
- We will not need SALCs here.
- H–F example:
 - Point group: $C_{\infty v}$. However, we will work within the C_{2v} subgroup (knowing why to choose this specific subgroup will come later; just accept it for now).
 - Choose basis functions (H_{1s} , F_{1s} , F_{2p_x} , F_{2p_y} , and F_{2p_z}).
 - Applying operations, we get

$$\begin{aligned}\Gamma_{H_{1s}} &= (1, 1, 1, 1) = A_1 \\ \Gamma_{F_{1s}} &= (1, 1, 1, 1) = A_1 \\ \Gamma_{F_{2p_z}} &= (1, 1, 1, 1) = A_1 \\ \Gamma_{F_{2p_x}} &= (1, -1, 1, -1) = B_1 \\ \Gamma_{F_{2p_y}} &= (1, -1, -1, 1) = B_2\end{aligned}$$

- Therefore, the H_{1s} , F_{1s} , F_{2p_x} , F_{2p_y} , and F_{2p_z} orbitals transform with the A_1 , A_1 , A_1 , B_1 , and B_2 irreducible representations, respectively.
- Note that we can also extract p -orbital information from the corresponding linear functions in the character table.
- Also note that one orbital \Rightarrow one degree of freedom. We do not have x, y, z -DOFs as before.

- When we bring the atoms together, those orbitals with similar symmetry (in this case, all the A_1 orbitals) will start mixing. However, from Figure III.7, the F_{1s} orbital has a very different energy (much lower) than the H_{1s} orbital, meaning that it has negligible mixing with H_{1s} . On the other hand, H_{1s} along with the $2p$ -orbitals in fluorine have comparable energies, so they will have significant mixing.
- We will get a low energy MO from F_{1s} , a higher bonding MO from H_{1s} and F_{2p_z} , two higher nonbonding MOs from F_{2p_x} and F_{2p_y} , and a higher anti-bonding MO from H_{1s} and F_{2p_z} .

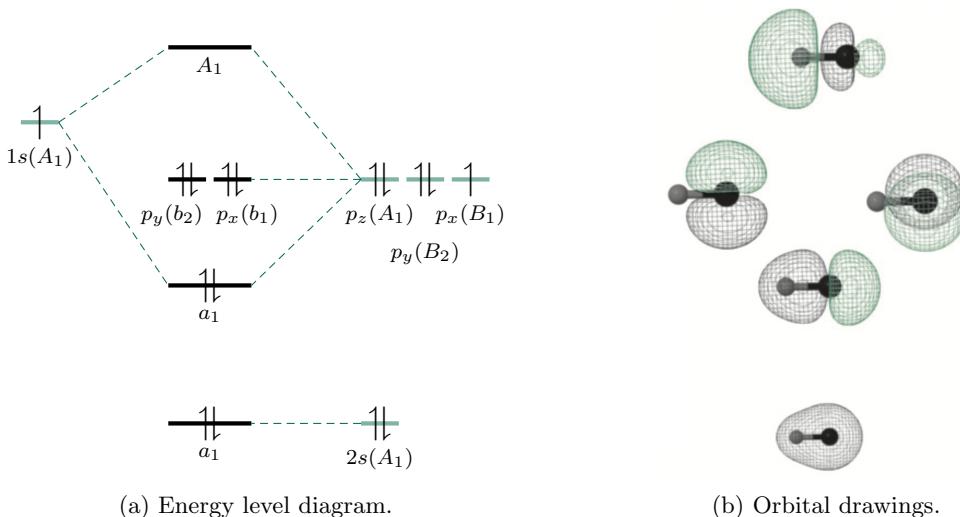


Figure III.10: HF orbital diagram.

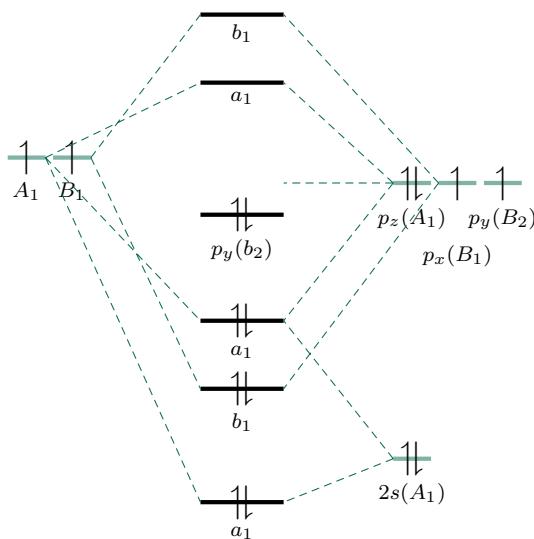
- We can now fill the MOs with the atomic electrons using the Aufbau principle, the Pauli exclusion principle, and Hund's rule. We can also draw these orbitals.
- Orbital energies can be found in Table 5.2 of Miessler et al. (2014).
- If the energy difference between orbitals is more than 10 eV, then we can *probably* (not always) ignore mixing.
 - The approximate scaling is that the stabilization energy (or energy gain) is inversely proportional to the energy gap.
 - The magnitude of the interaction integral will be approximately inversely proportional to the energy gap between the atomic orbitals participating.
- Organic chemists can go a long way with a hybridization approach even though it is not reflected in inorganic chemistry.

III.10 Module 16: Constructing Molecular Orbitals (Part 3, H_2O Molecule)

- Beware: Orientation of the point groups C_{2v} and D_{2h} .
 - Unlike all other groups, C_{2v} and D_{2h} present problems in assigning the irreducible representations. For most groups, the symmetry axis is obvious, or if there are several axes, the principal axis is obvious. For C_{2v} and D_{2h} , an ambiguity exists. The commonly (but not unanimously) used convention is the following:
 - If there are three C_2 axes, the one with the largest number of atoms unmoved by a C_2 operation is z . If there is only one C_2 axis, that is z .

- Once z is defined, the y -axis is defined as the axis of the remaining two axes which has the largest number of atoms unmoved by the σ symmetry operation.
- The x -axis is the remaining axis.
- The overall result is that the symmetry axes in ethylene are defined as: z - along the C=C bond; y - in the molecular plane, perpendicular to the C=C bond; and x - out of the plane.
- Similar approaches apply to the C_{2v} point group.
- We will need SALCs here.
 - Since the hydrogen atoms do not lie on the principal rotation axis, they will not individually obey the symmetry operations.
 - Thus, we need **Symmetry Adapted Linear Combinations**.
- **Symmetry Adapted Linear Combination** (of atomic orbitals): The linear combination of multiple atomic orbitals corresponding to peripheral atoms (atoms that do not lie on the principal axis). This linear combination will obey the symmetry modes. *Also known as SALC.*
- To generate SALCs, the steps are:
 - Group the atomic orbitals in the molecule into sets which are equivalent by symmetry.
 - Generate and reduce the reducible representation for each set.
 - Use the projection operator for one basis.
- H_2O example:
 - Point group: C_{2v} .
 - Choose basis functions (both H_{1s} orbitals, O the same as F in the last example).
 - Applying operations, we get

$$\Gamma_H = (2, 0, 2, 0) = A_1 + B_1$$
 - For the hydrogens, it makes sense that we need two irreducible representations to describe two atoms. More formally, the number of basis functions should match the number of irreducible representations to account for possible degeneracy of irreducible representations.
 - The O orbitals are the same as the F orbitals in the last example.
 - Combining orbitals by their symmetry and energy again, we get the following MOs.

Figure III.11: H_2O orbital diagram.

- We can now compare this with the results from photoelectron spectroscopy and see that our predictions are correct.

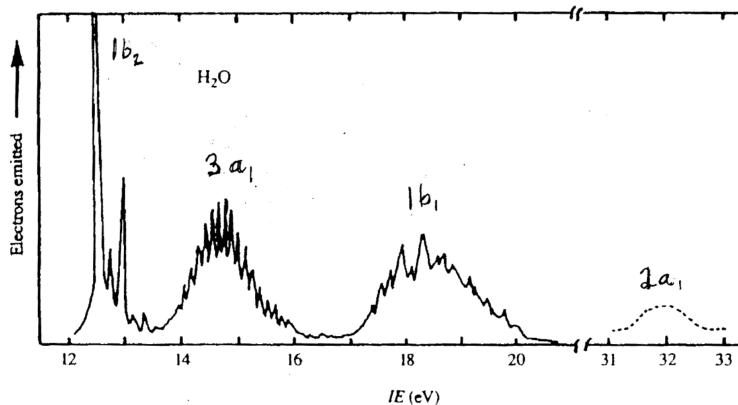


Figure III.12: Photoelectron spectrum for H₂O.

- We see four different states, which matches the prediction of Figure III.11.
- The lowest state is called 2a₁ because it is the second state that transforms as A₁ going out from the core (the first is 1a₁ corresponding to oxygen's 1s electrons [this state is not shown in Figure III.11 because it is so core as to not be significantly relevant to the chemistry of H₂O]).
- Then 1b₁ is the first A₁ state going out from the core, 3a₁ is the third A₁ state, and 1b₂ is the first B₂ state.
- There are no 4a₁ and 2b₁ electrons (see Figure III.11); hence, there are no such PES peaks (see Figure III.12).
- Note that the numbers along the bottom axis correspond to the orbital potential energies of the corresponding molecular orbitals.
- For H₂O, hybridization is qualitatively wrong.
- To generate SALCs:
 - Use the projection operator. See Nocera Lecture 6 for how to mathematically apply it.
 - Projection operators constitute a method of generating the symmetry allowed combinations.
 - Taking one AO and projecting it out using symmetry.
- Using the projection operator method to reconstruct SALC orbitals (NH₃ example):

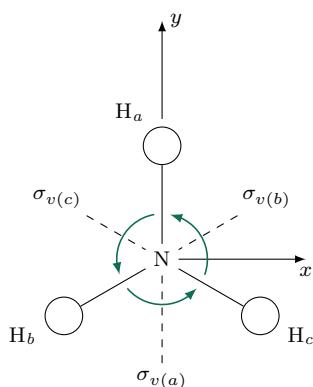


Figure III.13: Coordinate system for NH₃.

C_{3v}	E	$2C_3$	$3\sigma_v$	linear	quadratic
A_1	1	1	1	z	$x^2 + y^2, z^2$
A_2	1	1	-1	R_z	
E	2	-1	0	$(x, y), (R_x, R_y)$	$(x^2 - y^2, xy), (xz, yz)$

Table III.1: Character table for the C_{3v}

- Ungroup the operations in the classes in Table III.1. Choose H_a and see into which H it projects under each operation. If we order the operations $E, C_3, C_3^2, \sigma_{v(a)}, \sigma_{v(b)}, \sigma_{v(c)}$, then H_a becomes $H_a, H_b, H_c, H_a, H_c, H_b$, respectively.
- Multiply each projected atom by the corresponding character and sum. Thus, for the three representations, we have

$$\begin{aligned} A_1 &= H_a + H_b + H_c + H_a + H_c + H_b = 2H_a + 2H_b + 2H_c \\ A_2 &= H_a + H_b + H_c - H_a - H_c - H_b = 0 \\ E &= 2H_a - H_b - H_c + 0 + 0 + 0 = 2H_a - H_b - H_c \end{aligned}$$

- We have essentially just done what the projection operator would enable us to do.

- Continuing with the H_2O example:

- Using the projection operator, we get

$$\begin{aligned} P^{A_1} &= \frac{1}{4}[1E\phi_1 + 1C_2\phi_1 + 1\sigma_{xz}\phi_1 + 1\sigma_{yz}\phi_1] = \frac{1}{2}(\phi_1 + \phi_2) \\ P^{B_1} &= \frac{1}{4}[1E\phi_1 - 1C_2\phi_1 + 1\sigma_{xz}\phi_1 - 1\sigma_{yz}\phi_1] = \frac{1}{2}(\phi_1 - \phi_2) \end{aligned}$$

- P^{A_1} is a sum of the two hydrogen 1s orbitals.
- P^{B_1} is a difference of the two hydrogen 1s orbitals.
- This step allows us to construct electron density maps for the SALCs of the hydrogen's atomic orbitals.
- We can now see if the symmetry of various orbitals matches up or doesn't match up and use this information to sketch molecular orbitals.
- Oxygen atomic orbital coefficients can only be analytically derived with quantum mechanical calculations.
- How can we predict the shape of molecules from MO theory (i.e., without VSEPR theory)?
- **Walsh diagrams** help understand the molecular shapes (bond angles).
 - If H_2O is linear, it is part of the $D_{\infty h}$ point group (which can be reduced to D_{2h} for further mysterious reasons).
 - $1\sigma_g^+$: If we slowly decrease the bond angle, the hydrogen orbital overlap will increase, meaning that $1\sigma_g^+$ energy decreases.
 - $1\sigma_u^+$: If we slowly decrease the bond angle, the s and p orbitals will become less aligned, meaning that $1\sigma_u^+$ energy increases.
 - $\pi_u, 2a_1$: Energy decreases as hydrogen orbitals get closer to the appropriately signed region of the O_p orbital.
 - π_u, b_2 : Energy is constant.
 - Now all orbitals with electrons have been accounted for. Since 2 energies go down, 1 goes up, and 1 stays the same as bond angle decreases, we know that bond angle will decrease to an equilibrium.

III.11 Module 17: Constructing Molecular Orbitals (Part 4, NH₃ Molecule)

2/3: • NH₃ example:

- Point group: C_{3v} .
- Basis functions: H_{1s}, N_{2s}, N_{2p_x}, N_{2p_y}, and N_{2p_z}.
- Apply operations (pick any one operation from each class)/construct reducible representations:

$$\Gamma_H = (3, 0, 1) = A_1 + E$$

$$\Gamma_{N_{2s}} = A_1$$

$$\Gamma_{N_{2p_x}} = E$$

$$\Gamma_{N_{2p_y}} = E$$

$$\Gamma_{N_{2p_z}} = A_1$$

- N_{2p_x} and N_{2p_y} are doubly degenerate — their transformation cannot be decoupled and they transform together.
- Look at the relevant energies and plot them against each other.
- 3 A_1 type atomic orbitals form 3 a_1 type MOs.

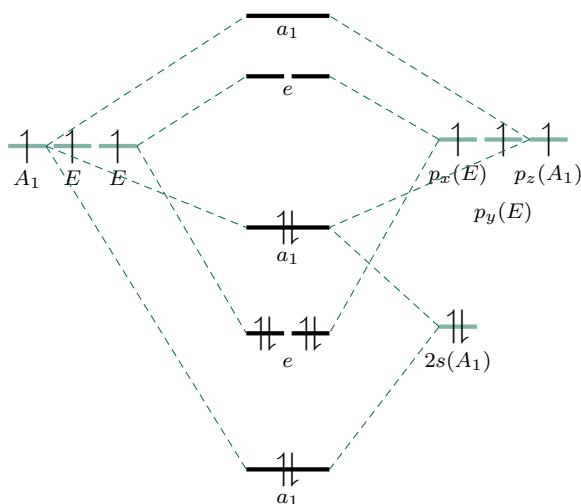


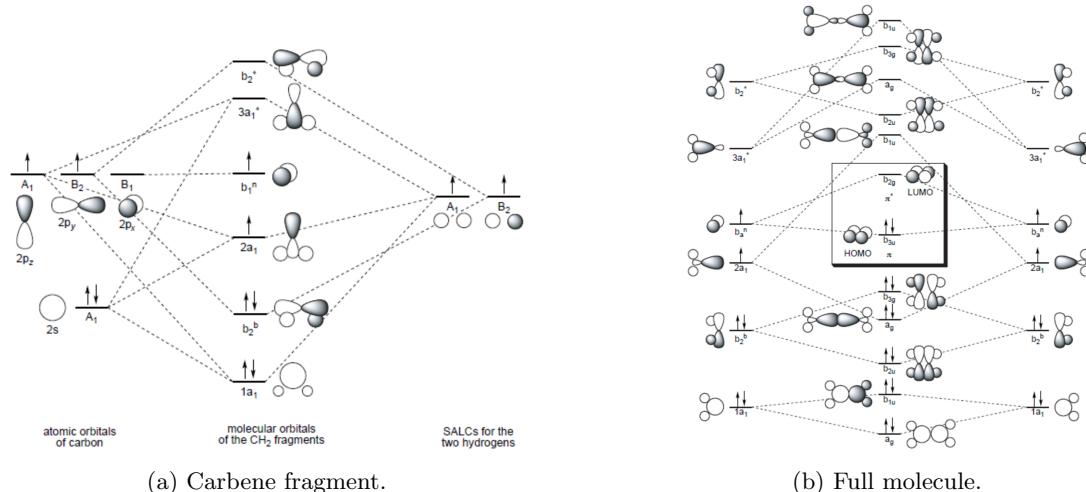
Figure III.14: NH₃ orbital diagram.

- Since 2s(A_1) and 1s(A_1) are so far away energetically, their combination will have very low energy.
- On the other hand, since the 1s(A_1) and $p_z(A_1)$ are close in energy, they have a lot of overlap.
- We can't analytically calculate orbital energies at this level. Take an educated guess on the homework and explain your reasoning. Note that here, e orbitals are more stabilizing because they are bigger and have larger overlap. Also, σ bonds are stronger than π bonds because there is a higher degree of overlap.
- To see what these orbitals look like, we need to apply the projection operator.
 - We need to construct one A_1 orbital and two E orbitals.
 - Modifying from the example from last time, we have $P^{A_1} \approx \phi_1 + \phi_2 + \phi_3$ and $P^E \approx 2\phi_1 - \phi_2 - \phi_3$.

- For the last E SALC, we apply the projection to each hydrogen atom (we can choose any basis function) giving us, in addition to the other projections, $P^E \approx 2\phi_2 - \phi_3 - \phi_1$ and $P^E \approx 2\phi_3 - \phi_1 - \phi_2$. By subtracting these two, we get $\Psi_E = \frac{1}{\sqrt{2}}(\phi_2 - \phi_3)$. Do we choose this linear combination because it's orthogonal to the other two? Because its structure is different and simple (you *can* do it with any linear combination, but it will make the analysis much more complicated if you do it with a trickier linear combination).
- We need to renormalize our SALCs so that the integral of their square over all space is 1. To normalize, we must guarantee that the sum of the squares of the coefficients is 1. For P^{A_1} for example, $1^2 + 1^2 + 1^2 = 3$, so we must scale the whole thing by $\frac{1}{\sqrt{3}}$.
- In the PES spectrum, we see three peaks, one being approximately twice as wide as the others (corresponding to the double degenerate state).
- Walsh diagrams (rely on the pictures that we draw of orbitals):
 - $1a'_1$ energy decreases, $1e'$ energy increases, a''_2 energy decreases rather dramatically.
 - If we promote electrons to the excited state, the relevant Walsh diagrams will change, and this will cause the molecule to change shape.
- Also try to sketch the MOs in the homework.

III.12 Module 18: Constructing Molecular Orbitals (Part 5, $\text{H}_2\text{C}=\text{CH}_2$ Molecule)

- $\text{H}_2\text{C}=\text{CH}_2$ example:



III.13 Module 19: Isolobal Principle

- **Isolobal** (fragment): A fragment with similar (not identical) numbers, symmetry properties, approximate energies, and shapes of the frontier orbitals and the number of electrons in them.
- This allows us to show, for example, that a carbene fragment and a d^7 metal complex have similar properties.
- Two CH_3 fragments will tend to couple into ethane. But so will two isolobal $\text{Mn}(\text{CO})_5$ fragments!

III.14 Module 20: Orbital Hybridization

- **Hybridization:** The concept of mixing atomic orbitals into new hybrid orbitals (with different energies, shapes, etc. than the component atomic orbitals) suitable for the pairing of electrons to form chemical bonds.
- Hybridization accomplishes the work of MO theory in steps: Creating new orbitals and then bonding atoms with them, as opposed to defining new orbitals and filling them.
- This was the dominant approach in chemistry before MO theory, and is still widely used in organic chemistry.
- For light elements, the energy gaps between orbitals aren't very large and this method is legitimate to some degree.
 - For carbon, it's acceptable.
 - For silicon and lead (in the same group), this is not a good approach; it will give you incorrect predictions (e.g., around bond angles and lengths).
- Steps to determine the hybridization of a bond:
 1. Assign a point group.
 2. Choose basis function (σ bonds).
 3. Apply operations.
 4. Reduce to irreducible representations.
 5. Compare symmetry of irreducible representations to central atom MOs.
- BF_3 example:
 - Point group: D_{3h} .
 - $\Gamma_\sigma = (3, 0, 1, 3, 0, 1)$.
 - $\Gamma_\sigma = A'_1 + E'$.
 - For boron, $\text{B}(s) = A'_1$, $\text{B}(p_x) = E'$, $\text{B}(p_y) = E'$, and $\text{B}(p_z) = A''_2$. Thus, since one orbital matches up on s and two match up on p (specifically, p_x, p_y), the hybrid orbitals will be sp^2 (hybrid of s, p_x, p_y).
- It is not generally justifiable to say that atomic orbitals in *individual* atoms mix before bonding.

III.15 Chapter 5: Molecular Orbitals

From Miessler et al. (2014).

- 1/29: • MO theory uses group theory to describe molecular bonding, complementing and extending Chapter 3.

- “In molecular orbital theory the symmetry properties and relative energies of atomic orbitals determine how these orbitals interact to form molecular orbitals” (Miessler et al., 2014, p. 117).
 - The molecular orbitals are filled according to the same rules discussed in Chapter 2.
 - If the total energy of the electrons in the molecular orbitals is less than that of them in the atomic orbitals, the molecule is stable relative to the separate atoms (and forms). If the total energy of the electrons in the molecular orbitals exceeds that of them in the atomic orbitals, the molecule is unstable (and does not form).
 - Homonuclear** (molecule): A molecule in which all constituent atoms have the same atomic number.
 - Heteronuclear** (molecule): A molecule that is not homonuclear, i.e., one in which at least two atoms differ in atomic number.
 - A less rigorous pictorial approach can describe bonding in many small molecules and help us to build a more rigorous one, based on symmetry and employing group theory, that will be needed to understand orbital interactions in more complex molecular structures.
 - Schrödinger equations can be written for electrons in molecules as they can for electrons in atoms. Approximate solutions can be constructed from the LCAO method.
 - In diatomic molecules for example, $\Psi = c_a\psi_a + c_b\psi_b$ where Ψ is the molecular wave function, $\psi_{a,b}$ are the atomic wave functions for atoms a and b , and $c_{a,b}$ are adjustable coefficients that quantify the contribution of each atomic orbital to the molecular orbital.
 - “As the distance between two atoms is decreased, their orbitals overlap, with significant probability for electrons from both atoms being found in the region of overlap” (Miessler et al., 2014, p. 117).
 - Electrostatic forces between nuclei and electrons in bonding molecular orbitals hold atoms together.
- 2/5:
- Precise calculations show that the coefficient of the σ^* hydrogen antibonding orbital is slightly larger than that of the σ orbital, but we typically neglect this distinction.
 - This technically implies that $\Delta E_{\sigma^*} > \Delta E_{\sigma}$, i.e., that the increase in energy from electrons in atomic orbitals moving into the σ^* antibonding molecular orbital is greater in magnitude than the decrease in energy from electrons in atomic orbitals moving into the σ bonding molecular orbital.
 - σ orbital:** A molecular orbital that is symmetric to rotation about the line connecting the nuclei.
 - Orbital asterisk:** Denotes antibonding orbitals for those molecules where bonding and antibonding orbital descriptions are unambiguous (i.e., smaller molecules).
 - When two regions of like sign overlap, the sum of the orbitals exhibits increased electron probability in the overlap region, and vice versa for regions of opposite sign.
 - When the z -axes are drawn pointing in the same direction, their difference is σ and their sum is σ^* .



(a) Same direction.

(b) Opposite directions.

Figure III.16: Choice of z -axis direction.

- Notice how in Figure III.16b, we can simply push the orbitals together (add them) to have regions of like sign overlap and form a bonding orbital.

- However, in Figure III.16a, with a more standard coordinate system, we have to flip the signs of one of the orbitals before merging them (multiply it by -1 and add it to the other orbital/subtract it from the other orbital) to create a bonding orbital. If we add them as they are, we will get an antibonding orbital.
- **π orbital:** A molecular orbital with a change in sign of the wave function under C_2 rotation about the z -axis (the bond axis).
 - π and π^* orbitals arise from the interactions between p_x and p_y atomic orbitals.
- MOs that form from p_z orbitals are σ and σ^* orbitals.
- “When orbitals overlap equally with both the same and opposite signs... the bonding and antibonding effects cancel, and no molecular orbital results” (Miessler et al., 2014, p. 121).
- d orbitals can also be involved in bonding in heavier elements such as the transition metals.
 - Two d_{z^2} orbitals participate in σ bonding.
 - Two d_{xz} or d_{yz} orbitals participate in π bonding.
 - Two $d_{x^2-y^2}$ or d_{xz} orbitals participate in **δ bonding**.
- **δ orbital:** A molecular orbital with a change in sign of the wave function under C_4 rotation about that z -axis.
 - δ and δ^* orbitals arise from the interactions of atomic orbitals meeting in parallel planes and combining side to side.
- “Sigma orbitals have no nodes that include the line connecting the nuclei, pi orbitals have one node that includes the line connecting the nuclei, and delta orbitals have two nodes that include the line connecting the nuclei” (Miessler et al., 2014, p. 121).
- Note that p_x and d_{xz} orbitals (for example) could, in theory, interact.
- Additional thoughts on Figure III.6:
 - Similarity in energy is correlated with similarity in structure.
 - When orbital energies are similar (Figure III.6a), there is a large difference between the atomic orbitals and the molecular orbitals, resulting in a great potential for stabilization through bonding.
- Molecular orbitals help us understand the structure of Li_2 , Be_2 , and other diatomic molecules that violate the octet rule.
 - They also explain experimental phenomena that clash with VSEPR theory — for example, the Lewis structure of O_2 predicts a diamagnetic molecule, but in reality, O_2 is paramagnetic with two unpaired electrons.
- See Figure III.17 for the molecular orbitals in the homonuclear diatomic molecules formed by the first 10 elements, neglecting interactions between atomic orbitals of differing energy levels.
- Bond order:

$$\text{B.O.} = \frac{1}{2}(\text{number of electrons in bonding orbitals} - \text{number of electrons in antibonding orbitals})$$
 - Generally, we need only consider valence electrons to calculate the bond order: For example, O_2 has $\text{B.O.} = 2$ whether or not you factor valence electrons into the calculation.
- Core electrons are generally assumed to reside mostly on the original atom and negligibly participate in bonding and antibonding interactions.

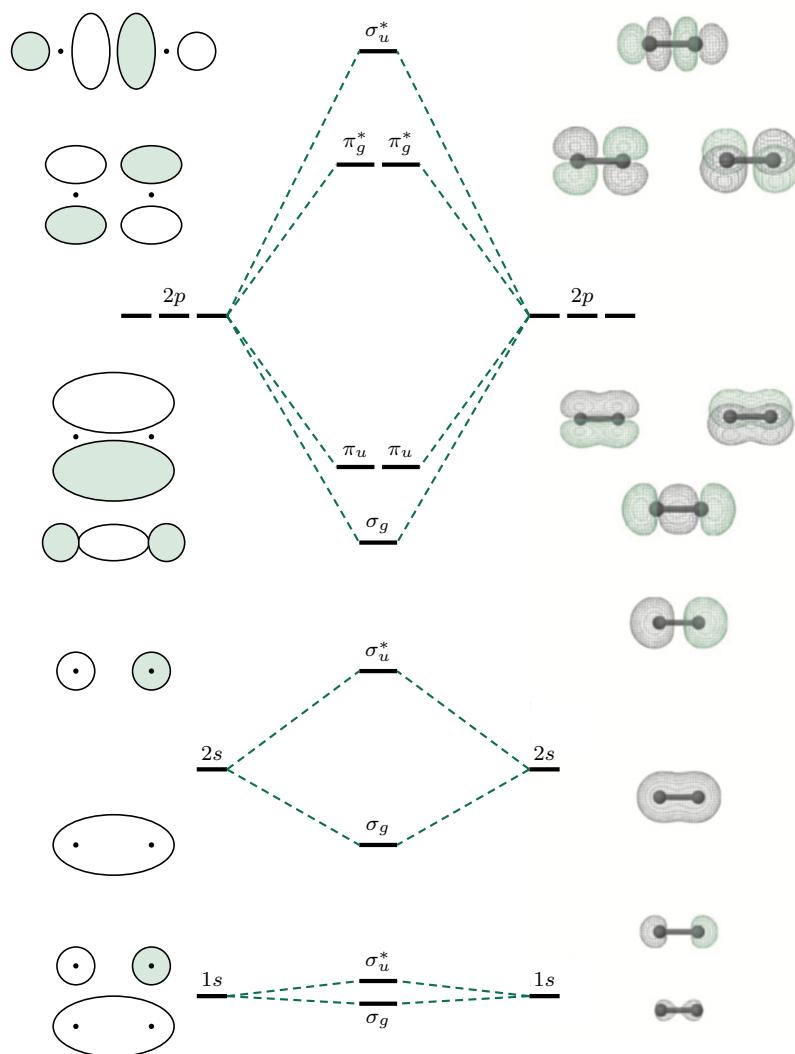


Figure III.17: Molecular orbitals for the first 10 elements.

- This is reflected by how much lower the 1s MOs are in Figure III.17 than the others, and by how slight the difference is between bonding and antibonding.
- “When two molecular orbitals of the same symmetry have similar energies, they interact to lower the energy of the lower orbital and raise the energy of the higher orbital” (Miessler et al., 2014, p. 124).
- **σ_g symmetry:** Symmetry to infinite rotation and inversion.
 - Possessed by the σ_g(2s) and σ_g(2p) orbitals in Figure III.17, for example.
- Molecular orbitals with the same symmetry and similar energies can **mix**, lowering the energy of the lower orbital and raising the energy of the higher orbital.
 - For example, the σ_g(2s) and σ_g(2p) orbitals in Figure III.17 can mix, lowering the energy of the 2s one and raising the energy of the 2p one.