

CHEM 20100 (Inorganic Chemistry I) Notes

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Topic 0

Course Prep

0.1 Chapter 1: Introduction to Inorganic Chemistry

From Miessler et al. (2014).

12/21:

- **Inorganic chemistry:** The chemistry of everything that is not organic chemistry, which is the chemistry of hydrocarbon compounds and their derivatives.
- **Organometallic chemistry:** The chemistry of compounds containing metal-carbon bonds and the catalysis of many organic reactions.
- There is also both **bioinorganic chemistry** and **environmental chemistry** (Miessler et al., 2014, p. 1), as well as **analytical chemistry**, **physical chemistry**, **petroleum chemistry**, and **polymer chemistry** (Miessler et al., 2014, p. 4).
 - Note, though, that there are no strict dividing lines between subfields of chemistry nowadays, and most professionals work in multiple fields.
- Single, double, and triple bonds (both metal-metal and metal-carbon bonds) are found in organic and inorganic chemistry.
- Quadruple bonds exist between metal atoms in some compounds.



Figure 0.1: Examples of bonding interactions.

- No such bonds exist between carbon atoms because two carbon atoms max out at a triple bond.
- Quadruple bonds possess one sigma bond, two pi bonds, and one delta (δ) bond.
- The delta bond is only possible with metal atoms because these atoms possess energetically accessible d orbitals.
- Quintuple bonds between transition metals have been reported, but scientists have not yet reached a consensus on to what extent these exist.
- Hydrogen atoms and alkyl groups can act as bridges in inorganic chemistry, excessively disobeying the octet rule (see Figure 0.2).
- **Coordination number:** The number of other atoms, molecules, or ions to which an atom is bonded.

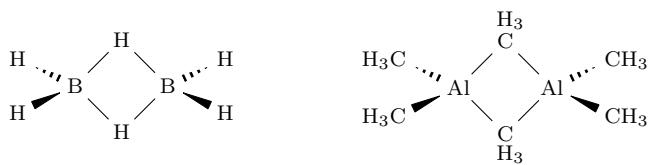


Figure 0.2: Inorganic compounds containing bridging hydrogens and alkyl groups.

- “Numerous inorganic compounds have central atoms with coordination numbers of five, six, seven, and higher” (Miessler et al., 2014, p. 2).
 - The most common coordination geometry for transition metals is octahedral.
- 4-coordinate carbon is almost always tetrahedral. 4-coordinate metals and nonmetals can be either tetrahedral or square planar.
- **Coordination complex:** A compound with a metal as the central atom or ion and some number of ligands bonded to it.
- **Ligand:** An anion or neutral molecule bonded to a central atom (frequently through N, O, or S).
- **Organometallic complex:** A coordination complex where carbon (potentially bonded to other things) is one of the ligands.



Figure 0.3: Tetrahedral geometry without a central atom.

- There are multiple kinds of tetrahedral structures. There is the standard arrangement seen in molecules such as methane, but there is also a form that lacks a central atom, as in elemental phosphorous P_4 (see Figure 0.3).
 - Other atoms such as boron and carbon also form units that surround a central cavity (e.g., icosahedral B_{12} and buckyballs C_{60}).
- Aromatic rings can bond to metals using all of their pi orbitals. This results in a metal suspended above the ring’s center.
- **Cluster compound:** A compound where “a carbon atom is at the center of a polyhedron of metal atoms” (Miessler et al., 2014, p. 3).
 - There exist examples of carbon surrounded by five, six, or more metal atoms^[1].
- Many new forms of elemental carbon have been discovered since the mid-1980s, notably including fullerenes (such as buckminsterfullerene, or buckyballs), carbon nanotubes, graphene, and polyyne wires.
- Miessler et al. (2014) give a brief history of inorganic chemistry for context.
 - Be aware of **crystal field theory** and **ligand field theory**.

^[1]This provides a challenge to theoretical inorganic chemists.

0.2 Chapter 2: Atomic Structure

From Miessler et al. (2014).

12/22:

- **Coinage metals:** Copper, silver, and gold, i.e., the transition metals in IUPAC Group 11.
- **Chalcogens:** Oxygen, sulfur, selenium, tellurium, and polonium, i.e., the nonmetals in Group 16.
- The energies of visible light emitted by the hydrogen atom are given by^[2]

$$E = R_H \left(\frac{1}{2^2} - \frac{1}{n_h^2} \right)$$

where n_h is an integer greater than 2 and R_H is the Rydberg constant for hydrogen (H).

- Note that $R_H = 1.097 \times 10^7 \text{ m}^{-1} = 2.179 \times 10^{-18} \text{ J} = 13.61 \text{ eV}$.
- This equation was first discovered by Johann Balmer in 1885.
- Infrared and ultraviolet emissions can be described by replacing 2^2 with integers n_l^2 in the above equation on the condition that $n_l < n_h$ ^[3].
- The energy of the light emitted is related to its wavelength, frequency, and wavenumber by the equations

$$E = h\nu = \frac{hc}{\lambda} = hc\bar{\nu}$$

where h is Planck's constant ($6.626 \times 10^{-34} \text{ Js}$), c is the speed of light ($2.998 \times 10^8 \text{ m s}^{-1}$), ν is the frequency of the light, λ is the wavelength of the light, and $\bar{\nu}$ is the **wavenumber** of the light.

- **Wavenumber:** The number of waves that exist between two points along the light's path a given distance apart. *Measured in waves per centimeter (SI: cm⁻¹)*.
 - Wavenumber is proportional to energy. This is why m⁻¹ or cm⁻¹ can be used as an energy unit.
- **Principal quantum number:** One of the quantities n in Balmer's equation.
- Bohr's atomic theory first explained the phenomenon of Balmer's equation, positing that "electrons may absorb light of certain specific energies and be excited to orbits of higher energy; they may also emit light of specific energies and fall to orbits of lower energy" (Miessler et al., 2014, pp. 11–12).
 - Bohr rewrote the Rydberg constant in terms of other quantities:

$$R = \frac{2\pi^2\mu Z^2 e^4}{(4\pi\epsilon_0)^2 h^2}$$

where...

- μ is the reduced mass of the electron/nucleus combination $\frac{1}{\mu} = \frac{1}{m_e} + \frac{1}{m_{\text{nucleus}}}$, where $m_e = 9.11 \times 10^{-31} \text{ kg}$ is the mass of the electron and m_{nucleus} is the mass of the nucleus.
- Z is the nuclear charge.
- $e = 1.602 \times 10^{-19} \text{ C}$ is the charge of the electron.
- h is Planck's constant.
- $4\pi\epsilon_0$ is the permittivity of a vacuum.
- Importantly, note that the Rydberg constant for hydrogen R_H is not universal, and changes for the atom at hand based on factors like the mass of the nucleus.
- Also note that some of the terms cancel or can be written in terms of others, making the above equation equivalent to the one given in Chapter 7 of Labalme (2020a).

²Refer to Labalme (2020a), specifically Figure 7.6 and the accompanying discussion.

³ n_l denotes the lower final energy level while n_h denotes the higher initial energy level.



Figure 0.4: Hydrogen atom energy levels.

- **Balmer series:** The four main electron transitions in hydrogen that release electromagnetic radiation in the visible spectrum.
- **Lyman series:** The five main electron transitions in hydrogen that release electromagnetic radiation in the ultraviolet spectrum.
- **Paschen series:** The three main electron transitions in hydrogen that release electromagnetic radiation in the infrared spectrum.
- “The inverse square dependence of energy on n results in energy levels that are far apart in energy at small n and become much closer in energy at larger n ” (Miessler et al., 2014, p. 12)

- Individual electrons can have more energy than they would possess in the infinite energy level, but at and above this point, the nucleus and electron are considered to be separate entities.
- **Heisenberg's uncertainty principle:** “There is a relationship between the inherent uncertainties in the location and momentum of an electron” (Miessler et al., 2014, p. 14).

$$\Delta x \Delta p_x \geq \frac{h}{4\pi}$$

- The above equation describes the x -component of the uncertainty, where Δx is the uncertainty in the position of the electron and Δp_x is the uncertainty in the momentum of the electron in the x -direction.
- Because of the uncertainty principle, we cannot treat electrons as particles with precisely described motion; instead, we must describe them in **orbitals**.
- **Orbital:** A region that describes the probable locations of an electron.
- **Electron density:** “The probability of finding the electron at a particular point in space” (Miessler et al., 2014, p. 14).
 - This can be calculated in principle.

- Schrödinger and Heisenberg published (in 1926 and 1927, respectively) papers on atomic wave mechanics. Although they used very different mathematical techniques, their theories can be shown to be equivalent. However, we will introduce Schrödinger's more commonly used differential equations.
- “The Schrödinger equation describes the wave properties of an electron in terms of its position, mass, total energy, and potential energy” (Miessler et al., 2014, p. 14).

$$H\Psi = E\Psi$$

- In its simplest form, it is given by the above, where H is the **Hamiltonian operator**, E is the energy of the electron, and Ψ is the **wave function**.
- **Wave function:** A function that describes an electron wave in space, i.e., an atomic orbital.
- Energy values are another way of describing the quantization introduced with the Bohr model; different orbitals, characterized by different wave functions, each have characteristic energies.
- **Operator:** “An instruction or set of instructions that states what to do with the function that follows it” (Miessler et al., 2014, p. 14).
- **Hamiltonian operator:** An operator including derivatives that transforms the wave function into a constant (the energy) times Ψ . *Also known as \hat{H} .*

$$H = \frac{-h^2}{8\pi^2 m} \left(\frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2} \right) - \frac{Ze^2}{4\pi\epsilon_0\sqrt{x^2 + y^2 + z^2}}$$

- In the form used for calculating the energy levels of one-electron systems, it is given by the above, where...
 - h is Planck's constant.
 - m is the mass of the electron.
 - e is the charge of the electron.
 - $\sqrt{x^2 + y^2 + z^2} = r$ is the distance from the nucleus.
 - Z is the charge of the nucleus.
 - $4\pi\epsilon_0$ is the permittivity of a vacuum.
- The first part describes the kinetic energy of the electron, its energy of motion.

- The second part describes the potential energy of the electron, the result of the electrostatic attraction between it and the nucleus. It is commonly designated as $V(x, y, z)$.
- “Because n varies from 1 to ∞ , and every atomic orbital is described by a unique Ψ , there is no limit to the number of solutions of the Schrödinger equation for an atom. Each Ψ describes the wave properties of a given electron in a particular orbital” (Miessler et al., 2014, p. 15).
- Electron density is proportional to Ψ^2 .
- Necessary conditions for a physically realistic solution for Ψ :
 1. Ψ must be single-valued: There cannot be two probabilities for an electron at any position in space.
 2. Ψ and its first derivatives must be continuous: The probability must be defined at all positions in space and cannot change abruptly from one point to the next.
 3. Ψ must approach zero as $r \rightarrow \infty$: For large distances from the nucleus, the probability must grow smaller and smaller (the atom must be finite).
 4. The integral

$$\int_{\text{all space}} \Psi_A \Psi_A^* d\tau = 1$$

The total probability of an electron being somewhere in space must be 1. Applying this stipulation is called **normalizing** the wave function^[4].

5. The integral

$$\int_{\text{all space}} \Psi_A \Psi_B^* d\tau = 0$$

Ψ_A and Ψ_B are different orbitals within the same atom, and this stipulation reflects the fact that all orbitals in the same atom must be **orthogonal** to each other.

12/23:

- The Particle in a Box.
 - Imagine a one-dimensional “box” between $x = 0$ and $x = a$ with potential energy everywhere zero inside the “box” and everywhere infinite outside the box.
 - These energy constraints mean that the particle is completely trapped within the box (it would take an infinite amount of energy to leave), but also that no forces act on it within the box.
 - It follows that the wave equation for locations within the box is

$$\frac{-\hbar^2}{8\pi^2 m} \left(\frac{\partial^2 \Psi(x)}{\partial x^2} \right) = E\Psi(x)$$

- Since sine and cosine functions have properties associated with waves, we propose that a general solution describing possible waves in the box is

$$\Psi(x) = A \sin rx + B \cos sx$$

where A, B, r, s are constants.

- As seen in Problem 2.8a, substituting the above general solution into the wave equation allows us to solve for

$$r = s = \sqrt{2mE} \frac{2\pi}{\hbar}$$

⁴ Ψ_A^* denotes the complex conjugate of Ψ_A . This is necessary because wave functions may have imaginary values. However, in many cases, the wave functions are real and the integrand reduces to Ψ_A^2 .

- Because Ψ must be continuous and equal 0 for $x < 0$ and $x > a$, Ψ must approach 0 as $x \rightarrow 0$ and $x \rightarrow a$. Since $\cos(s \cdot 0) = 1$, $\Psi(0)$ can only equal 0 if $B = 0$. Therefore, the wave function reduces to

$$\Psi(x) = A \sin rx$$

- We must also have $\Psi(a) = 0$, which is only possible if rx is an integer multiple of π . This means that

$$r = \frac{\pm n\pi}{a}$$

- Combining the two expressions for r , we can solve for E :

$$E = \frac{n^2 h^2}{8 m a^2}$$

- “These are the energy levels predicted by the particle-in-a-box model for any particle in a one dimensional box of length a . The energy levels are quantized according to quantum numbers $n = 1, 2, 3, \dots$ ” (Miessler et al., 2014, p. 17).
 - As seen in Problem 2.8d, substituting $r = n\pi/a$ into $\Psi(x)$ and applying the normalizing requirement allows us to solve for the total solution

$$\Psi(x) = \sqrt{\frac{2}{a}} \sin \frac{n\pi x}{a}$$

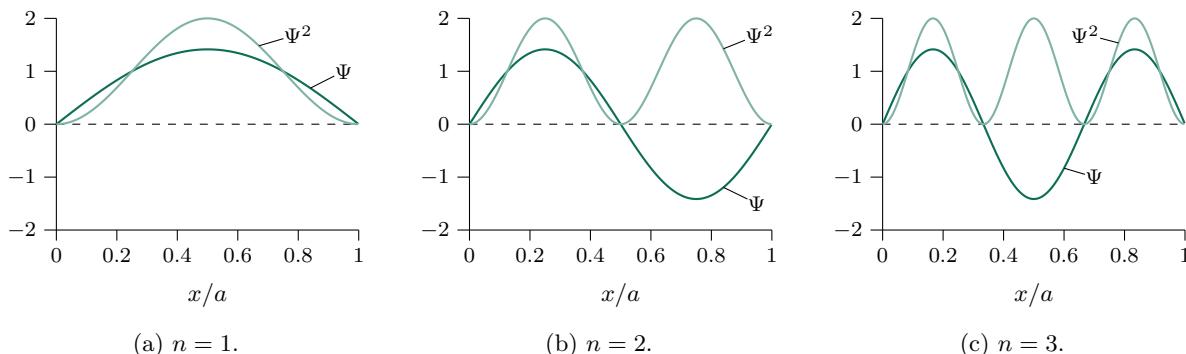


Figure 0.5: Particle in a box: Wave functions and their squares at different energy levels.

- The predicted probability densities for an electron at different energy levels (see Figure 0.5) showcase a difference between classical mechanics' and quantum mechanics' predictions about the behavior of this particle: Classical predicts an equal probability at any point in the box, but the (quantum) wave nature of the particle predicts varied probabilities in different locations.
 - “The greater the square of the electron amplitude, the greater the probability of the electron being located at the specified coordinate when at the quantized energy defined by Ψ ” (Miessler et al., 2014, p. 17).
 - Atomic orbitals, mathematically, are discrete solutions of the three-dimensional Schrödinger equations.
 - These orbital equations include four quantum numbers (see Table 0.1).
 - Orbitals with l values $0, 1, 2, 3, 4, \dots$ are known by the labels s, p, d, f, g, \dots (continuing alphabetically), respectively, derived from early terms for different families of spectroscopic lines.

Symbol	Name	Values	Role
n	Principal	1, 2, 3, ...	Determines the major part of the energy.
l	Angular momentum ^[5]	0, 1, 2, ..., $n - 1$	Describes angular dependence and contributes to the energy.
m_l	Magnetic	$0, \pm 1, \pm 2, \dots, \pm l$	Describes orientation in space (angular momentum in the z -direction).
m_s	Spin	$\pm \frac{1}{2}$	Describes orientation of the electron spin (magnetic moment) in space.

Table 0.1: Quantum numbers and their properties.

- n is the primary quantum number affecting the overall energy.
- l determines the type of shape. See Table 0.2.
- m_l determines the “orientation of the angular momentum vector in a magnetic field, or the position of the orbital in space” (Miessler et al., 2014, p. 18). See Table 0.2.
- m_s determines the electron spin^[6] in the orbital, or “the orientation of the electron’s magnetic moment in a magnetic field, either in the direction of the field ($+\frac{1}{2}$) or opposed to it ($-\frac{1}{2}$)” (Miessler et al., 2014, p. 18).
- n , l , and m_l specify an orbital; m_s specifies one of the two electrons in the orbital.
- A note on $i = \sqrt{-1}$ in Table 0.2:
 - Any linear combination of solutions to the wave equation is another solution.
 - Because it is more convenient to work with real functions as opposed to complex functions, we make use of the above fact to decomplexify the solutions, as in the following examples.

$$\begin{aligned}
 \Psi_{2p_x} &= \frac{1}{\sqrt{2}}(\Psi_{+1} + \Psi_{-1}) \\
 &= \frac{1}{\sqrt{2}} \left(\frac{1}{\sqrt{2\pi}} e^{i\phi} \cdot \frac{\sqrt{3}}{2} \sin \theta \cdot [R(r)] + \frac{1}{\sqrt{2\pi}} e^{-i\phi} \cdot \frac{\sqrt{3}}{2} \sin \theta \cdot [R(r)] \right) \\
 &= \left(\frac{1}{\sqrt{\pi}} \frac{\sqrt{3}}{2} [R(r)] \sin \theta \right) \frac{e^{i\phi} + e^{-i\phi}}{2} \\
 &= \frac{1}{2} \sqrt{\frac{3}{\pi}} [R(r)] \sin \theta \cos \phi
 \end{aligned}$$

$$\begin{aligned}
 \Psi_{2p_y} &= \frac{-i}{\sqrt{2}}(\Psi_{+1} - \Psi_{-1})^{\text{[7]}} \\
 &= \left(\frac{1}{\sqrt{\pi}} \frac{\sqrt{3}}{2} [R(r)] \sin \theta \right) \frac{e^{i\phi} - e^{-i\phi}}{2i} \\
 &= \frac{1}{2} \sqrt{\frac{3}{\pi}} [R(r)] \sin \theta \sin \phi
 \end{aligned}$$

⁵Also known as **azimuthal** (quantum number).

⁶Note that the spin of an electron is purely quantum mechanical, and should not be related back to classical mechanics.

⁷Errata: Miessler et al. (2014) incorrectly normalizes this expression with $i/\sqrt{2}$.

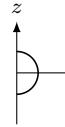
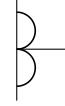
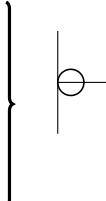
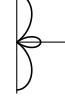
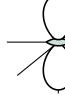
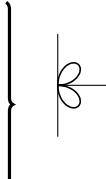
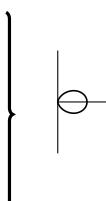
Angular factors			Real Wave Functions					
Related to Angular Momentum			Functions of θ	In Polar Coordinates	In Cartesian Coordinates	Shapes	Label	
l	m_l	Φ	Θ	$\Theta\Phi(\theta, \phi)$	$\Theta\Phi(x, y, z)$			
0(s)	0	$\frac{1}{\sqrt{2\pi}}$	$\frac{1}{\sqrt{2}}$		$\frac{1}{2\sqrt{\pi}}$	$\frac{1}{2\sqrt{\pi}}$		s
1(p)	0	$\frac{1}{\sqrt{2\pi}}$	$\frac{\sqrt{6}}{2} \cos \theta$		$\frac{1}{2}\sqrt{\frac{3}{\pi}} \cos \theta$	$\frac{1}{2}\sqrt{\frac{3}{\pi}} \frac{z}{r}$		p_z
$+1$	$\frac{1}{\sqrt{2\pi}} e^{i\phi}$	$\frac{\sqrt{3}}{2} \sin \theta$		$\frac{1}{2}\sqrt{\frac{3}{\pi}} \sin \theta \cos \phi$	$\frac{1}{2}\sqrt{\frac{3}{\pi}} \frac{x}{r}$		p_x	
				$\frac{1}{2}\sqrt{\frac{3}{\pi}} \sin \theta \sin \phi$	$\frac{1}{2}\sqrt{\frac{3}{\pi}} \frac{y}{r}$		p_y	
2(d)	0	$\frac{1}{\sqrt{2\pi}}$	$\frac{1}{2}\sqrt{\frac{5}{2}}(3\cos^2 \theta - 1)$		$\frac{1}{4}\sqrt{\frac{5}{\pi}}(3\cos^2 \theta - 1)$	$\frac{1}{4}\sqrt{\frac{5}{\pi}} \frac{2z^2 - x^2 - y^2}{r^2}$		d_{z^2}
$+1$	$\frac{1}{\sqrt{2\pi}} e^{i\phi}$	$\frac{\sqrt{15}}{2} \cos \theta \sin \theta$		$\frac{1}{2}\sqrt{\frac{15}{\pi}} \cos \theta \sin \theta \cos \phi$	$\frac{1}{2}\sqrt{\frac{15}{\pi}} \frac{xz}{r^2}$		d_{xz}	
				$\frac{1}{2}\sqrt{\frac{15}{\pi}} \cos \theta \sin \theta \sin \phi$	$\frac{1}{2}\sqrt{\frac{15}{\pi}} \frac{yz}{r^2}$		d_{yz}	
$+2$	$\frac{1}{\sqrt{2\pi}} e^{2i\phi}$	$\frac{\sqrt{15}}{4} \sin^2 \theta$		$\frac{1}{4}\sqrt{\frac{15}{\pi}} \sin^2 \theta \cos 2\phi$	$\frac{1}{4}\sqrt{\frac{15}{\pi}} \frac{x^2 - y^2}{r^2}$		$d_{x^2 - y^2}$	
				$\frac{1}{4}\sqrt{\frac{15}{\pi}} \sin^2 \theta \sin 2\phi$	$\frac{1}{4}\sqrt{\frac{15}{\pi}} \frac{xy}{r^2}$		d_{xy}	

Table 0.2: Hydrogen atom wave functions: Angular functions.

- Note that d_{z^2} actually uses the function $2z^2 - x^2 - y^2$, so we should write $d_{2z^2 - x^2 - y^2}$ but we shorthand for convenience.
- Since the functions in the polar and Cartesian coordinates columns of Table 0.2 are real, $\Psi = \Psi^*$ and $\Psi\Psi^* = \Psi^2$.
- 12/25: • Ψ may be expressed in terms of Cartesian coordinates (x, y, z) or in terms of spherical coordinates (r, θ, ϕ) .
 - Using spherical coordinates comes with the advantage that r is the distance from the nucleus (see Figure 0.6).

12/23:

Orbital	n	l	Radial Functions $R(r)$, with $\sigma = Zr/a_0$
$1s$	1	0	$R_{1s} = 2 \left[\frac{Z}{a_0} \right]^{3/2} e^{-\sigma}$
$2s$	2	0	$R_{2s} = \left[\frac{Z}{2a_0} \right]^{3/2} (2 - \sigma) e^{-\sigma/2}$ ^[8]
$2p$		1	$R_{2p} = \frac{1}{\sqrt{3}} \left[\frac{Z}{2a_0} \right]^{3/2} \sigma e^{-\sigma/2}$
$3s$	3	0	$R_{3s} = \frac{2}{27} \left[\frac{Z}{3a_0} \right]^{3/2} (27 - 18\sigma + 2\sigma^2) e^{-\sigma/3}$
$3p$		1	$R_{3p} = \frac{1}{81\sqrt{3}} \left[\frac{2Z}{a_0} \right]^{3/2} (6 - \sigma) \sigma e^{-\sigma/3}$
$3d$		2	$R_{3d} = \frac{1}{81\sqrt{15}} \left[\frac{2Z}{a_0} \right]^{3/2} \sigma^2 e^{-\sigma/3}$

Table 0.3: Hydrogen atom wave functions: Radial functions.

12/25:



Figure 0.6: Spherical coordinates.

- Know the following volume element conversions^[9].

$$dV = dx dy dz = r^2 \sin \theta d\theta d\phi dr$$

- The volume of the thin shell between r and $r + dr$ is as follows^[10], and is useful for describing the electron density as a function of distance from the nucleus.

$$\begin{aligned} \int_0^{2\pi} \int_0^\pi 1 r^2 \sin \theta d\theta d\phi dr &= \int_0^{2\pi} 2r^2 d\phi dr \\ &= 4\pi r^2 dr \end{aligned}$$

⁸Errata: This differs from the corresponding equation in Miessler et al. (2014) because the textbook is wrong.

⁹Refer to Labalme (2020b), specifically Figure 16.9 and the accompanying discussion.

¹⁰Errata: Miessler et al. (2014) flips the bounds over which θ and ϕ should be evaluated, but I use the right ones.

- Ψ can be factored into a **radial component** R and two **angular components** Θ and Φ , which are sometimes combined into a single component Y :

$$\Psi(r, \theta, \phi) = R(r)\Theta(\theta)\Phi(\phi) = R(r)Y(\theta, \phi)$$

- Refer to Table 0.2 for examples of Θ and Φ , and Table 0.3 for examples of R .
- Note that in Table 0.2, the shaded orbital lobes correspond to where the wave function is negative. Distinguishing regions of opposite signs is useful for bonding purposes.
- The radial component can also be manipulated to give the **radial probability function** $4\pi r^2 R^2$.
- R is the radial function, so R^2 is the probability function (think $\Psi = RY$, so $\Psi^2 = R^2 Y^2$).

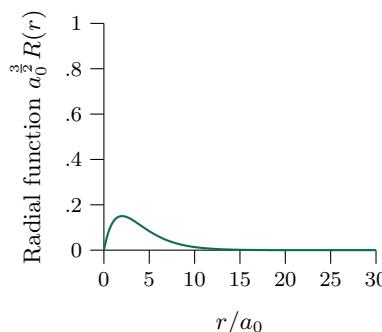
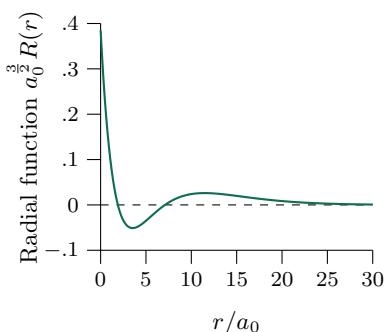
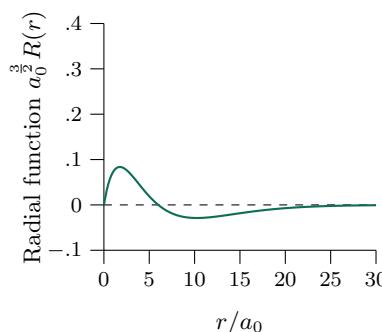
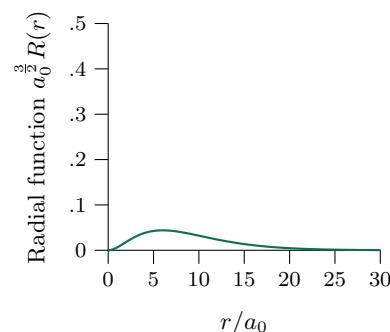
(a) $1s$.(b) $2s$.(c) $2p$.(d) $3s$.(e) $3p$.(f) $3d$.

Figure 0.7: Radial wave functions.



Figure 0.8: Radial probability functions.

- **Radial component:** The part of Ψ that describes “electron density at different distances from the nucleus” (Miessler et al., 2014, p. 20).
 - Determined by quantum numbers n and l .
 - Plotted for the $n = 1, 2, 3$ orbitals in Figure 0.7.
- **Angular components:** The parts of Ψ that describe “the shape of the orbital and its orientation in space” (Miessler et al., 2014, p. 20).
 - Determined by quantum number l and m_l .
- **Radial probability function:** The function describing “the probability of finding the electron at a given distance from the nucleus, summed over all angles, with the $4\pi r^2$ factor the result of integrating over all angles” (Miessler et al., 2014, p. 21).

- Plotted for the $n = 1, 2, 3$ orbitals in Figure 0.8.
- **Bohr radius:** “The value of r at the maximum of Ψ^2 for a hydrogen $1s$ orbital (the most probable distance from the hydrogen nucleus for the $1s$ electron), and it is also the radius of the $n = 1$ orbit according to the Bohr model” (Miessler et al., 2014, p. 23).
 - See Figure 0.8a.
 - Its value is $a_0 = 52.9 \text{ pm}$.
 - It is a common unit in quantum mechanics.
 - It is used to scale the functions in Figures 0.7 and 0.8 to give reasonable axis units^[11].
- Figure 0.8 shows that the electron density falls off rapidly after the absolute maximum in every case.
 - However, it falls off more quickly for lower energy levels.
- The electron density at the nucleus is always zero.
 - Mathematically, this is because $4\pi r^2 R^2$ naturally equals zero when $r = 0$.
- “Because chemical reactions depend on the shape and extent of orbitals at large distances from the nucleus, the radial probability functions help show which orbitals are most likely to be involved in reactions” (Miessler et al., 2014, p. 23).

12/26:

- **Nodal surface:** A surface within an orbital having zero electron density.
 - The wave function is zero at these because it is changing sign.
 - Knowing where the wave function is positive and negative can be useful when working with molecular orbitals.
 - To find nodal surfaces, we must find places where

$$0 = \Psi^2 = \Psi = R(r)Y(\theta, \phi)$$

or where

$$0 = R(r) \quad 0 = Y(\theta, \phi)$$

- “The total number of nodes in any orbital is $n - 1$ if the conical nodes of some d and f orbitals count as two nodes” (Miessler et al., 2014, p. 23).
 - The conical node of the d_{z^2} orbital (see Table 0.2) is mathematically one surface, but we must count it as two (perhaps think of it as a top and bottom section) to fit the pattern.
- **Angular node:** A node that results when $Y = 0$.
 - There are l angular nodes in any orbital.
- **Radial node:** A node that results when $R = 0$. *Also known as spherical nodes.*
 - There are $n - l - 1$ radial nodes in any orbital.
- Note that on the basis of relativity, electron density at nodes may not be zero but may in fact be a very small finite value.
 - Alternatively, if we think of an electron wave like a wave in a violin string, which also has places of zero vibration where the string still exists, nodes may just be places where electron waves have zero amplitude, not places where they don’t exist.

¹¹Errata?: Should the y -axis in Figure 0.8 be $a_0^3 r^2 R^2$?

- A further possible explanation is that electrons can ‘teleport’ between two places without ever passing in between.
- We now look at a few examples of finding nodes in orbitals.
- Nodal structure of p_z :
 - From Table 0.2, the angular factor of this orbital in Cartesian coordinates is
$$Y = \frac{1}{2} \sqrt{\frac{3}{\pi}} \frac{z}{r}$$
 - When we set this equal to zero, we find that the above is only equal to zero when $z = 0$.
 - Indeed, $z = 0$, i.e., the xy -plane, is an angular node in the p_z orbital, as we can see in Table 0.2.
 - Additionally, we can tell from the above equation that the wave function is positive when $z > 0$ and negative when $z < 0$; this is also reflected in Table 0.2.
- Nodal structure of $d_{x^2-y^2}$:
 - The angular factor:
$$Y = \frac{1}{4} \sqrt{\frac{15}{\pi}} \frac{x^2 - y^2}{r^2}$$
 - This is only equal to zero when
$$\begin{aligned} 0 &= x^2 - y^2 \\ y^2 &= x^2 \\ y &= \pm x \end{aligned}$$
 - This translates to Table 0.2, where we can see that the planes $x = y$ and $x = -y$, i.e., those that contain the z -axis and make 45° angles with the x - and y -axes, are nodes.
 - Additionally, the function is positive where $|x| > |y|$ and negative where $|x| < |y|^{[12]}$, and we can see this in Table 0.2.
- Be aware of lines/surfaces of constant electron density.
- When filling orbitals in polyelectronic atoms, “we start with the lowest n , l , and m_l values (1, 0, and 0, respectively) and either of the m_s values (we will arbitrarily use $+\frac{1}{2}$ first)” (Miessler et al., 2014, p. 26). We then utilize the three following rules.
 - See them used to fill a p orbital in Table 0.4.
- **Aufbau principle:** The buildup of electrons in atoms results from continually increasing the quantum numbers. *Etymology aufbau* from German “building up.”
 - “Electrons are placed in orbitals to give the lowest total electronic energy to the atom. This means that the lowest values of n and l are filled first. Because the orbitals within each subshell (p , d , etc.) have the same energy, the orders for values of m_l and m_s are indeterminate” (Miessler et al., 2014, p. 26).
- **Pauli exclusion principle:** Each electron in an atom must have a unique set of quantum numbers.
 - Note that this is experimentally derived, and does not follow from the Schrödinger equation.
- **Hund’s rule** (of maximum multiplicity): Electrons must be placed in orbitals to give the maximum total spin (the maximum number of parallel spins).
 - This is a consequence of the Aufbau principle, as two electrons in the same orbital have higher energy than two in different orbitals due to electrostatic repulsions.

¹²Errata: Miessler et al. (2014) incorrectly use $x > y$ and $x < y$.

Number of electrons	Arrangement	Unpaired e ⁻	Multiplicity
1	$\begin{array}{c} \uparrow \\ \hline \end{array}$ ____	1	2
2	$\begin{array}{cc} \uparrow & \uparrow \\ \hline & \hline \end{array}$ ____	2	3
3	$\begin{array}{ccc} \uparrow & \uparrow & \uparrow \\ \hline & \hline & \hline \end{array}$	3	4
4	$\begin{array}{ccc} \uparrow \downarrow & \uparrow & \uparrow \\ \hline & \hline & \hline \end{array}$	2	3
5	$\begin{array}{ccc} \uparrow \downarrow & \uparrow \downarrow & \uparrow \\ \hline & \hline & \hline \end{array}$	1	2
6	$\begin{array}{ccc} \uparrow \downarrow & \uparrow \downarrow & \uparrow \downarrow \\ \hline & \hline & \hline \end{array}$	0	1

Table 0.4: Hund's rule and multiplicity.

- **Spin multiplicity:** The number of unpaired electrons plus 1.



Figure 0.9: Coulombic energy of repulsion and exchange energy.

- **Coulombic energy of repulsion:** The potential energy of two negatively charged electrons that occupy the same orbital (as opposed to separate orbitals). *Denoted by Π_c .*
 - Note that all polyelectronic species are subject to *some* Coulombic repulsions, but the contribution is significantly higher for paired electrons.
- **Exchange energy:** The potential energy of two electrons that have opposite spins (as opposed to the same spin). *Denoted by Π_e .*
 - This arises from purely quantum mechanical considerations.
 - Basically, there are three possible ways to arrange two electrons in degenerate orbitals (see Figure 0.9). The exchange energy between two of the states reflects the fact that electrons with opposite spins are distinguishable, but electrons with identical spins can be exchanged (can switch places) and no observer would be the wiser. This event, when it happens, is called **one exchange of parallel electrons**; the fact that it can happen lowers the energy of the system with electrons of identical spin simply by virtue of creating more possible states that each electron can occupy.

12/27:

- Energies with oxygen:

- Three possible configurations: $\begin{array}{cccc} \uparrow \downarrow & \uparrow \downarrow & _ & _ \end{array}$, $\begin{array}{cccc} \uparrow \downarrow & \uparrow & _ & \downarrow \end{array}$, and $\begin{array}{cccc} \uparrow \downarrow & \uparrow & _ & \uparrow \end{array}$.
- In the first one, we have two pairs of electrons (so $2\Pi_c$) and two possible exchanges (one for each pair of electrons with like spin, so $2\Pi_e$).

- In the second one, we have one pair of electrons (so $1\Pi_c$) and two possible exchanges (one for each pair of electrons with like spin, so $2\Pi_e$).
 - In the third one, we have one pair of electrons (so $1\Pi_c$) and three possible exchanges (1-2, 1-3, and 2-3, so $3\Pi_e$).
 - Thus, the energies from greatest to least are 1, 2, 3, and indeed the three rules suggest that the third one is how we should put four electrons into a p orbital.
- **Degenerate** (orbitals): Orbitals with the same energy.
 - Energy minimization is the driving force determining the ground state, and this usually means that lower subshells are filled before higher ones.
 - However, in some transition elements, subshells are so close in energy ($4s$ and $3d$, for example) that the sum of the Coulombic and exchange terms can exceed the energy difference between subshells.
 - **Klechkowsky's rule:** “The order of filling of the orbitals proceeds from the lowest available value for the sum $n + l$. When two combinations have the same value, the one with the smaller value of n is filled first” (Miessler et al., 2014, p. 29).
 - The periodic table can also be used to predict orbital filling.
 - Both Klechkowsky's rule and the periodic table method are imperfect, but they work for the majority of atoms and provide starting points for the others.
 - Miessler et al. (2014) includes a table listing ground state electron configurations for every atom, including variant ones.
 - **Shielding:** “Each electron acts as a shield for electrons farther from the nucleus, reducing the attraction between the nucleus and the more distant electrons” (Miessler et al., 2014, p. 30).
 - As Z increases, orbital energies decrease (electrons are pulled closer to the nucleus).
 - When ranking higher energy orbitals, l must be considered in addition to n (e.g., to know that $4s$ fills before $3d$)^[13].
 - **Slater's rules:** Four rules that approximately determine the magnitude of the effective nuclear charge $Z^* = Z - S$, where Z is the nuclear charge and S is the shielding constant.
 1. Write the atom's electronic structure in order of increasing quantum numbers n and l , grouped as follows:

(1s) (2s, 2p) (3s, 3p) (3d) (4s, 4p) (4d) (4f) (5s, 5p) (5d) (and so on)
 2. Electrons in groups to the right in this list do not shield electrons to their left.
 3. We may now determine S . For ns and np valence electrons:
 - a. Each electron in the same group contributes 0.35 to the value of S for each other electron in the group, with one exception: A $1s$ electron contributes 0.30 to S for another $1s$ electron.
 - b. Each electron in $n - 1$ groups contributes 0.85 to S .
 - c. Each electron in $n - 2$ or lower groups contributes 1.00 to S .
 4. For nd and nf valence electrons:
 - a. Each electron in the same group contributes 0.35 to the value of S for each other electron in the group^[14].
 - b. Each electron in groups to the left contributes 1.00 to S .

¹³See Labalme (2020a), specifically Figure 7.12.

¹⁴This is the same as Rule 3a.

- Slater's rules for nickel:
 - Rule 1: The electron configuration is written $(1s^2)(2s^2, 2p^6)(3s^2, 3p^6)(3d^8)(4s^2)$.
 - For a $3d$ electron:
 - Rule 4a: Each other electron in the $(3d^8)$ group contributes 0.35 to S , so the total contribution is $7 \times 0.35 = 2.45$.
 - Rule 4b: Each electron in groups to the left of $(3d^8)$ contributes 1.00 to S , so the total contribution is $18 \times 1.00 = 18.00$.
 - Thus, $S = 2.45 + 18.00 = 20.45$, so $Z^* = 28 - 20.45 = 7.55$.
 - For a $4s$ electron:
 - Rule 3a: The other electron in the $(4s^2)$ group contributes 0.35 to S .
 - Rule 3b: Each electron in the $n - 1$ groups $(3s^2, 3p^6)(3d^8)$ contributes 0.85 to S , so the total contribution is $16 \times 0.85 = 13.60$.
 - Rule 3c: Each other electron to the left contributes 1.00 to S , so the total contribution is $10 \times 1.00 = 10.00$.
 - Thus, $S = 0.35 + 13.60 + 10.00 = 23.95$, so $Z^* = 28 - 23.95 = 4.05$.
 - Note that the fact that the $4s$ electrons are held less tightly than the $3d$ electrons is reflected in the fact that when $\text{Ni} \longrightarrow \text{Ni}^{2+} + 2e^-$, the two electrons removed are the $4s$ electrons.
- Slater's rules are justified by electron probability curves (see Figure 0.8) and experimental data.
- Rich provided another explanation of orbital filling (refer to Figure 0.10 throughout the following discussion).

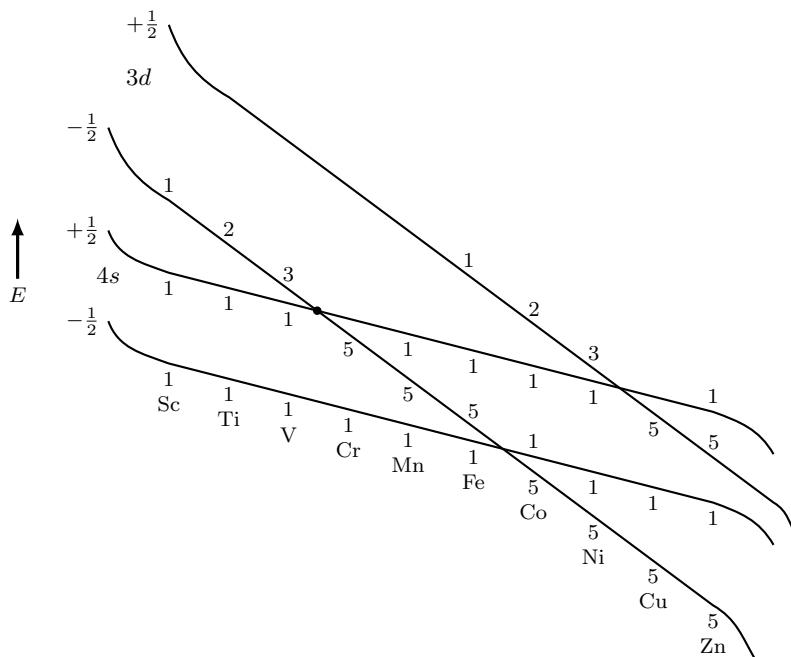
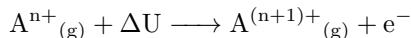


Figure 0.10: Schematic energy levels for transition elements.

- Each orbital can be conceived as two half-orbitals, separated energetically by a variable electron pairing energy of the form $a\Pi_c + b\Pi_e$ where $a, b \in \mathbb{Z}^+$.
 - Take Fe, for example. If one of the $3d$ electrons with $m_s = -\frac{1}{2}$ were changed to have $m_s = +\frac{1}{2}$ (i.e., paired with another electron), we would expect the energy of the system to increase by $\Pi_c + 4\Pi_e$.

- Each of these half orbitals naturally decreases in energy (is pulled closer to the nucleus) as Z increases.
 - Since orbitals in lower energy levels have shorter most probable distances to the nucleus, they are stabilized more as Z increases (this is why the slope of the $3d$ half orbitals is steeper).
 - Electrons are placed into the lowest energy positions possible.
 - Thus, the $4s$ orbital usually fills up first, but there are exceptions since, for example, the $3d$ half orbital with $m_s = -\frac{1}{2}$ crosses over the $4s$ half orbital with $m_s = +\frac{1}{2}$ between V and Cr (at the dot) as opposed to somewhere after Cr.
 - Note that as electrons are sequentially added, Figure 0.10 shows them generally having the same spin, in accordance with Hund's rule.
 - For ions, the crossover points shift left.
 - This is because the removal of the electron causes Z^* to increase dramatically for all electrons, but more for $(n-1)d$ orbitals than ns orbitals.
 - “This approach to electron configurations of transition metals does not depend on the stability of half-filled shells or other additional factors” (Miessler et al., 2014, p. 36).
 - Introductory chem: Electrons in the highest energy level are always removed first when ionizing transition metals. This chem: “Regardless of which electron is lost to form a transition metal ion, the lowest energy electron configuration of the resulting ion will always exhibit the vacancy in the ns orbital” (Miessler et al., 2014, p. 36).
 - Similar but more complex diagrams can treat other situations in higher energy levels and subshells.
- Since the periodic table arranges^[15] atoms on the basis of similar electronic configurations, a given atom's position provides information about its properties.
 - **Ionization energy:** The energy required to remove an electron from a gaseous atom or ion. *Also known as ionization potential.*



- When $n = 0$, ΔU is the first ionization energy IE_1 . When $n = 1$, ΔU is the second ionization energy IE_2 . The pattern continues.
 - Ionization energy trends^[16]:
 - Increases across a period (the major change).
 - Decreases down a group (a minor change; occurs because an increase in quantum number is associated with a much higher energy, and electrons with more energy need less to make it to the infinite energy level).
 - The trend breaks at boron and oxygen due to core shielding and Π_c , respectively.
 - Such trends are most pronounced in the main group elements; in the transition metals, the lanthanides, and the actinides, “the effects of shielding and increasing nuclear charge [are] more nearly in balance” (Miessler et al., 2014, p. 37).
- **Electron affinity:** The energy required to remove an electron from a negative ion. *Also known as zeroth ionization energy.*



- Electron affinity is often thought of as the energy released by adding an electron to a neutral atom, but taking this perspective facilitates comparisons with ionization energy.
- This reaction is endothermic, except for the noble gases and the alkaline earth elements.
- Trends are similar to those for ionization energy, but “for one larger Z value (one more electron for each species^[17]) and with much smaller absolute numbers” (Miessler et al., 2014, p. 37).
 - The magnitude decreases because outer electrons in anions are better shielded.

¹⁵Errata: Miessler et al. (2014) misspells arrangement as “arrangment” (p. 36).

¹⁶Refer to Labalme (2020a), specifically Figure 7.19 as well as Exercise on Z_{eff} .

¹⁷Think isoelectricity (see Figure 0.11).



Figure 0.11: First and second ionization energies and electron affinities.

- Atomic radius decreases across periods and increases down groups.
- **Nonpolar covalent radius:** The radius of an atom as determined from bond lengths in nonpolar molecules.
- **Van der Waals radius:** The radius of an atom as determined from collisions with other atoms.
 - Because compounds vary so much, it is hard to use either of these or any method to definitively determine atomic radius.
- **Pauling's approach:** A method for determining ionic radii based on the assumption that the ratio of the radii of isoelectronic atoms equals the ratio of their effective nuclear charges.
 - This is not entirely accurate. Modern cation **crystal radii** differ by +14 pm, and such anion radii differ by -14 pm.
- Factors that influence ionic size: “the coordination number of the ion, the covalent character of the bonding, distortions of regular crystal geometries, and delocalization of electrons (metallic or semiconducting character...)” (Miessler et al., 2014, p. 40).

Topic I

Review of VSEPR Theory

I.1 Module 1: Course Logistics and History

1/11:

- Homework questions will be similar to exam questions, so although you probably *can* find answers online, you shouldn't.
- Submit Psets to chem201hw@gmail.com.
- Watch modules before office hours and bring questions.
- If you have a question outside of office hours, post it on Slack.
- It's a difficult class, but he is open to and welcomes our feedback (via Slack, again).
- You do need to read from Miessler et al. (2014), too; his class is not a replacement for this textbook.
 - He is a big fan of Cotton (1990).
 - There is an extra, new textbook to look for!
- Convince yourself not to be afraid of time-independent quantum mechanics (we won't go too deep, but know wave functions and the like).
- Exams will probably be open book/open note.
- Inorganic chemistry contains too much information to rationalize empirically, so we need a system (the development of this system will be the focus of this course).
- Reviews history of chemistry from Miessler et al. (2014) Chapter 1.
- What is Nickel's electron configuration?
 - When Nickel is a free atom, the $[Ar]4s^23d^8$ electron configuration is the lowest energy.
 - When Nickel is chemically bound, the $[Ar]3d^{10}$ electron configuration is the lowest energy because it is energetically unfavorable to have a large 4s orbital pushing the bounds of the atom.
 - What is a **term symbol**?
- Homework: Refresh Chapter 2 in Miessler et al. (2014).
- **Covalent bond:** The sharing of pairs of electrons...?
- G. N. Lewis predicts in 1916 (before Rutherford) that the atom has a positive **kernel** surrounded by a shell containing up to 8 electrons.
 - Also orbital penetration.
 - He recommends that we read the full paper: Lewis (1916).

I.2 Module 2: Molecular Geometries and VSEPR

- The easiest way to approach a new Lewis structure:
 1. Draw a valid Lewis structure for a molecule.
 2. Place electron pairs in the valence shell as far away from each other as possible. Use the σ -bond framework first.
 3. Add π -bonds to complete the molecule.
 - Through the VSEPR approach, think of a molecule as arranged around a central atom A by m atoms or groups of atoms X and n lone electron pairs E .
 - **Steric number:** The sum $n + m$ of groups and electron pairs around the central atom.
 - Steric numbers correspond to geometries.
 - VSEPR is ok but it doesn't capture reality too well.
 - Consider trimethyl boron ($B\text{Me}_3$).
 - Trigonal planar (D_{3h}).
 - Octahedral: O_h .
 - Bent: C_{2v} .
 - **Order of the repulsive forces:** lone pair - lone pair > lone pair - bonding pair > bonding pair - bonding pair.
 - In SF_4 (see-saw), will the lone pair be axial or equatorial?
 - Equatorial — $2 \times 120^\circ$ and $2 \times 90^\circ$ vs. $3 \times 90^\circ$.
 - In BrF_3 ...

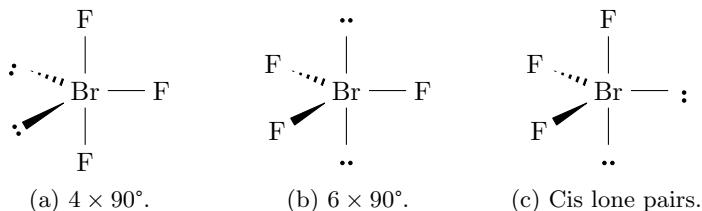


Figure I.1: VSEPR structure of BrF_3 .

- T-shaped → Distorted T — $4 \times 90^\circ$ vs. $6 \times 90^\circ$ or lone pairs in cis-position.
 - In ions such as ICl_4^- , we get square planar (D_{4h}).
 - With mixed substituents (such as PF_2Cl_2)...

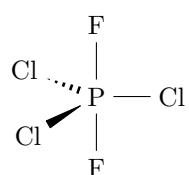


Figure I.2: Lewis structure of PF_2Cl_3 .

- We need **Bent's rule**, which tells us that atoms share electrons from *p*- or *d*-orbitals to a greater extent than they do from *s*-orbitals.
- Thus, when phosphorous excites $3s^23p^3$ to $3s^13p_x^13p_y^13p_z^13d_{z^2}^1$ and then rehybridizes to create three sp^2 orbitals (each composed of $s + p_x + p_y$) and two “*pd*” hybrid orbitals (each composed of $p_z + d_{z^2}$), the equatorial sp^2 orbitals bond to the more **electropositive** chlorines and the axial “*pd*” hybrid orbitals bond to the remaining more electronegative fluorines.
- There is also sometimes a tendency for symmetry.
- **Bent's rule:** Atomic *s*-character concentrates in orbitals directed toward electropositive substituents.
- **Electropositive** (species): A species that has relatively lower electronegativity than another.
- For molecules with multiple bonds, ignore π -bonds.
- Problems with VSEPR:
 - XeF_6 with 14 bonding electrons (7 pairs) is supposed to be pentagonal bipyramidal, but is actually octahedral (a known problem for $14 e^-$ systems).
 - Heavy main group elements with no hybridization.
 - $H-C\equiv C-H$ is linear, but $H-Si\equiv Si-H$ is not.
 - No σ -bond exists in the latter species — it's all π -bonding interactions.
- You maybe don't have to watch the modules and textbook *and* attend class.

I.3 Chapter 3: Simple Bonding Theory

From Miessler et al. (2014).

1/14:

- **Hypervalent** (central atom): A central atom that has an electron count greater than the atom's usual requirement.
- There are rarely more than 18 electrons around a central atom (2 for *s*, 6 for *p*, 10 for *d*). Even heavier atoms with energetically accessible *f* orbitals usually don't have more surrounding electrons because of crowding.
- With BeF_2 , instead of getting the predicted double-bonded Lewis structure, it forms a complex network with Be having coordination number 4.
 - $BeCl_2$ dimerizes to a 3-coordinate structure in the vapor phase.
- Boron trihalides exhibit partial double bond character.
 - It is also possible that the high polarity of B–X bonds and the **ligand-close packing** (LCP) model account for the observed shorter bond length.
 - Boron trihalides also act as Lewis acids.

1/17:

- The variety of structures means that one unified VSEPR theory will not likely^[1] work.
- **Not stereochemically active** (lone pair): “A lone pair that appears in the Lewis-dot structure but has no apparent effect on the molecular geometry” (Miessler et al., 2014, p. 54).
- Double and triple bonds have slightly greater repulsive effects than single bonds in the VSEPR model.
- Multiple bonds tend to occupy the same positions as lone pairs.
- Electronegativity varies for a given atom based on the neighboring atom to which it is bonded.

^[1]Errata: “unlikely.”

1/14:

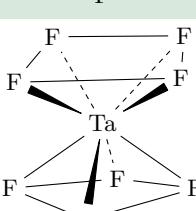
	Steric Number	Geometry	Examples	Calculated Bond Angles	
2	Linear	CO_2	180°	$\text{O} = \text{C} = \text{O}$	
3	Trigonal (triangular)	SO_3	120°		
4	Tetrahedral	CH_4	109.5°		
5	Trigonal bipyramidal	PCl_5	120°, 90°		
6	Octahedral	SF_6	90°		
7	Pentagonal bipyramidal	IF_7	72°, 90°		
8	Square antiprismatic	$[\text{TaF}_8]^{3-}$	70.5°, 9.6°, 109.5°		

Table I.1: VSEPR predictions.

1/17:

- “With the exception of helium and neon, which have large calculated electronegativities and no known stable compounds, fluorine has the largest value” (Miessler et al., 2014, p. 59).
- Although usually classified with Group 1, hydrogen’s chemistry is distinct from that of the alkali metals and actually all of the groups.
- Some bond angle trends can be explained by electronegativity (see Table I.2).
 - For instance, electronegative outer atoms pull electrons away from the central atom, allowing lone pairs to further push together such atoms.
 - Electronegative central atoms pull electrons toward the central atom, pushing bonding pairs farther apart.
- Atomic size can also have effects on VSEPR predictions.

Molecule	X–P–X Angle (°)	Molecule	Bond Angle (°)
PF ₃	97.8	H ₂ O	104.5
PCl ₃	100.3	H ₂ S	92.1
PBr ₃	101.0	H ₂ Se	90.6

Table I.2: Electronegativity and bond angles.

- For example, the C–N–C angle in N(CF₃)₃ is larger than that of N(CH₃)₃ despite the prediction we'd make based on electronegativity alone. This is because F atoms are significantly larger than H atoms so we get some steric hindrance.
- In molecules with steric number 5, axial bond length is greater than equatorial.
- Symmetric structures are often preferred.
- Groups (such as CH₃ and CF₃) have the ability to attract electrons, too — thus, they are also assigned electronegativities.
- **Ligand close-packing:** A model that uses the distances between outer atoms in molecules as a guide to molecular shapes. *Also known as LCP.*
 - Works off of the observation that the nonbonded distances between outer atoms are consistent across molecules with the same central atom, but the bond angles and lengths change.
 - This contrasts with VSEPR theory's concern with the central atom, as opposed to the ligands.
- **Dielectric constant:** “The ratio of the capacitance of a cell filled with the substance to be measured to the capacitance of the same cell with a vacuum between the electrodes” (Miessler et al., 2014, p. 66).
 - This is measured to experimentally determine the polarity of molecules.
- **Dipole moment:** The product Qr of the distance r between two charges' centers and the difference Q between the charges. *Also known as μ .*
 - This is calculated by measuring the dielectric constant at different temperatures.
 - SI unit: Coulomb meter; C m. Common unit: Debye; 1 D = $3.335\,64 \times 10^{-30}$ C m.

Topic II

Symmetry and Group Theory in Chemistry

II.1 Module 3: Symmetry Elements and Operations

1/13:

- He will upload lecture slides in advance in the future.
- An object is symmetric if one part is the same as other parts.
- The symmetry of discrete objects is described using **point symmetry**.
- **Point groups** (~ 32 for molecules) provide us with a way to indicate the symmetry unambiguously.
 - These have symmetry about a single point at the center of mass of the system.
- Extended objects (e.g., crystals) have **translational symmetry** described by **Space groups^[1]** (230 total).
- Readings: Miessler et al. (2014) Chapter 4 and https://en.wikipedia.org/wiki/Molecular_symmetry.
- **Symmetry elements**: Geometric entities about which a **symmetry operation** can be performed. In a point group, all symmetry elements must pass through the center of mass (the point).
- **Symmetry operation**: The action that produces an object identical to the initial object.

Element	Operation
Identity, E	nothing
Rotation axis, C_n	n -fold rotation
Improper rotation axis, S_n	n -fold improper rotation
Plane of symmetry, σ	Reflection
Center of symmetry, i	Inversion

- **Identity**: Does nothing to the object, but is necessary for mathematical completeness.
- **n -fold rotation**: A rotation of $360^\circ/n$ about the C_n axis ($n \in [1, \infty)$).
 - In H_2O , there is a C_2 axis, so we can perform a 2-fold (180°) rotation to get the same molecule.
 - Remember, because of quantum mechanical properties, the hydrogens are indistinguishable so when we rotate it 180° , we cannot tell it apart from the unrotated molecule.
 - Rotations are considered positive in the counterclockwise direction.

¹Not covered in this course.

- Each possible rotation operation is assigned using a superscript integer m of the form C_n^m . m is the number of sequential applications.
- The rotation $C_n^n \equiv E$ is equivalent to the identity operation (nothing is moved).
- Linear molecules have an infinite number of rotational options C_∞ because any rotation on the molecular axis will give the same arrangement.

- **Principal axis:** The highest order rotation axis.

- By convention, the principal axis is assigned to the z -axis if we are using Cartesian coordinates.

- **Reflection:** Exchanges one half of the object with the reflection of the other half.

- **Vertical mirror plane:** A mirror plane that contains the principal axis. *Also known as σ_v .*

- **Horizontal mirror plane:** A mirror plane that is perpendicular to the principal axis. *Also known as σ_h .*

- **Dihedral mirror planes:** A special type of σ_v that is between sides or planes. *Also known as σ_d .*

- For example, we might have vertical mirror planes in the xz - or yz -planes. In this case, the dihedral planes would contain the lines $y = \pm x$.

- Two successive reflections are equivalent to the identity operation.

- **Inversion:** Every part of the object is reflected through the inversion center, which must be at the center of mass of the object.

- $(x, y, z) \xrightarrow{i} (-x, -y, -z)$.

- **n -fold improper rotation:** This operation involves a rotation of $360^\circ/n$ followed by a reflection perpendicular to the axis. It is a single operation and is labeled in the same manner as “proper” rotations. *Also known as S_n^m , rotation-reflection operation.*

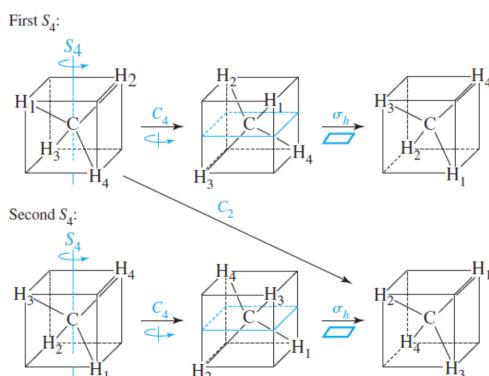


Figure II.1: Methane’s S_4 symmetry.

- Methane has S_4 symmetry.
- Note that $S_1 \equiv \sigma_h$, $S_2 \equiv i$, and sometimes $S_{2n} \equiv C_n$. In methane, for example, $S_4^2 \equiv C_2$.
- Applied to a triangular prism, is a good example.
- If n is even, we have n unique operations. There should be $C_{n/2}$.
- If n is odd, we have $2n$ unique operations. There should be C_n and σ_h .
- The absence of an S_n axis is the defining symmetry property of **chiral** molecules.
 - Formerly, we learned that chiral molecules should not have mirror planes and inversion centers.
 - Rigorously, chiral molecules must not have any improper rotation axes.

II.2 Module 4: Symmetry Point Groups

- Identifying the point groups:
 1. Determine if the symmetry is special (e.g., octahedral).
 2. Determine if there is a principal rotation axis.
 3. Determine if there are rotation axes perpendicular to the principal axis.
 4. Determine if there are mirror planes.
 5. Assign point groups.
- High symmetry and low symmetry groups are the most difficult to identify.
- High symmetry:
 - Perfect tetrahedral (T_d), e.g., P_4 and CH_4 .
 - Perfect octahedral (O_h), e.g., SF_6 .
 - Perfect icosahedral (I_h), e.g., C_{60} and $B_{12}H_{12}^{2-}$.
- Low symmetry:



Figure II.2: Low symmetry point groups.

- Only an improper axis: S_n .
- Only a mirror plane: C_s .
- Only an inversion center: C_i .
- No symmetry: C_1 .
- C_n groups:
 - Only a C_n axis. Note that conformation is important.
- C_{nh} groups have a C_n axis and a σ_h reflection plane (such as $B(OH)_3$).
 - H_2O_2 has C_{2h} symmetry.
- All symmetry elements are listed in the top row of the corresponding characters table (Appendix C in Miessler et al. (2014)).
- C_{nv} groups have a C_n axis and a σ_v reflection plane.
 - NH_3 has C_{3v} symmetry.
 - CO has $C_{\infty v}$ symmetry since there are an infinite number of both C_n axes and σ_v mirror planes.
- D_{nh} groups: A C_n axis, n perpendicular C_2 axes, and a σ_h reflection plane.

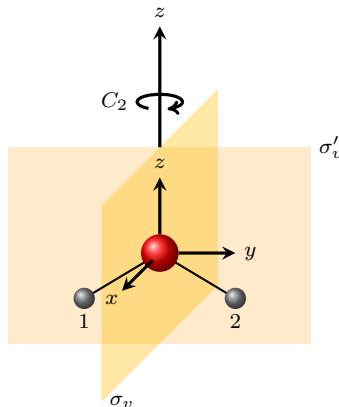
- BH_3 has D_{3h} symmetry.
- A square prism has D_{4h} symmetry.
- CO_2 has $D_{\infty h}$ symmetry.
- D_n groups: A C_n axis, n perpendicular C_2 axes, and no mirror planes.
 - A 3-bladed propeller has D_3 symmetry.
- D_{nd} groups: A C_n axis, n perpendicular C_2 axes, and a σ_d .
 - Ethane in the staggered conformation has D_{3d} symmetry.
- Local symmetry:
 - Sometimes, rigorous math analysis needs to be adjusted to physical reality.
 - If a cyclopentane ring is bonded through the center to $\text{Mn}(\text{CO})_3$, this molecule has only C_s symmetry.
 - However, spectroscopically, there is fast rotation about the $\text{Mn}-\text{Cp}$ bond. This means that the $\text{Mn}(\text{CO})_3$ fragment exhibits pseudo- C_{3v} symmetry while the C_5H_5 ligand exhibits pseudo- C_{5v} symmetry.
 - Often, the absolute symmetry of a molecule is very low, but the interactions are far away from the centers of interest, and do not perturb them significantly.
 - If we have platinum as a central atom bonded to two chlorines and two $\text{P}(\text{Et})_3$ groups, this molecule technically has C_1 symmetry due to the orientations of atoms within R groups (staggered), but IR spectroscopy is characteristic of highly symmetric species (D_{2h}).

II.3 Module 5: Group Theory 101

1/15:

- **Group:** A set of elements together with an operation that combines any two of its elements to form a third element satisfying four conditions called the group axioms.
- **Closure:** All binary products must be members of the group.
- **Associativity:** Associative law of multiplication must hold.
- **Identity:** A group must contain the identity operator.
- **Inverse:** Every operator must have an inverse.
- The integers with the addition operation form a group, for example.
- History:
 - Early group theory was driven by the quest for solutions of polynomial equations of degree 5 and above.
 - Early 1800s: Évariste Galois realized that the algebraic solution to a polynomial equation is related to the structure of a group of permutations associated with the roots off the polynomial, the Galois group of the polynomial.
 - Link to Galois video here.
 - 1920s: Group theory was applied to physics and chemistry.
 - 1931: It is often hard or even impossible to obtain a solution of the Schrödinger equation — however, a large part of qualitative results can be obtained by group theory. Almost all the rules of spectroscopy follow from the symmetry of a problem.
- We will use group theory for describing symmetry of molecules. We will use group theory to understand the bonding and spectroscopic features of molecules.

- For us, a group consists of a set of symmetry elements (and associated symmetry operations) that completely describes the symmetry of a molecule.
- **Order** (of a group): The total number of elements (i.e., symmetry operations) in the group. *Also known as h .*
- Rule 1: Closure.

Figure II.3: Symmetry elements for H_2O .

- H_2O is of the C_{2v} point group (refer to Figure II.3).
 - Symmetry operations: E , C_2 , $\sigma_{v(xz)}$, and $\sigma'_{v(yz)}$.
 - $\sigma_v \cdot C_2 = \sigma'_v = C_2 \cdot \sigma_v$.
 - The above property (order *does not* matter) shows that C_{2v} is an **Abelian group**.
- NH_3 is of the C_{3v} point group.
 - Symmetry operations: E , C_3^+ , C_3^- , σ_v , σ'_v , and σ''_v .
 - $\sigma''_v \cdot C_3 = \sigma_v$, but $C_3 \cdot \sigma''_v = C_3^- = C_3^2$.
 - The above property (order *does* matter) shows that C_{3v} is a **non-Abelian group**.
- Rule 2: Associativity.

- H_2O is of the C_{2v} point group (refer to Figure II.3).

$$\begin{aligned} \sigma'_v C_2 \sigma_v(1, 2) &= \sigma'_v C_2(2, 1) & \sigma'_v(C_2 \sigma_v)(1, 2) &= \sigma'_v E(1, 2) & (\sigma'_v C_2) \sigma_v(1, 2) &= \sigma_v \sigma_v(1, 2) \\ &= \sigma'_v(1, 2) & &= \sigma'_v(1, 2) & &= \sigma_v(2, 1) \\ &= (1, 2) & &= (1, 2) & &= (1, 2) \end{aligned}$$

- Rule 3: Identity.
- Rule 4: Inverse.
- For a C_{2v} point group:

$$E \cdot E = E \quad C_2 \cdot C_2 = E \quad \sigma_v \cdot \sigma_v = E \quad \sigma'_v \cdot \sigma'_v = E$$

- Group multiplication tables.

C_{2h}	E	C_2	σ_h	i
E	E	C_2	σ_h	i
C_2	C_2	E	i	σ_h
σ_h	σ_h	i	E	C_2
i	i	σ_h	C_2	E

Table II.1: Group multiplication table for the C_{2h} point group.

- Table II.1 corresponds to the C_{2h} point group, which has E , C_2 , σ_h , and i operations.
- Note that the operation in the top row is the one that's applied first, while the one in the left column will be applied second.
- **Subgroup:** Fractional parts of groups that are groups, too.

C_{3v}	E	C_3	C_3^2	σ_v	σ'_v	σ''_v
E	E	C_3	C_3^2	σ_v	σ'_v	σ''_v
C_3	C_3	C_3^2	E	σ''_v	σ_v	σ'_v
C_3^2	C_3^2	E	C_3	σ'_v	σ''_v	σ_v
σ_v	σ_v	σ'_v	σ''_v	E	C_3	C_3^2
σ'_v	σ''_v	σ_v	σ'_v	C_3	C_3^2	E
σ''_v	σ'_v	σ''_v	σ_v	C_3^2	E	C_3

Table II.2: Group multiplication table for the C_{3v} point group.

- If $h = 6$ (as in the C_{3v} group), subgroup order can be $h = \textcolor{purple}{3}, \textcolor{blue}{2}, \textcolor{red}{1}$. Why only these?
- The order $\textcolor{red}{1}$ and $\textcolor{purple}{3}$ charts are subgroups.
- The order $\textcolor{blue}{2}$ chart is not a subgroup because C_3^2 is not an operation in the group (therefore, the “subgroup” is not closed).
- We use subgroups because they can make complex problems simpler.
 - For example, calculating the vibrational modes of CO_2 .
 - As another example, D_{2h} is a subgroup of $D_{\infty h}$.

II.4 Module 6: Representations

- Items of the same point group have the same vibration modes.
- **Representation** (of a group): Any collection of quantities (or symbols) which obey the multiplication table of a group. *Also known as Γ .*
- For our purposes, these quantities are the matrices that show how certain characteristic of a molecule behave under the symmetry operations of the group.
- Operations (on a point (x, y, z) in Cartesian coordinates):
 - $E(x, y, z) = (x, y, z)$.
 - $\sigma_{xz}(x, y, z) = (x, -y, z)$.
 - $i(x, y, z) = (-x, -y, -z)$.
 - C_n : Convention is a counterclockwise rotation of the point by $\theta = \frac{2\pi}{n}$ radians.
 - S_n : Convention is a clockwise rotation of the point C_n followed by a σ through a plane perpendicular to C .

- Matrix forms of operations:

– Identity: $E = \begin{bmatrix} 1 & 0 & 0 \\ 0 & 1 & 0 \\ 0 & 0 & 1 \end{bmatrix}$.

– One example of a reflection (there are two more): $\sigma_{xy} = \begin{bmatrix} 1 & 0 & 0 \\ 0 & 1 & 0 \\ 0 & 0 & -1 \end{bmatrix}$.

– Inversion: $i = \begin{bmatrix} -1 & 0 & 0 \\ 0 & -1 & 0 \\ 0 & 0 & -1 \end{bmatrix}$.

– Rotation: Counterclockwise is $C_n(\theta) = \begin{bmatrix} \cos \theta & -\sin \theta & 0 \\ \sin \theta & \cos \theta & 0 \\ 0 & 0 & 1 \end{bmatrix}$ and clockwise is $C_n(\theta) = \begin{bmatrix} \cos \theta & \sin \theta & 0 \\ -\sin \theta & \cos \theta & 0 \\ 0 & 0 & 1 \end{bmatrix}$.

- A derivation of this matrix can be found in the slides.
- Improper rotation: $S_n(\theta) = \sigma_h C_n(\theta)$.

- **Reducible representations (Γ):**

- A representation of a symmetry operation of a group.
- Can be expressed in terms of a representation of lower dimension.
- Can be broken down into a simpler form.
- Characters can be further diagonalized.
- Are composed of the direct sum of irreducible representations.
- Infinite possibilities.

- **Irreducible representations (Γ_i):**

- A fundamental representation of a symmetry operation of a group.
- Cannot be expressed in terms of a representation of lower dimension.
- Cannot be broken down into a simpler form.
- Characters cannot be further diagonalized.
- Small finite number dictated by point group.

- Good example of reducible/irreducible representations?
- A representation shows how certain characteristics of an object (a basis) behave under the symmetry operation of the group.
- **Conjugate elements:** Two elements X and Y for which there exists an element Z in the group such that

$$Z^{-1} \cdot X \cdot Z = Y$$

- Every element is conjugated with itself (let $Z = E$).
- If X is conjugated with Y , then Y is conjugated with X .
- If X is conjugated with Y and W , then Y and W are also conjugate.

- **Class:** A complete set of elements of a group that are conjugate to one another.
 - Geometric meaning: operations in the same class can be converted into one another by changing the axis system through application of some symmetry operation of the group.
- Find the conjugates to C_3 in the C_{3v} point group (refer to Table II.2 throughout the following discussion).

- Let $X = C_3$, let Z iterate through the six symmetry elements $(E, C_3, C_3^2, \sigma_v, \sigma'_v, \sigma''_v)$, and let Z^{-1} iterate through the corresponding inverses $(E, C_3^2, C_3, \sigma_v, \sigma'_v, \sigma''_v)$.
- Thus, we have

$$\begin{aligned} E \cdot C_3 \cdot E &= C_3 \\ C_3^2 \cdot C_3 \cdot C_3 &= C_3 \\ C_3 \cdot C_3 \cdot C_3^2 &= C_3 \\ \sigma_v \cdot C_3 \cdot \sigma_v &= C_3^2 \\ \sigma'_v \cdot C_3 \cdot \sigma'_v &= C_3^2 \\ \sigma''_v \cdot C_3 \cdot \sigma''_v &= C_3^2 \end{aligned}$$

- It follows from the above that C_3 and C_3 are conjugates, and C_3 and C_3^2 are conjugates.
- Thus, C_3 and C_3^2 are in the same class.
- We can use a similar method to prove that σ_v, σ'_v , and σ''_v are all in the same class within the C_{3v} point group.
- Likewise E is in a class by itself.
- Thus, for the C_{3v} point group, E forms a class of order 1, C_3, C_3^2 form a class of order 2, and $\sigma_v, \sigma'_v, \sigma''_v$ form a class of order 3.

- **Similarity transformation:** The transformation

$$v^{-1} \cdot A \cdot v = A'$$

- A is a representation for some type of symmetry operation.
- v is a similarity transform operator.
- v^{-1} is the inverse of the similarity transform operator.
- A' is the product.
- A and A' are conjugates, and we say that A' is the similarity transform of A by v .

- **Block-diagonal (matrix):** A matrix with nonzero values only in square blocks along the diagonal from the top left to the bottom right.

$$\begin{bmatrix} 2 & 3 & 0 & 0 & 0 \\ 1 & 2 & 0 & 0 & 0 \\ 0 & 0 & 1 & 1 & 0 \\ 0 & 0 & 1 & 1 & 0 \\ 0 & 0 & 0 & 0 & 2 \end{bmatrix}$$

- The above matrix is an example of a block-diagonal matrix.
- Irreducible representations are the corresponding blocks within a set of block-diagonalized matrices representing each operation in a group.

II.5 Module 7: Characters and Character Tables

1/19:

- Nocera Lecture 3 notes (on Canvas) explain how reducible and irreducible transformations are related to each other through the similarity transformations.
- For H_2O , each atom has 3 Cartesian coordinates, so our transformation matrix^[2] is 9-square.

²Some values in it are negative because of the cosine/sine definition of a rotation matrix for $\theta = 180^\circ$.

- However, we can also apply a smaller matrix to the molecule as a whole and invoke symmetry to find the position of the individual atoms.
- **Characters** (of a representation): The traces (i.e., sums of the diagonal matrix elements) of the representation matrices for each operation. *Also known as χ .*
 - The character is an invariant for each type of symmetry operation (e.g., regardless of the axis about which a C_n operation is performed, the trace of the corresponding matrix will be the same).
- Common characters:
 - C_n character: $\chi = 2 \cos \theta + 1$.
 - σ_v, σ_d character: $\chi = 1$.
 - $S_2 \equiv i$ character: $\chi = -3$.
 - S_n character: $\chi = 2 \cos \theta - 1$.
 - $\perp C_2$ axes character: $\chi = -1$.
- **Character table:** The collection of characters for a given irreducible representation, under the operations of a group. *Also known as χ table.*

	Group Symbol	Symmetry Elements				linear	quadratic
		E	C_2	$\sigma_v(xz)$	$\sigma'_v(yz)$		
Irreducible Representations	A_1	1	1	1	1	z	x^2, y^2, z^2
	A_2	1	1	-1	-1	R_z	xy
	B_1	1	-1	1	-1	x, R_y	xz
	B_2	1	-1	-1	1	y, R_x	yz
Characters						Basis Functions	

Figure II.4: A character table.

- Character tables for all point groups are listed in Appendix C of Miessler et al. (2014).
- **Mulliken symbols** are used to classify irreducible representations based on degeneracy and symmetry.
 - A or B: singly degenerate (the maximum block size in the block-diagonalized irreducible transformation matrix is 1×1).
 - E: Doubly degenerate (the maximum block size in the block-diagonalized irreducible transformation matrix is 2×2).
 - T: Triply degenerate (the maximum block size in the block-diagonalized irreducible transformation matrix is 3×3).
 - A: symmetric (+) with respect to C_n .
 - B: anti-symmetric (-) with respect to C_n .
 - Subscript g: symmetric (+) with respect to i . *Etymology* short for gerade (German for symmetric).
 - Subscript u: anti-symmetric (-) with respect to i . *Etymology* short for ungerade (German for unsymmetric).
 - If the molecule has a center of inversion, we label irreducible representations with g or u .
 - Subscript 1: symmetric (+) with respect to $\perp C_2$ or σ_v .
 - Subscript 2: anti-symmetric (-) with respect to $\perp C_2$ or σ_v .
 - Superscript ': symmetric (+) under σ_h (if no i).
 - Superscript "': anti-symmetric (-) under σ_h (if no i).

- Don't mistake the operation E for the Mulliken symbol E .
- To assign Mulliken symbols, use the character table.
 - Assigning the main letter:
 - If E -character = 1 and C_n -character = 1: A .
 - If E -character = 1 and C_n -character = -1: B .
 - If E -character = 2: E .
 - If E -character = 3: T .
 - Assigning a subscript g or u :
 - If i -character = 1: g .
 - If i -character = -1: u .
 - This subscript can be assigned to A, B, E, T representations.
 - Assigning a superscript ' or '':
 - If σ_h -character = 1: '.
 - If σ_h -character = -1: ''.
 - This subscript can be assigned to A, B representations.
 - Assigning a subscript 1 or 2:
 - If $\perp C_2$ or σ_v -character = 1: 1.
 - If $\perp C_2$ or σ_v -character = -1: 2.
 - This subscript can be assigned to A, B representations.
- σ , π , and δ bonds come from the Mulliken symbols!
 - Infinity tables use Greek rather than Latin letters.

D_3	E	$2C_3(z)$	$3C'_2$	linear	quadratic
A_1	1	1	1		$x^2 + y^2, z^2$
A_2	1	1	-1	z, R_z	
E	2	-1	0	$(x, y)(R_x, R_y)$	$(x^2 - y^2, xy)(xz, yz)$.

Table II.3: Character table for the D_3 point group.

- In character tables, we need to multiply each symmetry operation by the number of types there are (see Table II.3).
- Basis functions show us how different functions transform under different symmetry operations.
- In the C_{2v} point group:
 - The p_x orbital has B_1 symmetry.
 - p_x transforms as B_1 .
 - p_x has the same symmetry as B_1 .
 - p_x forms a basis for the B_1 irreducible representation.
 - p_x is B_1 because (see Table 0.2 and Figure II.4) it does not change under E , it inverts under C_2 , it does not change under $\sigma_v(xz)$, and it inverts under $\sigma'_v(yz)$ (note that the C_2 axis is the z -axis, not the x -axis).
- We can apply the same procedure to other more complex functions.
 - For example, in the C_{2v} point group, we know that (with respect to orbitals):

- p_y is B_2 .
 - p_z is A_1 .
 - d_{z^2} is A_1 .
 - $d_{x^2-y^2}$ is A_1 .
 - d_{yz} is B_2 .
 - d_{xy} is A_2 .
 - d_{xz} is B_1 .
 - We can even go into the cubic functions describing the f orbitals and assign them Mulliken symbols.
 - Essentially, the right hand side of a character table tells you how atomic orbitals will transform under certain symmetry operations.
 - Properties of a character table:
 1. The characters of all matrices belonging to the operations in the same class are identical in a given irreducible representation.
 - We most commonly form a **rotational class** and a **reflection class**.
 2. The number of irreducible representations in a group is equal to the number of classes of that group.
 3. There is always a totally symmetric representation for any group.
 - I.e., a representation where every character is 1.
 4. The sum of the squares of the **dimensionality** of all the irreducible representations is equal to the order of the group. Mathematically,
- $$h = \sum_i [\chi_i(E)]^2$$
- For example, the dimensionalities of the D_3 point group (see Table II.3) are 1, 1, and 2, and the order is, indeed, $6 = 1^2 + 1^2 + 2^2$.
 - 5. The sum of the squares of the characters multiplied by the number of operations in the class equals the order of the group. Mathematically,
- $$h = \sum_{R_c} g_c [\chi_i(R_c)]^2$$
- For example, with respect to the D_3 point group (see Table II.3),

$$\begin{aligned} 6 &= (1)(1)^2 + (2)(1)^2 + (3)(1)^2 \\ &= (1)(1)^2 + (2)(1)^2 + (3)(-1)^2 \\ &= (1)(2)^2 + (2)(-1)^2 + (3)(0)^2 \end{aligned}$$
 - 6. The sum of the products of the corresponding characters of any two different irreducible representations of the same group is zero. Mathematically,
- $$\sum_{R_c} g_c \chi_i(R_c) \chi_f(R_c) = 0$$
- Basically, this means that if we treat irreducible representations as vectors in h -space, they are orthogonal.
 - For example, with respect to the D_3 point group (see Table II.3),

$$\begin{aligned} 0 &= 1(1)(1) + 2(1)(1) + 3(1)(-1) \\ &= 1(1)(2) + 2(1)(-1) + 3(1)(0) \\ &= 1(1)(2) + 2(1)(-1) + 3(-1)(0) \end{aligned}$$

- **Dimensionality:** The character of the identity operation E . Also known as **dimension**.

II.6 Module 8: Using Character Tables

- A reducible representation of a group is any representation Γ of the form

$$\Gamma = \sum_i a_i \Gamma_i$$

where each Γ_i is an irreducible representation of the group and a_i is a real scalar.

- Basically, a reducible representation is any linear combination of irreducible representations.
- For example, with respect to the C_{2v} point group (see Figure II.4), $\Gamma = (7, 1, 5, 3) = 4A_1 + 2B_1 + B_2$ is a reducible representation.
- We may “factor” reducible representations by inspection, or by the...
- Decomposition/reduction formula for a reducible representation:

$$a_i = \frac{1}{h} \sum_Q N \cdot \chi(R)_Q \cdot \chi_i(R)_Q$$

- a_i is the number of times the irreducible representation appears in Γ .
- h is the order of the group.
- N is the number of operations in class Q .
- $\chi(R)_Q$ is the character of the reducible representation.
- $\chi_i(R)_Q$ is the character of the irreducible representation.
- This formula cannot be applied to $D_{\infty h}$ and $C_{\infty v}$.
- Let’s look at decomposing $\Gamma = (7, 1, 5, 3)$ into its component irreducible point groups using the above formula.

$$\begin{aligned} a_{A_1} &= \frac{1}{4}(1 \cdot 7 \cdot 1 + 1 \cdot 1 \cdot 1 + 1 \cdot 5 \cdot 1 + 1 \cdot 3 \cdot 1) = 4 \\ a_{A_2} &= \frac{1}{4}(1 \cdot 7 \cdot 1 + 1 \cdot 1 \cdot 1 + 1 \cdot 5 \cdot -1 + 1 \cdot 3 \cdot -1) = 0 \\ a_{B_1} &= \frac{1}{4}(1 \cdot 7 \cdot 1 + 1 \cdot 1 \cdot -1 + 1 \cdot 5 \cdot 1 + 1 \cdot 3 \cdot -1) = 2 \\ a_{B_2} &= \frac{1}{4}(1 \cdot 7 \cdot 1 + 1 \cdot 1 \cdot -1 + 1 \cdot 5 \cdot -1 + 1 \cdot 3 \cdot 1) = 1 \end{aligned}$$

- You can also find websites that will apply the formula for you.
- Basis → reducible representation → irreducible representations workflow:
 1. Assign a point group.
 2. Choose a basis function (bond, vibration, orbital, angle, etc.).
 3. Apply operations.
 - The following shortcuts allow us to skip matrix math in certain situations.
 - If the basis stays the same: +1.
 - If the basis is reversed: -1.
 - If it is a more complicated change: 0.
 4. Generate a reducible representation.
 5. Reduce to irreducible representation.
- We now look at an example of applying the above method to H_2O .

- H_2O is of the C_{2v} point group.
- The 9×9 identity matrix represents the identity operation on the 3 atoms in H_2O , each described by 3 Cartesian coordinates. Thus, $\chi(E) = 9$.
- The following matrix represents the C_2 symmetry operation. Thus, $\chi(C_2) = -1$.

$$\begin{array}{l} \text{O} \left\{ \begin{bmatrix} x' \\ y' \\ z' \end{bmatrix} \right. \\ \text{H}_a \left\{ \begin{bmatrix} x' \\ y' \\ z' \end{bmatrix} \right. \\ \text{H}_b \left\{ \begin{bmatrix} x' \\ y' \\ z' \end{bmatrix} \right. \end{array} = \begin{bmatrix} -1 & 0 & 0 & 0 & 0 & 0 & 0 & 0 & 0 \\ 0 & -1 & 0 & 0 & 0 & 0 & 0 & 0 & 0 \\ 0 & 0 & 1 & 0 & 0 & 0 & 0 & 0 & 0 \\ 0 & 0 & 0 & 0 & 0 & 0 & -1 & 0 & 0 \\ 0 & 0 & 0 & 0 & 0 & 0 & 0 & -1 & 0 \\ 0 & 0 & 0 & 0 & 0 & 0 & 0 & 0 & 1 \\ 0 & 0 & 0 & -1 & 0 & 0 & 0 & 0 & 0 \\ 0 & 0 & 0 & 0 & -1 & 0 & 0 & 0 & 0 \\ 0 & 0 & 0 & 0 & 0 & 1 & 0 & 0 & 0 \end{bmatrix} \begin{bmatrix} x \\ y \\ z \end{bmatrix} \end{array} \left. \begin{array}{l} \text{O} \\ \text{H}_a \\ \text{H}_b \end{array} \right\}$$

- Note that atoms moved during the transformation do not contribute to the character of the transformation matrix.
- Since under $\sigma_v(xz)$ only O is unshifted, we need only consider its part of the transformation matrix (as follows) when looking for the character. Thus, $\chi(\sigma_v(xz)) = 1$ ³.

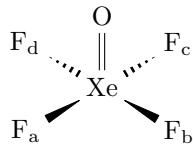
$$\begin{bmatrix} x' \\ y' \\ z' \end{bmatrix} = \begin{bmatrix} 1 & 0 & 0 \\ 0 & -1 & 0 \\ 0 & 0 & 1 \end{bmatrix} \begin{bmatrix} x \\ y \\ z \end{bmatrix}$$

- Since under $\sigma_v(yz)$ no atom is shifted, we need to consider each (identical) part of the transformation matrix (as follows) when looking for the character. Thus, $\chi(\sigma_v(yz)) = 3$.

$$\begin{bmatrix} x' \\ y' \\ z' \end{bmatrix} = \begin{bmatrix} -1 & 0 & 0 \\ 0 & 1 & 0 \\ 0 & 0 & 1 \end{bmatrix} \begin{bmatrix} x \\ y \\ z \end{bmatrix}$$

- Thus, the characters of our final reducible representation is $\Gamma_{3N} = (9, -1, 1, 3)$.
- This representation represents fully unrestricted motion of all 3 ambiguities of freedom.

- Another example: XeOF_4 .

Figure II.5: XeOF_4 .

- Point group: C_{4v} .
- Basis function: F atoms.
- Let's see what happens to the fluorine atoms under the C_{4v} operations (remember that if a basis element (a fluroine atom) stays the same, it contributes +1 to the character of an operation, and if it moves, it contributes 0 to the character).

³Note also that $a_{11} = 1$ since the x -vector (a basis) is unchanged under $\sigma_v(xz)$, $a_{22} = -1$ since the y -vector is reversed under $\sigma_v(xz)$, and $a_{33} = 1$ since the z vector is unchanged under $\sigma_v(xz)$.

E	all unchanged	4
C_4	all move	0
C_2	all move	0
$2\sigma_v$	2 move, 2 unchanged	2
$2\sigma_d$	all move	0

Table II.4: Changes in the fluorine atoms of XeOF_4 under the C_{4v} symmetry operations.

- Thus, $\Gamma = (4, 0, 0, 2, 0)$.
- With the C_{4v} character table and the decomposition formula, we can discover that $\Gamma = A_1 + B_1 + E$.

II.7 Nocera Lecture 3

From Nocera (2008).

- 1/25:
- Corresponding blocks within block-diagonalized matrices obey the same multiplication properties as the block-diagonalized matrices, themselves.
 - One basis may not uncover all irreducible representations.
 - For example, we can use a Cartesian basis to uncover the A_1 and E irreducible representations of the D_3 point group (see Table II.3) and the basis function R_z ($E : R_z \rightarrow R_z$, $C_3 : R_z \rightarrow R_z$, $\sigma_v : R_z \rightarrow R_z$) to uncover the A_2 irreducible representation.
 - We can also construct a character table algebraically. Continuing with the D_3 example...
 - There is one totally symmetric representation $\Gamma_i = (1, 1, 1)$.
 - $h = 6 = \sum_i [\chi_i(E)]^2$. Since each $\chi_i(E)$ can only be 1, 2, or 3, we know that $\chi_1(E) = \chi_2(E) = 1$ and $\chi_3(E) = 2$.
 - This combines with orthogonality to reveal that

$$\begin{aligned} 0 &= \sum_{R_c} g_c \chi_1(R_c) \chi_2(R_c) \\ &= 1 \cdot 1 \cdot 1 + 2 \cdot 1 \cdot \chi_2(C_3) + 3 \cdot 1 \cdot \chi_2(\sigma_v) \\ &= 1 + 2\chi_2(C_3) + 3\chi_2(\sigma_v) \end{aligned}$$

i.e., $\Gamma_i = (1, 1, -1)$.

- To determine the last representation, we can use two rules and solve as a system of equations.

$$\begin{aligned} 0 &= \sum_{R_c} g_c \chi_1(R_c) \chi_3(R_c) & 6 &= \sum_{R_c} g_c [\chi_3(R_c)]^2 \\ &= 2 + 2\chi_3(C_3) + 3\chi_3(\sigma_v) & &= 4 + 2[\chi_3(C_3)]^2 + 3[\chi_3(\sigma_v)]^2 \end{aligned}$$

From this, we can determine that $\chi_3(C_3) = -1$ and $\chi_3(\sigma_v) = 0$. Therefore, $\Gamma_i = (2, -1, 0)$.

II.8 TA Review Session 1

- 1/22:
- We also don't need to show the equatorial electron pair in the shape picture.
 - For SeCl_4 , we need to show that the axial bonds are longer and are not straight up and down.
 - For I_3^- , it's trigonal bipyramidal EPA and linear molecular geometry — there's an extra electron pair around the central iodine atom. This makes it $D_{\infty h}$ point group.

- For SeOCl_4 , we should also make the axial longer and bent away from the oxygen atom.
 - This molecule is sp^3d hybridized, not sp^2d^2 , as it would be if the oxygen were axial.
- For $\text{IO}(\text{OH})_5$, we should show the equatorial OH's pushed away from the oxygen atom. The oxygen bond should also be a bit shorter.
- In the exam, they will specify whether we need to consider the hydrogens or not when calculating symmetry.
- For ClOF_4^- , one or the other of the lone pair/oxygen will push the equatorial fluorines a bit. You don't have to know which, just show bent.
- For XeO_2F_2 , show the lone pair pushing the axial fluorines away.
- For IF_3^{2-} , show that this is a *distorted T*.
- A tennis ball belongs to the D_{2d} point group since you have perpendicular C_2 axes punching through the seam and mirror planes at 45° angles to the C_2 axes (i.e., dihedral).
- FeF_6^{3-} also loses 4 C_2 axes.
- There are three possible isomers of IF_3O_2 (see Figure I.1), but the one with equatorial oxygens is the most stable.
 - The structure analogous to Figure I.1a has C_{2v} symmetry.
 - The structure analogous to Figure I.1b has D_{3h} symmetry.
 - The structure analogous to Figure I.1c has C_s symmetry.
- Determine the point group of NO_3^{2-} , $\text{HFC}=\text{C}=\text{CHF}$, $\text{H}_2\text{C}=\text{CF}_2$, and C_2H_6 (consider three possible conformers).
 - NO_3^{2-} is of the D_{3h} point group.
 - $\text{HFC}=\text{C}=\text{CHF}$ is of the C_2 point group.
 - $\text{H}_2\text{C}=\text{CF}_2$ is of the C_{2v} point group.
 - C_2H_6 eclipsed has D_{3h} symmetry, staggered has D_{3d} symmetry, and others have D_3 symmetry.

II.9 Module 9: Molecular Vibrations

1/25:

- If you have a nonlinear triatomic molecule (like H_2O), we have 3 atoms \times 3 DOF = 9 DOF.
 - This accounts for any possible perturbations of our molecules.
- We can break this down into three types of motion:
 - Translational motion along the x -, y -, and z -axes (3 DOFs).
 - Rotational motion about the x -, y -, and z -axes (3 more DOFs).
 - The remaining degrees of freedom are vibrational.
- Every *nonlinear* molecule has 3 translational and 3 rotational DOFs.
 - The number of vibrations is $3N - 6$, where N is the number of atoms.
- Every *linear* molecule has 3 translational and 2 rotational DOFs.
 - The number of vibrations is $3N - 5$, where N is the number of atoms.
- **Vibrational mode:** A perturbation of molecular structure that keeps the center of mass of the molecule in one place.

- For H_2O , there are three vibrational modes: antisymmetric stretch (one H moves away from the O as the other moves in, and then the process reverses), symmetric stretch (both H's move away from the O and then move back in), and scissoring bend (the bond angle changes).
- For molecules in general, there are these three and an additional three: wagging (in H_2O , the H's rotate out of the plane of the molecule and then back), twisting (in H_2O , both H's rotate out of the plane of the molecule, but one in one direction and one in the other), and rocking (in H_2O , both H's rotate about the O within the plane of the molecule preserving the bond angle, and then go back).
- What vibrational modes a molecule can and cannot have is governed by its point group and group theory.
- Each normal mode of vibration forms a basis for an irreducible representation of the point group of the molecule.
- Workflow to identify the vibrational properties of a molecule:
 1. Find number/symmetry of vibrational modes.
 2. Assign the symmetry of known vibrations.
 3. What does the vibration look like?
 4. Find if a vibrational mode is IR or Raman Active.
- To find vibrational modes, follow the 5-step Basis \rightarrow reducible representation \rightarrow irreducible representation workflow and then subtract translational and rotational motion.
 - For example, with H_2O , as before, we have C_{2v} point group, $3N$ basis, $\Gamma_{3N} = (9, -1, 1, 3)$, and $\Gamma_{3N} = 3A_1 + A_2 + 2B_1 + 3B_2$.
 - We now need to use the basis functions in the C_{2v} character table to determine the translational and rotational degrees of freedom. From Figure II.4, the x, y, z -translational modes come from the B_1, B_2 , and A_1 Γ_i 's, respectively. Thus, the total translational mode, sufficient to describe all possible translations in 3 DOFs, is $\text{trans} = A_1 + B_1 + B_2$. Similarly, we can determine that the total rotational mode is $\text{rot} = A_2 + B_1 + B_2$.
 - Thus, we can determine that the total vibrational mode is $\Gamma_{3N} - \text{trans} - \text{rot} = 2A_1 + B_2$.
 - When H_2O scissors or symmetrically stretches, it maintains its C_{2v} symmetry. Thus, both of these vibrations are represented by the totally symmetric irreducible representation A_1 (hence the $2A_1$ component).
 - As to asymmetric stretch, there exist points in time where H_2O lowers its symmetry from C_{2v} to C_s , losing its C_2 and $\sigma_v(xz)$ symmetry elements but maintaining E and $\sigma_v(yz)$. This shift is encapsulated by the B_2 irreducible representation.
 - We can see all of these modes in H_2O 's infrared absorption spectrum.
 - Note that this spectrum does not tell us the energy of vibrations (we can model this with quantum mechanics), or if it is IR or Raman active (which modes will appear in the respective spectrum).
- Determine the number and symmetry types for the translations, vibrations, and rotations of PF_5 .
 - Point group: D_{3h} .
 - We could write out all 18 x, y, z -vectors, or we can use a shortcut: To construct Γ_{3N} , we can, for each symmetry element, multiply the number of atoms atoms unmoved under the operation by the corresponding character the representation that accounts for x, y, z -movements.

$$\Gamma_{3N} = (\Gamma_{x,y,z}) \cdot (\# \text{ unmoved atoms})$$

Applying this shortcut, we have $\Gamma_{x,y,z} = E' + A''_2 = (3, 0, -1, 1, -2, 1)$. We also have # unmoved atoms = (6, 3, 2, 4, 1, 4). Thus, $\Gamma_{3N} = (18, 0, -2, 4, -2, 4)$.

D_{3h}	E	$2C_3$	$3C_2$	σ_h	$2S_c$	$3\sigma_v$	linear	quadratic
A'_1	1	1	1	1	1	1		$x^2 + y^2, z^2$
A'_2	1	1	-1	1	1	-1	R_z	
E'	2	-1	0	2	-1	0	(x, y)	$(x^2 - y^2, xy)$
A''_1	1	1	1	-1	-1	-1		
A''_2	1	1	-1	-1	-1	1	z	
E''	2	-1	0	-2	1	0	(R_x, R_y)	(xz, yz)

Table II.5: Character table for the D_{3h} point group.

- We can now reduce it into $\Gamma_{3N} = 2A'_1 + A'_2 + 4E' + 3A''_2 + 2E''$.
 - Note that we still have our 18 degrees of freedom because each E is double degenerate, meaning that $4E'$ counts for 8 degrees of freedom and $2E''$ counts for 4 degrees of freedom.
- This combined with the fact that $\Gamma_{\text{trans}} = \Gamma_{x,y,z} = E' + A''_2$ and $\Gamma_{\text{rot}} = A'_2 + E''$ implies that $\Gamma_{\text{vibs}} = 2A'_1 + 3E' + 2A''_2 + E''$.
 - For a similar reason to the note above, we can count $12 = 3N - 6$ degrees of freedom (vibrational modes) in Γ_{vibs} .
 - Although degenerate modes count for multiple DOFs, the multiple modes they count for are of the same type.
 - In an ideal world, those modes corresponding to E would be twice as intense as the others.
 - From the definition of Γ_{vibs} and the above discussion, we know that PF_5 has $2+3+2+1=8$ distinct vibrational modes.
- There are (broadly) two types of vibrations: **stretches (ν)** and **bends (δ)**.
 - These two types appear in different energies in IR and Raman spectra.
 - To differentiate between stretching and bending modes: Look at how the stretching modes transform. Each stretch happens along a single vector from central atom to ligand. Thus, if we let these vectors be our basis and look at how many vectors stay the same (+1) or change (-1) under each operation, we will have a reducible representation that can be decomposed. Its components will be the stretching modes and $\Gamma_{\text{vibs}} - \Gamma_{\nu}$ will equal Γ_{δ} .
 - Applied to PF_5 , we have $\Gamma_{\nu} = (5, 2, 1, 3, 0, 3) = 2A'_1 + E' + A''_2$.
 - Thus, both A'_1 representations, one E' representation, and one A''_2 representation correspond to stretching modes (or [spectroscopic] bands; note that the E' band will be twice as large). Consequently, the other two E' representations, the other A''_2 representation, and the E'' representation correspond to bending modes.

II.10 Module 10: IR and Raman Active Vibrations (part 1)

- Today, we will learn what we need to know for PSet 2. Wednesday, we will introduce a more rigorous, quantum mechanical foundation.
- Molecular vibrations can be experimentally observed by **infrared spectroscopy** or **Raman spectroscopy**.
- **Infrared spectroscopy:** A method of measuring the change in dipole moment during a vibration^[4]. *Also known as IR spectroscopy.*

⁴Refer to Labalme (2020a), specifically the discussion of spectroscopy in Chapter 13.

- **Raman spectroscopy:** A method of measuring the change in the polarizability during a vibration. We impinge upon a sample of a substance with an incident light (a powerful laser). Most of the light that scatters will be a **Raleigh scatter**, but some will be a **Raman scatter** (a new wavelength).
- **Raleigh scatter:** The same wavelength is emitted as was impinged by the incident light because the laser was elastically scattered. The electrons fell the same number of energy levels that they rose after absorbing the light.
- **Raman scatter:** A new wavelength is emitted, different than the one impinged by the incident light. The electrons fell fewer (**Stokes Raman scattering**) or more (**Anti-Stokes Raman scattering**) energy levels than they rose after absorbing the light.
- **IR active** (vibration): A vibration that transforms with the same symmetry as the \vec{x} , \vec{y} , and \vec{z} vectors (see the χ tables).
- **Raman active** (vibration): A vibration having the same symmetry as the quadratic x, y, z terms, i.e., x^2, z^2, yz , etc.
- For molecules possessing inversion centers, IR and Raman activity will be mutually exclusive.
- Activity data comes from character tables.
 - All vibration representations that have a linear term are IR active, while those that have a quadratic term are Raman active.
- Continuing with the PF_5 example...
 - Table II.5 tells us that E' and A_2'' vibrations are IR active, and A'_1 , E' , and E'' are Raman active.
 - Thus, since $\Gamma_{\text{vibs}} = 2A'_1 + 3E' + 2A_2'' + E''$, there are $3 + 2 = 5$ IR active bonds (corresponding to $3 \cdot 2 + 2 = 8$ vibrations) and $2 + 3 + 1 = 6$ Raman active bonds (corresponding to $2 + 3 \cdot 2 + 1 \cdot 2 = 10$ vibrations).

II.11 Chapter 4: Symmetry and Group Theory

From Miessler et al. (2014).

1/18:

- **Coincident** (axes): Two identical axes.
 - For example, the C_3 rotation axis of CHCl_3 is **coincident** with the C–H bond axis.
- Snowflakes, which are often planar and have hexagonal symmetry, have a twofold (C_2), threefold (C_3), and sixfold (C_6) axis through their center and perpendicular to their plane (see Figure II.6).
 - Rotations C_3^2 and C_6^5 are also symmetry operations.
- “When necessary, the C_2 axes perpendicular to the principal axis are designated with primes; a single prime (C_2') indicates that the axis passes through several atoms of the molecule, whereas a double prime (C_2'') indicates that it passes between the outer atoms” (Miessler et al., 2014, p. 77).

1/19:

- Even though $S_2 \equiv i$ and $S_1 \equiv \sigma$, the i and σ notations are preferred because of the group theory requirement of maximizing the number of unique classes of symmetry operations associated with a molecule.
- **Point group:** A set of symmetry operations that describes a molecule’s overall symmetry.
- Alternative steps for assigning point groups:
 1. Determine whether the molecule exhibits very low symmetry (C_1, C_s, C_i) or high symmetry ($T_d, O_h, C_{\infty v}, D_{\infty h}, I_h$).



(a) About the principal axis. (b) About perpendicular axes.

Figure II.6: Rotations of a snowflake design.

2. If not, find the highest order C_n axis for the molecule.
 3. Does the molecule have any C_2 axes perpendicular to the principal C_n axis? If it does, there will be n of such C_2 axes, and the molecule is in the D set of groups. If not, it is in the C or S set.
 4. Does the molecule have a mirror plane (σ_h) perpendicular to the principal C_n axis? If so, it is classified as C_{nh} or D_{nh} . If not, continue with Step 5.
 5. Does the molecule have any mirror planes that contain the principal C_n axis (σ_v or σ_d)? If so, it is classified as C_{nv} or D_{nd} . If not, but it is in the D set, it is classified as D_n . If the molecule is in the C or S set, continue with Step 6.
 6. Is there an S_{2n} axis collinear with the principal C_n axis? If so, it is classified as S_{2n} . If not, the molecule is classified as C_n .
- Groups of high symmetry:
 - $C_{\infty v}$ (linear): These molecules are linear, with an infinite number of rotations and an infinite number of reflection planes containing the rotation axis. They do not have a center of inversion.
 - $D_{\infty h}$ (linear): These molecules are linear, with an infinite number of rotations and an infinite number of reflection planes containing the rotation axis. They also have perpendicular C_2 axes, a perpendicular reflection plane, and an inversion center.
 - T_d (tetrahedral): Most (but not all) molecules in this point group have the familiar tetrahedral geometry. They have four C_3 axes, three C_2 axes, three S_4 axes, and six σ_d planes. They have no C_4 axes.
 - Look for C_3 and C_2 axes.
 - O_h (octahedral): These molecules include those of octahedral structure, although some other geometrical forms, such as the cube, share the same set of symmetry operations. Among their 48 symmetry operations are four C_3 rotations, three C_4 rotations, and an inversion.
 - Look for C_4 , C_3 , and C_2 axes.
 - I_h (icosahedral): Icosahedral structures are best recognized by their six C_5 axes, as well as many other symmetry operations — 120 in all.
 - Look for C_5 , C_3 , and C_2 axes.
 - T_h : Adds i to T_d . Example: $\text{W}[\text{N}(\text{CH}_3)_2]_6$.

1/24:

- When we have a block-diagonalized reducible representation, the irreducible representations can be obtained from the individual blocks.
 - This is because we will find that an analogous set of blocks satisfies the multiplication table for the group.

- For the C_{2v} point group, we have the following reducible representation that can be applied to each coordinate.

$$E : \begin{bmatrix} 1 & 0 & 0 \\ 0 & 1 & 0 \\ 0 & 0 & 1 \end{bmatrix} \quad C_2 : \begin{bmatrix} -1 & 0 & 0 \\ 0 & -1 & 0 \\ 0 & 0 & 1 \end{bmatrix} \quad \sigma_v(xz) : \begin{bmatrix} 1 & 0 & 0 \\ 0 & -1 & 0 \\ 0 & 0 & 1 \end{bmatrix} \quad \sigma_v(yz) : \begin{bmatrix} -1 & 0 & 0 \\ 0 & 1 & 0 \\ 0 & 0 & 1 \end{bmatrix}$$

- The above matrices are block diagonalized. Thus, the sets of a_{11} , a_{22} and a_{33} form three irreducible representations (and we can confirm that these obey the C_{2v} group multiplication table).

$$\begin{array}{llll} E : [1] & C_2 : [-1] & \sigma_v(xz) : [1] & \sigma_v(yz) : [-1] \\ E : [1] & C_2 : [-1] & \sigma_v(xz) : [-1] & \sigma_v(yz) : [1] \\ E : [1] & C_2 : [1] & \sigma_v(xz) : [1] & \sigma_v(yz) : [1] \end{array}$$

- Three of the C_{2v} irreducible representations can be found in this way. The fourth can be found w/ linear algebra and the character table properties (or by inspection).

- If one matrix in the reducible representation cannot be blocked past having a 2×2 chunk, per se, then the others must also have that 2×2 chunk (so that we have an irreducible 2×2 representation).

- “Matching the symmetry operations of a molecule with those listed in the top row of the character table will confirm any point group assignment” (Miessler et al., 2014, p. 99).

- Symmetric:** Character of 1.

- Antisymmetric:** Character of -1.

- Mulliken symbol rules:

1. “Letters are assigned according to the dimension of the irreducible representation” (Miessler et al., 2014, p. 99).

- Assign *A* if the dimension is 1 and the representation is symmetric to the principal rotation operation ($\chi(C_n) = 1$).
- Assign *B* if the dimension is 1 and the representation is antisymmetric to the principal rotation operation ($\chi(C_n) = -1$), or if $\chi(S_{2n}) = -1$ even if $\chi(C_n) = 1$.
- Assign *E* if the dimension is 2.
- Assign *T* if the dimension is 3.

2. “Subscript 1 designates a representation symmetric to a C_2 rotation perpendicular to the principal axis, and subscript 2 designates a representation antisymmetric to the C_2 . If there are no perpendicular C_2 axes, 1 designates a representation symmetric to a vertical plane, and 2 designates a representation antisymmetric to a vertical plane” (Miessler et al., 2014, p. 100).

3. “Subscript *g* designates representations symmetric to inversion, and subscript *u* designates representations antisymmetric to inversion” (Miessler et al., 2014, p. 100).

4. “Single primes are symmetric to σ_h and double primes are antisymmetric to σ_h when a distinction between representations is needed (C_{3h} , C_{5h} , D_{3h} , D_{5h})” (Miessler et al., 2014, p. 100).

1/26:

- Chiral** (molecule): A molecule that is not superimposable on its mirror image. *Also known as disymmetric.*

- “In general, a molecule or object is chiral if it has no symmetry operations (other than *E*), or if it has only proper rotation axes” (Miessler et al., 2014, p. 100).

- Raman spectroscopy makes use of higher energy radiation than IR, exciting molecules to higher electronic states that are envisioned as short-lived “virtual” states.

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

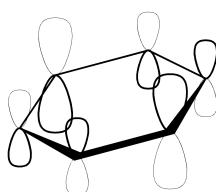
$$\begin{array}{ll} \Psi_1(A) = \frac{1}{\sqrt{6}} (\phi_1 + \phi_2 + \phi_3 + \phi_4 + \phi_5 + \phi_6) & \Psi_2(B) = \frac{1}{\sqrt{6}} (\phi_1 - \phi_2 + \phi_3 - \phi_4 + \phi_5 - \phi_6) \\ \Psi_3(E_1) = \frac{1}{\sqrt{12}} (2\phi_1 + \phi_2 - \phi_3 - 2\phi_4 - \phi_5 + \phi_6) & \Psi_4(E_1) =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



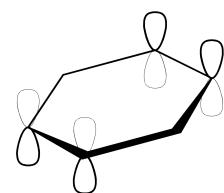
(a) $\Psi_1(A) \sim \Psi(A_{2u})$.



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



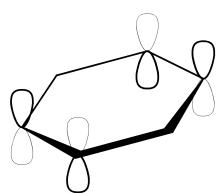
(c) $\Psi_3(E_1) \sim \Psi(E_{1g}^a)$.



(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).

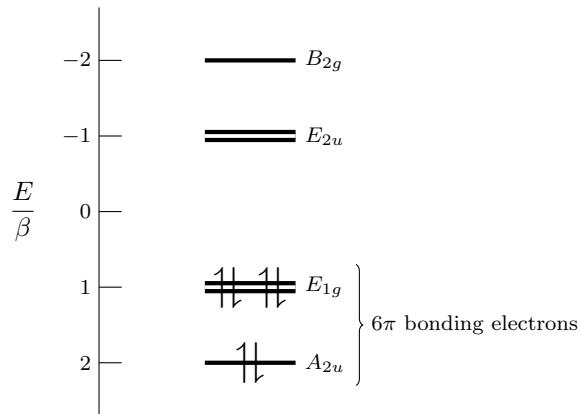


Figure III.2: Energy level diagram of benzene.

- 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.

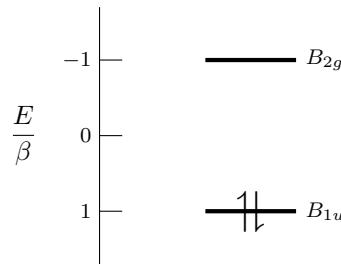


Figure III.3: Energy level diagram of ethene.

- 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 (2020a), 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.

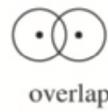
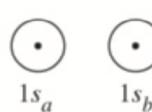
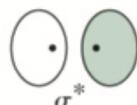
- The bonding orbital(s) will reside predominantly on the atom of lower orbital energy (the more electronegative atom).
- The anti-bonding orbital(s) will reside predominantly on the atom with greater orbital energy (the less electronegative atom).
- The energies of atomic orbitals (measured by PES) 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?

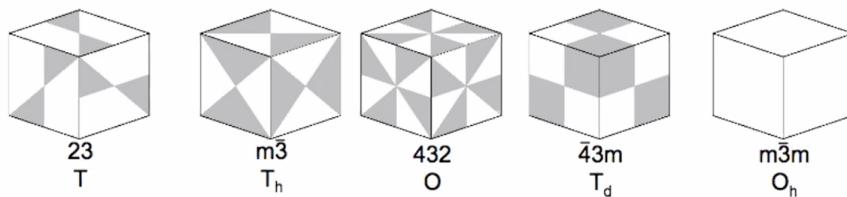


Figure III.9: Tetrahedral point groups.

- 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.

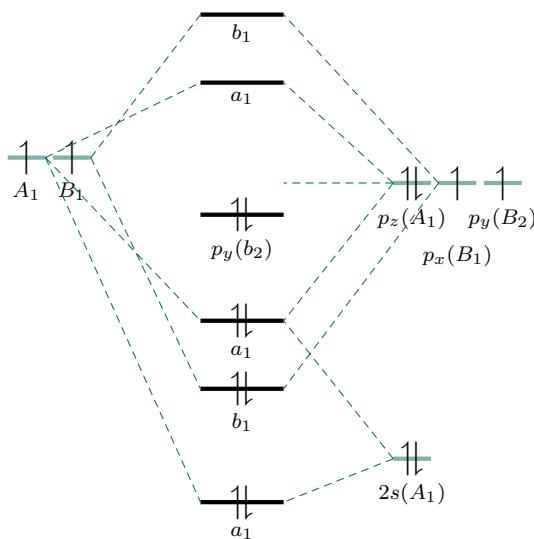
- We can now fill the MOs with the atomic electrons using the Aufbau principle, the Pauli exclusion principle, and Hund's rule. We can also draw these orbitals.
- Orbital energies can be found in Table 5.2 of Miessler et al. (2014).
- If the energy difference between orbitals is more than 10 eV, then 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 B₁ 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):

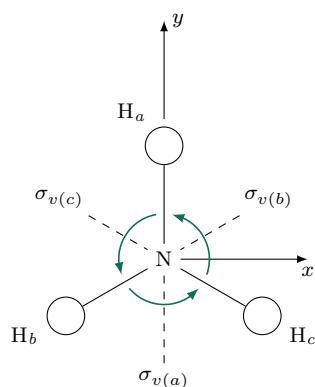


Figure III.13: Coordinate system for NH₃.

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 A₁ type atomic orbitals form 3 a₁ type MOs.



Figure III.14: NH₃ orbital diagram.

- Since 2s(A₁) and 1s(A₁) are so far away energetically, their combination will have very low energy.
- On the other hand, since the 1s(A₁) and p_z(A₁) 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 A₁ 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:

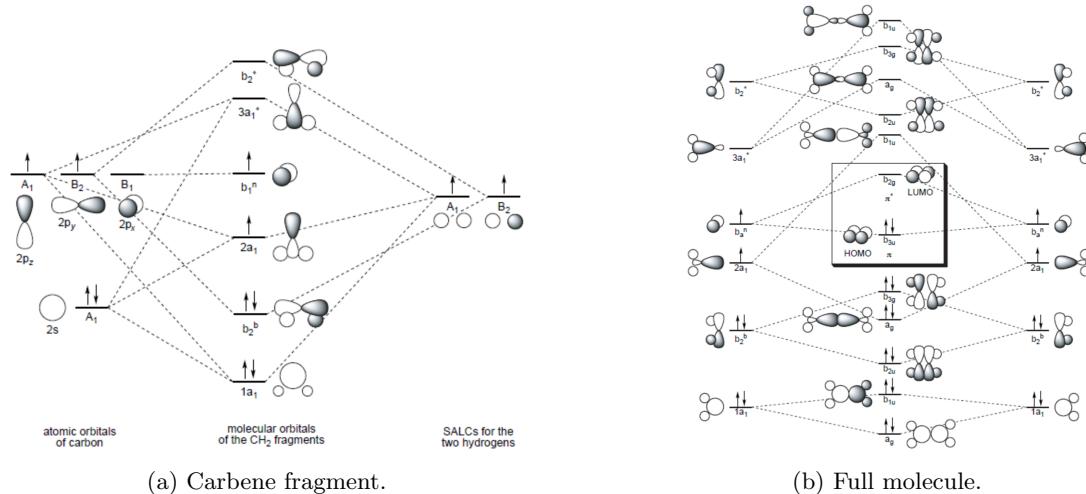


Figure III.15: $\text{H}_2\text{C}=\text{CH}_2$ orbital diagram.

- Two different approaches (D_{2h}): $\text{C}_1 + \text{C}_2$ then H_{1-4} , or two carbene fragments (CH_2).
 - We can read about the first approach in the linked JCE article (*cite*).
- C_{2v} for the CH_2 fragment.
- When we make the MO diagram, we have two unpaired electrons, one in each of the upper two orbitals. These will be used for bonding to the other carbene fragment.
- We will then combine the MO diagrams and will need to reassign orbital names because ethene isn't C_{2v} ; it's D_{2h} .

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^* .

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 $\sigma_g(2s)$ and $\sigma_g(2p)$ 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 $\sigma_g(2s)$ and $\sigma_g(2p)$ orbitals in Figure III.17 can mix, lowering the energy of the 2s one and raising the energy of the 2p one.
- 2/6: • Atomic orbital energies decrease across a row in the periodic table.
- Characterization of selected homonuclear diatomic molecules:

- He₂’s bond order is 0, with two bonding and two antibonding electrons. Thus, it has no significant tendency to form (it’s bond energy is 0.01 J/mol, compared to H₂’s 436 kJ/mol).
- B₂ is paramagnetic (because of its two unpaired π_u electrons). The 2p valence electrons do not occupy the σ_g MO because of orbital mixing and because the energy difference between $\sigma_g(2p)$ and $\pi_u(2p)$ is greater than Π_c (even if there were no mixing, the electrons would still occupy the π_u orbitals if it was more energetically favorable to do this than to occupy the same orbital).
- C₂ is a rarely encountered allotrope of carbon, but it has two π bonds and no σ bonds.
- In N₂, some larger trends become apparent:
 - Since electrons in different orbitals vary in their shielding abilities and electron-electron interactions, the difference between the 2s and 2p energies increase as Z increases.
 - More specifically, 2s electrons have higher probabilities close to the nucleus than 2p electrons (Figure 0.8), so they are more affected by increasing nuclear charge (inverse square law).
 - As a result of this separation, the $\sigma_g(2s)$ and $\sigma_g(2p)$ orbitals mix less in N₂ than in C₂ or B₂.
- The ions of O₂ (which include dioxygenyl O₂⁺, superoxide O₂⁻, and peroxide O₂²⁻) reveal that bond order and bond distance are inversely proportional. Here, mixing finally becomes small enough that the normal filling order (as in Figure III.17) returns.
- Notice how atomic radius consistently decreases across the second period, but bond length decreases until N₂ and then increases (because of the addition of antibonding electrons).
- At this point, we’ll assume (for PES purposes) the energy levels in the uncharged molecule to be essentially the same as those in the charged ions generated when we seek to examine core energy levels^[3].
- Comparing the PES spectrums for N₂ and O₂, it can be seen that the N₂ energy levels are closer together than the O₂ ones.
- PES also provides evidence for the existence of **vibrational energy levels**, energy levels that are much more closely spaced than electronic levels that cause the multiple peaks within one “peak.”
- Orbitals that are strongly involved in bonding have **vibrational fine structure** (multiple peaks), and vice versa for less involved orbitals.
 - Thus, the 10 π_u vs. the 6 σ_g vibrational peaks in O₂ indicate that the π_u orbitals are more strongly involved in bonding than the σ_g orbital.
- “The atomic orbitals of atoms that form homonuclear diatomic molecules have identical energies, and both atoms contribute equally to a given MO. Therefore, in the molecular orbital equations, the coefficients associated with the same atomic orbitals of each atom...are identical. In heteronuclear diatomic molecules...the atomic orbitals have different energies, and a given MO receives unequal contributions from these atomic orbitals; the MO equation has a different coefficient for each of the atomic orbitals that contribute to it” (Miessler et al., 2014, pp. 134, 136).
- Atomic orbitals closer in energy to an MO contribute more to that MO, and thus have larger coefficients in the wave equation.
- CO example:
 - The 1 π orbitals are lower than the 3 σ orbitals because of the strong mixing between the 2p_z orbital of oxygen and the 2s and 2p_z orbitals of carbon.
 - The 3 σ orbital has a large lobe on the carbon end because two carbon orbitals (2s and 2p_z) mix with one oxygen orbital (2p_z).

³However, we should note that this is an oversimplification; a rigorous treatment of PES considers how the energy levels and orbital shapes vary between the neutral and ionized species.

- The π orbital has electron density concentrated on oxygen because of the better energy match between the MO and the $2p_{x,y}$ orbitals of oxygen; the π^* orbital has electron density concentrated on carbon because of the better energy match between the MO and the $2p_{x,y}$ orbitals of carbon.
- Atomic orbitals with energy differences greater than 10-14 eV usually do not interact significantly.
- MOs in HF predict a polar bond since electron density is concentrated on F to such a greater extent in every occupied orbital.
- **Frontier orbitals:** The HOMO and LUMO, so-named because they lie at the occupied-unoccupied frontier.
- Frontier orbitals help explain reaction chemistry with transition metals.
 - In CO for example, we'd predict based on electronegativity that the O would be more reactive, and hence we'd get a preponderance of M–O–C structures.
 - However, carbonyl complexes are typically of the form M–C–O because the frontier orbitals both lie more on C. Indeed, the HOMO contains the least stabilized (most reactive) electrons in the molecule, and the LUMO is ready to accept whatever electrons are donated first.
- Ionic compounds can still be treated, MO theory-wise, like covalent compounds; we just get MOs that are almost identical to the more favored constituent atomic orbitals.
- Crystalline lattice salts are much more stable than diatomic ones.
 - Such crystal lattices are held together by a combination of electrostatics (ionic) attraction and covalent bonding.
 - Salts do not exhibit directional bonds; instead, the orbitals form energy bands (see Chapter 7).
- Considers lattice enthalpies and Born-Haber cycles^[4].
- **Group orbitals:** Collections of matching orbitals on outer atoms.
 - Sets of orbitals that potentially could interact with the central atom orbitals.
 - The same combinations that formed bonding and antibonding orbitals in diatomics.
- FHF⁻ example:
 - The group orbitals are formed by adding and subtracting the F orbitals as we would in F₂.



Figure III.18: Interaction of fluorine group orbitals with the hydrogen 1s orbital.

- The only two group orbitals eligible for bonding with the H(1s) orbital based on symmetry are $2s + 2s$ and $2p_{z_a} - 2p_{z_b}$. Having the H orbital with the same sign as the surrounding lobes gives bonding orbitals, and the opposite sign generates antibonding orbitals.

⁴See Labalme (2020a), specifically Figure 8.2 and the associated discussion

- Since the H(2s) orbitals match better energetically with the F(2p_z) orbitals than the F(2s) orbitals, the 2p_z interactions will be stronger.
 - Notice that the lone pairs are delocalized over both fluorine atoms, not confined to one as the Lewis dot model predicts.
 - Additionally, the Lewis approach predicts two, 2-electron bonds, resulting in 4 electrons around the central H atom. However, the MO model suggests 2 electrons (in a σ bond [a_g]) decentralized over all three atoms.
- A stepwise approach to building MOs for more complex molecules (Miessler et al., 2014, p. 143):
1. Determine the point group of the molecule. If it is linear, substituting a simpler point group that retains the symmetry of the orbitals (ignoring the wave function signs) makes the process easier. It is useful to substitute D_{2h} for $D_{\infty h}$ and C_{2v} for $C_{\infty v}$. This substitution (known as a **Descent in Symmetry**) retains the symmetry of the orbitals without the need to use infinite-fold rotation axes.
 2. Assign x, y, z coordinates to the atoms, chosen for convenience. Experience is the best guide here. A general rule is that the highest order rotation axis of the molecule is assigned as the z -axis of the central atom. In nonlinear molecules, the y -axes of the outer atoms are chosen to point toward the central atom.
 3. Construct a (reducible) representation for the combination of the valence s orbitals on the outer atoms. If the outer atom is not hydrogen, repeat the process, finding the representations for each of the other sets of outer atom orbitals (for example, p_x , p_y , and p_z). As in the case of the vectors described in Chapter 4, any orbital that changes position during a symmetry operation contributes 0 to the character of the resulting representation; any orbital that remains in its original position — such as a p orbital that maintains its position and direction (signs of its orbital lobes) — contributes 1; and any orbital that remains in the original position, with the signs of its lobes reversed, contributes -1.
 4. Reduce each representation from Step 3 to the sum of its irreducible representations. This is equivalent to finding the symmetry of the group orbitals or the **symmetry-adapted linear combinations** (SALCs) of the orbitals. The group orbitals are then the combinations of atomic orbitals that match the symmetry of the irreducible representations.
 5. Identify the atomic orbitals of the central atom with the same symmetries (irreducible representations) as those found in Step 4.
 6. Combine the atomic orbitals of the central atom and those of the group orbitals with matching symmetry and similar energy to form molecular orbitals. The total number of molecular orbitals formed must equal the number of atomic orbitals used from all the atoms. Note that the MOs are assigned lowercase Mulliken symbols (e.g., a_1), whereas atomic orbitals and representations in general are assigned uppercase Mulliken symbols (e.g., A_1).
- CO₂ example:
- We find the same group orbitals as in FHF⁻ but by using the stepwise procedure.
 - Note that saying $\Gamma_{2s} = A_g + B_{1u}$, for example, expresses the fact that there are two 2s orbitals (the sum and difference of the O(2s) orbitals) and that one has A_g symmetry while the other has B_{1u} symmetry.
 - The orbitals on the central atom are 2s, 2p_z, 2p_x, 2p_y.
 - The A_g O(2s) orbital can be added to and subtracted from the C(2s) orbital, and the B_{1u} O(2s) orbital can be added to and subtracted from the C(2p_z) orbital.
 - However, due to the large energy difference between the O and C orbitals, these do not generally form and are not included in Figure III.19.
 - Instead, the two O(2s) orbitals are included as σ 2a_g and 2b_{1u}, respectively, nonbonding orbitals.

Figure III.19: Molecular orbitals for CO_2 .

- The A_g O($2p_z$) orbital can be added to and subtracted from the C($2s$) orbital, and the B_{1u} O($2p_z$) orbital can be added to and subtracted from the C($2p_z$) orbital.
 - From the first interaction: The σ $3a_g$ bonding orbital formed is the most stable, with only two nodes and an uninterrupted probability region between the three nuclei. The σ^* $4a_g$ antibonding orbital is the second least stable with four nodes.
 - From the second interaction: The σ $3b_{1u}$ bonding orbital formed is the second most stable, with only three nodes. The σ^* $4b_{1u}$ antibonding orbital is the least stable with a whopping five nodes.
- The B_{2u} O($2p_y$) orbital can be added to and subtracted from the C($2p_y$) orbital, yet the B_{3g} O($2p_y$) orbital does not have matching symmetry with any C orbital.
 - From the first interaction: The π $1b_{2u}$ bonding orbital formed is less stable than the σ $3b_{1u}$ orbital since it lacks such direct probability between the nuclei but also has few nodes. The π^* $2b_{2u}$ antibonding orbital is more stable than the σ^* $4a_g$ antibonding orbital since it has fewer nodes.

- From the second interaction: The $\pi 1b_{3g}$ nonbonding orbital formed resides in the middle of the energy diagram.
- The B_{3u} O($2p_x$) orbital can be added to and subtracted from the C($2p_x$) orbital, yet the B_{2g} O($2p_x$) orbital does not have matching symmetry with any C orbital.
- The cases are symmetric to the previous two.
- This process may be used to obtain numerical values for the coefficients of the atomic orbitals, but the computational methods are beyond the scope of Miessler et al. (2014).
- H₂O example:
 - We canonically select that xz -plane as the plane of the molecule when we have a choice.
 - Because the H($1s$) orbitals have no directionality, it is not necessary to assign coordinate axes to the hydrogen atoms.

Symmetry	Molecular Orbitals	=	Oxygen Atomic Orbitals	Group Orbitals from Hydrogen Atoms	Description
B_1	Ψ_6	=	$c_9\psi(p_x)$	$+ c_{10}[\psi(H_a) - \psi(H_b)]$	Antibonding (c_{10} is negative)
A_1	Ψ_5	=	$c_7\psi(s)$	$+ c_8[\psi(H_a) + \psi(H_b)]$	Antibonding (c_8 is negative)
B_2	Ψ_4	=	$\psi(p_y)$		Nonbonding
A_1	Ψ_3	=	$c_5\psi(p_z)$	$+ c_6[\psi(H_a) + \psi(H_b)]$	Slightly bonding (c_6 is small)
B_1	Ψ_2	=	$c_3\psi(p_x)$	$+ c_4[\psi(H_a) - \psi(H_b)]$	Bonding (c_4 is positive)
A_1	Ψ_1	=	$c_1\psi(s)$	$+ c_2[\psi(H_a) + \psi(H_b)]$	Bonding (c_2 is positive)

Table III.2: Molecular orbital equations for H₂O.

- Note that the nonbonding pairs afforded by the MO model are not equivalent as in the Lewis model.
- NH₃ example:
 - Is it not a bit circular to assign the point group C_{3v} and then later use Walsh diagrams to determine that it has a C_{3v} structure?
 - Here, we can no longer just add and subtract atomic orbitals. Instead, we must apply “the projection operator method, a systematic approach for deduction of group orbitals” (Miessler et al., 2014, p. 152).
 - This method illustrates how atomic orbitals should be combined to afford the SALCs that define the group orbitals.
 - First, determine the impact of each point group symmetry operation (individually, not within classes) on one atomic orbital.
 - Linear combinations of these atomic orbitals that match the symmetries of the group’s irreducible representations can be obtained via (1) multiplication of each outcome by the characters associated with each operation for these irreducible representations, followed by (2) addition of the results.
 - Finding SALC(A_1) = $\frac{1}{\sqrt{3}}[\Psi(H_a) + \Psi(H_b) + \Psi(H_c)]$ and SALC(E) = $\frac{1}{\sqrt{6}}[2\Psi(H_a) - \Psi(H_b) - \Psi(H_c)]$ is easy.
 - Now SALC(E) has y symmetry (note that this means that choosing the H_a that we did in Figure III.13 [i.e., the one lying along a coordinate axis] has simplified calculations). Thus, the other SALC should have x symmetry.

– H_a should not contribute to the new SALC because of the orthogonal node defined by the yz -plane (if it contributed, there would be no yz node). Thus, only H_b and H_c can contribute, and we can deduce that their coefficients are $\frac{1}{\sqrt{2}}$ and $-\frac{1}{\sqrt{2}}$, respectively, to satisfy the normalization requirement while also leading to identical total contributions from all three $1s$ wave functions across the three group orbitals.

- In the projection operator method, orbitals can become the inverses of other orbitals.
- BF_3 example:

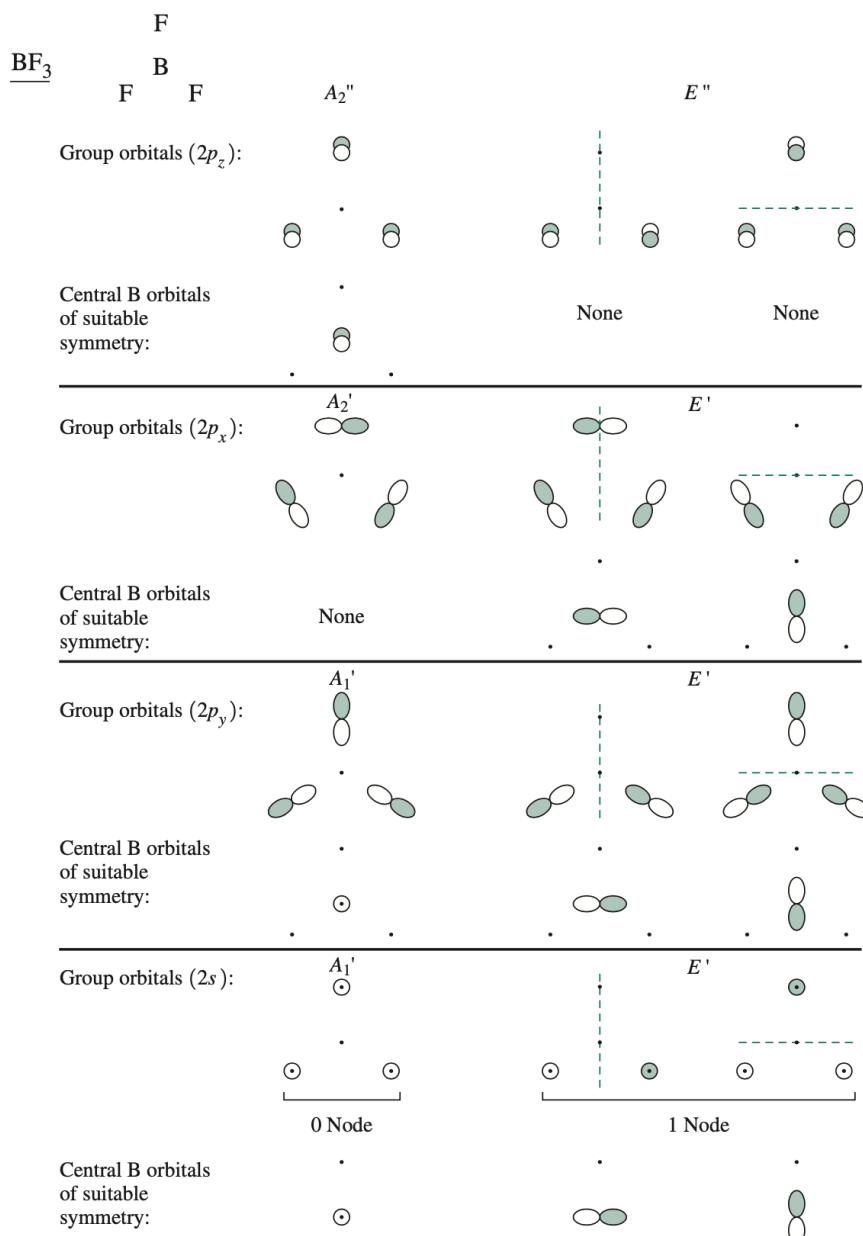


Figure III.20: Group orbitals for BF_3 .

- The fluorine ligands have $2p$ electrons, too, now. As such, let the C_3 axis be the z -axis, the p_y axes point towards the central boron atom, and the p_x axes lie in the molecular plane.

- The partial double bonding character of the B–F bonds discussed in Chapter 3 can now be explained.
- The LUMO of BF_3 is an empty antibonding π orbital with large lobes on boron that can act as electron pair acceptors; this is why BF_3 is a Lewis acid.
- This approach is also applicable to other isoelectronic trigonal planar species such as SO_3 , NO_3^- , and CO_3^{2-} .
- “Because the extent of orbital overlap in π interactions is generally less than that in most σ interactions, a double bond composed of one filled σ orbital and one filled π orbital is not twice as strong as a single bond” (Miessler et al., 2014, p. 161).
- Note that single bond energies between the same atoms can vary widely based on steric crowding and adjacent bonding.
- A qualitative group orbital approach (what we’re doing) does not allow for the determination of the precise MO energies, but it generally allows for the placement of the MOs in the approximate order based on their shapes and expected orbital overlaps.
- “In the hybrid concept, the orbitals of the central atom are combined into sets of equivalent hybrids. These hybrid orbitals form bonds with orbitals of other atoms” (Miessler et al., 2014, p. 161).
- Hybrid orbitals correctly describe methane’s PES spectrum.
- “Like all bonding models, hybrids are useful so long as their limits are recognized” (Miessler et al., 2014, p. 161).
- CH_4 example:
 - Point group: T_d .
 - Basis functions: σ bond vectors.
 - $\Gamma = (4, 1, 0, 0, 2)$.
 - $\Gamma = A_1 + T_2$.
 - “The atomic orbitals of carbon used in the hybrids must have symmetry matching $A_1 + T_2$; more specifically, one orbital must match A_1 , and a set of three (degenerate) orbitals must match T_2 ” (Miessler et al., 2014, p. 162).
 - For A_1 , we choose the $2s$ orbital. For T_2 , we could choose d_{xy}, d_{xz}, d_{yz} , but since they’re much higher energy (and thus won’t mix well with the $\text{H}(1s)$ orbitals), we’ll go for p_x, p_y, p_z .
 - Therefore, the hybridization is sp^3 , combining four atomic orbitals into four equivalent hybrid orbitals, one directed toward each hydrogen atom.
- In addition to the sp^3 explanation of H_2O , it is sometimes explained through a slightly more MO theory lens as sp^2 since the oxygen orbitals used in MO bonding are $2s, 2p_x, 2p_z$ where the H_2O molecule lies in the xz -plane.
- Only σ bonding is considered when determining the orbitals used in hybridization.

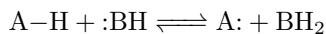
Topic IV

Hard-Soft Acid-Base and Donor-Acceptor Concepts of Transition Metals

IV.1 Module 24: Acid-Base Chemistry

2/8:

- Brønsted-Lowry Acid-Base Theory of Acids and Bases (1923).
 - Acid: Any chemical species (molecule or ion) that is able to lose, or “donate,” a hydrogen ion (proton).
 - Base: Any chemical species that is able to gain, or “accept”, a proton.
 - A base must have a pair of electrons available to share with the proton; this is usually present as an unshared pair, but sometimes is in a π orbital.
 - Acid-base reactions: The transfer of a proton from an acid to a base.



- Protons do not exist free in solution but must be attached to an electron pair.
- Water is amphoteric.
- In the Brønsted-Lowry paradigm, we cannot separate the acids/bases from the solvent (no protons; only *solvated* protons). In a non-aqueous medium such as DMSO, however, we have much broader scope of acids and bases.
- **Carbon acid:** Any molecule containing a C–H bond can lose a proton forming the carbanion.
- Carborane ($\text{H}(\text{CHB}_{11}\text{Cl}_{11})$) is a superacid one million times stronger than sulfuric acid since its conjugate basis is incredibly stable (super easy to delocalize the charge).
- The base dissociation constant or K_b is a measure of basicity. $\text{p}K_b$ is the negative log of K_b and related to the $\text{p}K_a$ by the simple relationship $\text{p}K_a + \text{p}K_b = 14$. The larger the $\text{p}K_b$, the more basic the compound.
- **Superacid:** An acid with acidity greater than that of 100% pure sulfuric acid.
 - In water, the strongest acid you can have is H_3O^+ .
 - The strongest superacids are prepared by the combination of two components, a strong Lewis acid and a strong Brønsted-Lowry acid.
 - Fluoroantimonic acid HF-SbF_5 is 2×10^{19} stronger than 100% sulfuric acid.

- Olah's magic acid ($\text{FSO}_3\text{H}-\text{SbF}_5$) can dissolve paraffin (candle wax; extremely inert), converting methane into the t-butyl carbocation.

- **Hammett acidity function:** Can replace the pH in concentrated solutions. *Also known as H_0 .*

$$H_0 = \text{p}K_{\text{BH}^+} + \log \frac{[\text{B}]}{[\text{BH}^+]}$$

- Let BH^+ be the conjugate acid of a very weak base B, with a very negative $\text{p}K_{\text{BH}^+}$. In this way, it is rather as if the pH scale has been extended to very negative values.
- Hammett originally used a series of anilines with EWGs for the bases.

- **Superbase:** A compound that has a high affinity for protons.

- Again, these do not exist in water.
- Often destroyed by water, CO_2 , and O_2 .
- A superbase has been defined as an organic compound whose basicity is greater than that of proton sponge, which has conjugate $\text{p}K_a$ of 12.1.
- These are valuable in organic chemistry, which abounds in very weak acids.
- A common superbase is lithium diisopropylamide.

IV.2 Module 25: Acid-Base Chemistry: Lewis Acids and Bases

2/10:

- The second midterm will be Saturday, February 20.
- Talapin will hold office hours this Friday from 3:00-4:00 PM.
- **Lewis acid:** Any species with a vacant orbital.
 - An atomic or molecular species that has an empty atomic or molecular orbital of low energy (LUMO) that can accommodate a pair of electrons.
- **Lewis base:** A compound with an available pair of electrons, either unshared or in a π orbital.
 - An atomic or molecular species that has a lone pair of electrons in the HOMO.
- Examples of Lewis acids:
 - The proton, onium ions (e.g., NH_4^+ , H_3O^+), metal cations (e.g., Li^+ , Mg^{2+}), trigonal planar species (e.g., BF_3 , CH_3^+), and electron-poor π -systems (e.g., enones, tetracyanoethylene [TCNE]).
- **Lewis base adduct:** The product of a Lewis acid reaction.
 - Simple Lewis acids:
 - Examples include BF_4^- in $\text{BF}_3 + \text{F}^- \longrightarrow \text{BF}_4^-$ and BF_3OMe_2 in $\text{BF}_3 + \text{OMe}_2 \longrightarrow \text{BF}_3\text{OMe}_2$, which are Lewis base adducts of BF_3 .
 - In many cases, Lewis base adducts violate the octet rule.
 - In some cases, Lewis acids can bind to two Lewis bases (e.g., $\text{SiF}_4 + 2\text{F}^- \longrightarrow \text{SiF}_6^{2-}$).
 - Complex Lewis acids:
 - Most compounds considered to be Lewis acids require an activation step prior to the formation of the adduct.
 - For example, the reaction $\text{B}_2\text{H}_6 + 2\text{H}^- \longrightarrow 2\text{BH}_4^-$ goes through the intermediate B_2H_7^- .
- The proton (H^+) is one of the strongest but is also one of the most complicated Lewis acids. It is convention to ignore the fact that a proton is heavily solvated (bound to solvent). With this simplification in mind, acid-base reactions can be viewed as the formation of adducts.

- For example, $\text{H}^+ + \text{NH}_3 \longrightarrow \text{NH}_4^+$ and $\text{H}^+ + \text{OH}^- \longrightarrow \text{H}_2\text{O}$.
- In the first reaction, essentially what happens is the frontier orbitals of H^+ and NH_3 are very similar in energy, and thus strongly combine when they form bonding and antibonding orbitals. Two electrons previously in the HOMO of NH_3 drop the bonding energy into the new lower MO of NH_4^+ . More rigorously, we have to account for the change in symmetry group from C_{3v} to T_d , but the above description basically encapsulates what happens.
- Examples of Lewis bases:
 - $\text{NH}_{3-x}\text{R}_x$ where R = alkyl or aryl.
 - $\text{PR}_{3-x}\text{A}_x$ where R = alkyl and A = aryl.
 - Compounds of O, S, Se, and Te in oxidation state 2, including water, ethers, and ketones.
 - Simple anions (e.g., H^- , F^-), other lone pair-containing species (e.g., H_2O , NH_3 , OH^- , CH_3^-), complex anions (e.g., sulfate), and electron-rich π -systems (e.g., ethyne, ethene, benzene).
- To quantify the strength of Lewis acids and bases, compare the standard enthalpies of complexation in kJ/mol.
 - For example, heats of binding of Lewis bases to BF_3 .
 - We do this because we cannot use $\text{p}K_b$ in the same way as we can in the Brønsted-Lowry description.
- **Electrides:** Compounds where the anions are electrons.
 - Put an alkali metal in liquid ammonia; the deep blue color comes from solvated electrons floating freely in solution surrounded by ammonia molecules.
 - If you add a very strong complexing agent for alkali metal ions to the solution, then the electrons will not be able to react back with the alkali metal and we will have crystallized an electride.
 - Electrides are dielectrics.
- The power of the Lewis approach is that it allows us to classify many reactions as acid-base reactions.
 - For example, Wilkinson's catalyst ($\text{Rh}(\text{PPh}_3)\text{Cl}$ where Ph is a phenyl group) is probably the first hydrogenation catalyst to be explained.
 - The Lewis approach allows us to view every step in the mechanism^[1] as an acid-base reaction.

IV.3 Module 26: Hard and Soft Acids and Bases (HSAB) Principle

- HSAB principle can't be derived, but it's been proven to be a very powerful tool for chemists.
- The affinity of hard acids and hard bases for each other is mainly ionic in nature, whereas the affinity of soft acids and bases for each other is mainly covalent in nature.
- Hard acids and hard bases tend to have:
 - Small atomic/ionic radius.
 - High oxidation state.
 - Low polarizability.
 - High electronegativity.
 - Energy low-lying HOMO (bases) or energy high-lying LUMO (acids).

¹We will study this mechanism in depth in CHEM 20200.

- Examples: H^+ , alkali ions, OH^- , F^- .
- Small stabilization energy when bonding, so attraction is mostly electrostatic.
- Soft acids and soft bases tend to have:
 - Large atomic/ionic radius.
 - Low or zero oxidation state.
 - High polarizability.
 - Low electronegativity.
 - Energy high-lying HOMO (bases) or energy low-lying LUMO (acids).
 - Examples: CCH_3Hg^+ , Pt^{2+} , H^- , R_3P , I^- .
 - Large stabilization energy when bonding, so attraction is mostly covalent.
- Hard acids don't readily react with soft bases and vice versa.
- You should be able to intuitively classify acids/bases as hard or soft.

¹ H															
³ Li	⁴ Be														
¹¹ Na	¹² Mg														
¹⁹ K	²⁰ Ca	²¹ Sc	²² Ti	²³ V	²⁴ Cr	²⁵ Mn	²⁶ Fe	²⁷ Co	²⁸ Ni	²⁹ Cu	³⁰ Zn	³¹ Ga	³² Ge	³³ As	
³⁷ Rb	³⁸ Sr	³⁹ Y	⁴⁰ Zr	⁴¹ Nb	⁴² Mo	⁴³ Te	⁴⁴ Ru	⁴⁵ Rh	⁴⁶ Pd	⁴⁷ Ag	⁴⁸ Cd	⁴⁹ In	⁵⁰ Sn	⁵¹ Sb	
⁵⁵ Cs	⁵⁶ Ba		⁷² Hf	⁷³ Ta	⁷⁴ W	⁷⁵ Re	⁷⁶ Os	⁷⁷ Ir	⁷⁸ Pt	⁷⁹ Au	⁸⁰ Hg	⁸¹ Ti	⁸² Pb	⁸³ Bi	
⁸⁷ Fr	⁸⁸ Ra														

hard soft intermediate

⁵⁷ La	⁵⁸ Ce	⁵⁹ Pr	⁶⁰ Nd	⁶¹ Pm	⁶² Sm	⁶³ Eu	⁶⁴ Gd	⁶⁵ Tb	⁶⁶ Dy	⁶⁷ Ho	⁶⁸ Er	⁶⁹ Tm	⁷⁰ Yb	⁷¹ Lu
⁸⁹ Ac	⁹⁰ Th	⁹¹ Pa	⁹² U	⁹³ Np	⁹⁴ Pu	⁹⁵ Am	⁹⁶ Cm	⁹⁷ Bk						

Figure IV.1: Hard vs. soft metal ions.

- The affinity that metal ions have for ligands is controlled by size, charge, and electronegativity.
- This can be refined further by noting that for some metal ions, their chemistry is dominated by size and charge, while for others it is dominated by their electronegativity. These two categories of metal ions have been termed by Pearson as **hard** (metal ions) and **soft** (metal ions).
- Soft Lewis acids are transition metals with unusually high electronegativities (these go against typical periodic trends).
 - Note that QMech can't explain this island of high electronegativities in the transition metals; we need relativistic corrections and Dirac equations.

- The higher electronegativity stabilizes the LUMO of the acid.
- These ions are a treasure trove for catalytic chemistry.
- **Chemical hardness:** One half the difference between the ionization potential I and the electron affinity A .

$$\eta = \frac{I - A}{2}$$

- Note that the above equation cannot be applied to anions because electron affinity cannot be measured for them; the assumption is made that η for an anion X^- is same as that for the radical $X\cdot$.
- Related to the **Mulliken electronegativity**:

$$\chi = \frac{I + A}{2}$$

- **HSAB principle:**

- (i) Hard acids prefer to bond to hard bases, and soft acids prefer to bond to soft bases.
 - Note that this rule has nothing to do with acid or base strength but merely says that the product $A-B$ will have extra stability if A and B are both hard or both soft.
- (ii) A soft Lewis acid and a soft Lewis base tend to form a covalent bond, while a hard acid and a hard base tend to form ionic bonds.
- Solubility: Hard solvents (e.g., HF, H_2O , and the protic solvents) tend to solvate strong solute bases (e.g., F^- and the oxygen anions). Dipolar aprotic solvents (e.g., Me_2SO and CH_3COCH_3) are soft solvents with a preference for solvating large anions and soft bases.
 - For example, $LiI + AgF \longrightarrow LiF + AgI$. In this reaction, two mixed reactants recombine exothermically to form a hard-hard product and a soft-soft precipitate. Because of the Lewis definition, we can treat this as an acid-base reaction.

Classification		Hard	Intermediate	Soft	
		F^-	Cl^-	Br^-	I^-
Soft	Ag^+	0.4	3.3	4.7	6.6
Intermediate	Pb^{2+}	1.3	0.9	1.1	1.3
Hard	Fe^{3+}	6.0	1.4	0.5	-

Table IV.1: Hard and soft formation constants.

- As another example, we can see in Table IV.1 that soft-soft and hard-hard ions have higher formation constants ($\log K_1$) than any other combination, and hard-soft ions have lower formation constants than any other combination.
- Coordination chemistry: Numerous experiments have been done to determine the relative ordering of ligands and transition metal ions in terms of their hardness and softness.
- Au (I) is the softest metal ion.
 - Soft to the extent that compounds such as AuF and Au_2O are unknown.
 - Nevertheless, it forms stable compounds with soft ligands such as cyanide.
- Al (III) is a very hard metal ion.
 - It has very high formation constants with both F^- and OH^- ; additionally, it has virtually no affinity in solution for heavier halides such as Cl^- .

- Its solution chemistry is dominated by its affinity for F^- and for ligands with negative O-donors.
- **Symbiosis effect:** Taking a borderline element and reacting it with a soft Lewis base to soften it, or a hard Lewis base to harden it.
 - For example, B^{3+} can be reacted with 3H^- , we make the soft compound BH_3 . If we react it with 3F^- , we make the hard compound 3F^- .
 - Now BH_3 and F^- do not react, but we can further soften it in the reaction $\text{BH}_3 + \text{H}^- \longrightarrow \text{BH}_4^-$.
 - Similarly, BF_3 and H^- don't react, but we can further harden BF_3 by reacting it with F^- to form BF_4^- .
- There is a preference against combining hard and soft ligands in the first coordination sphere. Examples:
 - $\text{CH}_3\text{F} + \text{CF}_3\text{I} \rightleftharpoons \text{CH}_3\text{I} + \text{CF}_4$ since CH_3^+ and I^- are soft while CF_3^+ and F^- are hard.
 - $\text{Co}(\text{NH}_3)_5\text{F}$ is stable since it has all hard ligands, but $\text{Co}(\text{NH}_3)_5\text{I}$ will react with H_2O to produce $\text{Co}(\text{NH}_3)_5\text{OH}$ since the former combines hard and soft and the latter does not.
 - Thiocyanate is **ambidentate**. Thus, we have $\text{Fe}-\text{N}$ bonds in $\text{Fe}(\text{NCS})_6^{3+}$ (with the hard Fe (III) ion) and $\text{Au}-\text{S}$ bonds in $\text{Au}(\text{SCN})_2^-$ (with the soft Au (I) ion).
 - Note that intermediate metal ions tend to bond to thiocyanate through nitrogen; indeed, intermediate Cu (II) forms $\text{Cu}(\text{NCS})_4^{2-}$.
- **Ambidentate (ligand):** A polyatomic ligand that can bind through more than one of its constituent atoms.

IV.4 Module 27: Steric Effects in Inorganic Chemistry

2/12:

- Chemistry is a game of two players: Electronic and steric effects.
- Gallanes, gallenes, cyclogallenes, and gallynes:
 - Various compounds with gallium.
 - React gallium (III) chloride with a Grignard reagent containing mesityl (2,4,6-trimethylphenyl; Mes; $\text{Me}_3\text{C}_6\text{H}_2$): $3(\text{Me}_3\text{C}_6\text{H}_2)\text{MgBr} + \text{GaCl}_3 \xrightarrow{-3\text{MgBrCl}} \text{Ga}(\text{Me}_3\text{C}_6\text{H}_2)_3$.
 - A molecule with even more steric hindrance: a three gallium ring with a 2,6-dimesitylphenyl groups attached to each gallium.
 - The triphenylcyclopropenium cation is aromatic but the cyclogallene dianion is metalloaromatic.
 - In another compound, a $\text{Ga}\equiv\text{Ga}$ triple bond holds two massive groups together.
 - The debate over this is a good example of chemistry not being black and white.
 - Acetylene and gallyne both show band structure.
 - Additionally, the former is linear while the latter is trans-bent with a donor-acceptor bond lacking a sigma bond.
 - To learn more about such bonding, read Power (2010).
- Steric effects are much more common in Lewis acid-base reactions in which larger acids are used.
 - With a small acid, increasing size in the base (i.e., replacing hydrogens with methyl groups) correlates with increasing strength.
 - With a large acid, decreasing size in the base correlates with increasing strength.
- Frustrated Lewis pairs:
 - Excessive steric bulk can prevent acids and bases from reacting to form an adduct.

- However, you can also have alternate reactions that yield compounds with Lewis acidic *and* Lewis basic regions.
- Frustrated Lewis pairs can activate (cleave) H₂.
 - They can also activate CO₂; this has applications in the realm of CO₂ sequestration.

IV.5 Chapter 6: Acid-Base and Donor-Acceptor Chemistry

From Miessler et al. (2014).

- IUPAC calls hydronium “oxonium” and uses “hydrogen ion” instead of “proton.”
- In the Brønsted-Lowry paradigm, “the equilibrium always favors the formation of *weaker* acids and bases” (Miessler et al., 2014, p. 171).
- Common amphoteric solvents: sulfuric acid, hydrofluoric acid, water, acetic acid, methanol, ammonia, and acetonitrile.
- H₃O⁺ and OH⁻ are the strongest acid and base, respectively, that^[2] can exist in H₂O.
- **Leveling effect:** Acids stronger than H₃O⁺ cannot be differentiated by their aqueous ionization.
- Because of leveling, the strength of the strong acids cannot be differentiated in aqueous solution.
 - Thus, we use more strongly acidic solvents, such as **glacial** acetic acid.
 - The use of this solvent allows us to determine that HClO₄ > HCl > H₂SO₄ > HNO₃.
 - Note that basic solvents similarly permit the differentiation of strong bases.
- **Glacial** (substance): A 100% pure, concentrated sample of a substance.
- “Nonamphoteric solvents... do not limit solute acidity or basicity because the solute does not react with the solvent. In these solvents, the inherent solute acid or base strength determines the reactivity, without a leveling effect” (Miessler et al., 2014, p. 173).
- **Superelectrophilic activation:** The result of generating small organic ions bearing a large amount of positive charge.
- Introduces the Hammett acidity function^[3].
- Protonating CH₄ is of particular interest since the natural abundance of methane (from natural gas) makes it attractive as a starting point for synthesis of more complex molecules.
 - CH₅⁺ has been isolated; CH₆²⁺ and CH₇³⁺ have been proposed.
- Dissolving SO₃ in H₂SO₄ results in **fuming sulfuric acid**, which contains higher polysulfuric acids such as H₂S₂O₇, all of which are stronger than H₂SO₄.
- Water is a strong base in superacid media: Adding H₂O to superacids produces hydronium salts.
 - As a consequence, we can form [H₃O][Ln][AsF₆]₃ where Ln is a lanthanide element as AsF₆ is part of a superacid.
- Thermodynamic measurements:
 - Hess’s law is often used to express the enthalpy of weak acid reactions in terms of reactions that do go to completion; this is not perfect, but it’s an ok starting point.

²Errata: Miessler et al. (2014, p. 173) has a typo: “than” instead of “that.”

³Errata: The – sign should be a + sign.

- One can also measure K_a via titration curves at different temperatures and use the van't Hoff equation

$$\ln K_a = \frac{-\Delta H}{R} \frac{1}{T} + \frac{\Delta S}{R}$$

- According to the above equation the plot of $\ln K_a$ vs. $\frac{1}{T}$ will be linear with a slope that allows us to determine ΔH and a y -intercept that allows us to determine ΔS .
- Brønsted basicity scale: Enthalpy changes associated with protonation by the superacid fluorosulfonic acid (HSO_3F).

- The best measure of acid/base strength is gas-phase acidity/basicity, since there are no solvent effects.
 - Wrt. $\text{HA(g)} \longrightarrow \text{A}^-(g) + \text{H}^+(g)$, $\Delta G = \text{Gas-Phase Acidity (GA)}$ and $\Delta H = \text{Proton Affinity (PA)}$.
 - There exist analogous definitions for bases.
 - Modern measurement techniques can measure these values very accurately for a select few molecules; from these molecules and Hess's law, we can build pretty good approximations of other reactions.
- **Superbase:** A base with a gas-phase proton affinity greater than 1000 kJ mol^{-1} .
 - Examples such as Grignard and organolithium reagents are ubiquitous in organic synthesis.
- Considers inductive effects.
 - “Solvation [of amines] is dependent on the number of hydrogen atoms available to form $\text{O} \cdots \text{H}-\text{N}$ hydrogen bonds with water” (Miessler et al., 2014, p. 181).
 - Steric effects are less obvious — for example, 2,6-dimethylpyridine is more basic than 4-methylpyridine, but the latter is more basic than 2-*t*-butylpyridine.
- Binary hydrogen compounds:
 - Acidity increases left-to-right across a period and down a group (wrt. the non-hydrogen atom).
 - The three heaviest hydrohalic acids are all equally acidic in water due to leveling.
- Considers relative strengths of various chlorine-based oxyacids (e.g., HClO_4 vs. HClO_3).
 - As the number of oxygens increases, the electronegativity of the terminal oxygen increases. Less and less electron density binds the hydrogen, making the bond increasingly weak and susceptible to heterolytic cleavage.
- In polyprotic acids, the pK_a increases about 5 units with each successive proton removal.
- Aqueous cations get surrounded by H_2O .
 - Then these complexes react in acid-base reactions.
 - When the concentration is sufficiently high, they can two complexes can combine via hydroxide or oxide bridges between the metal atoms.
 - Metal ions with charges of 4+ or higher are the really acidic ones. They're acidic to the extent that they do not exist by themselves in solution, but are present in oxygenated forms such as permanganate (MnO_4^-).
- **Coordinate covalent** (bond): A bond that links a Lewis acid and base into their adduct. *Also known as dative* (bond).
- **Coordination compound:** A Lewis acid-base adduct involving one or more metal ions.
- “In most Lewis acid-base reactions, the HOMO-LUMO combination forms new HOMO and LUMO orbitals of the product” (Miessler et al., 2014, p. 186).

- When the symmetries match and the energies are close, a stable adduct forms.
 - Since every molecule has a HOMO and LUMO, technically any molecule can act as acid, base, oxidizing agent, or reducing agent when combined with the right reactants.
 - For example, H_2O can oxidize Ca metal, forming $\text{Ca}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g})$. H_2O^- is not formed because the addition of electrons to antibonding orbitals in H_2O weakens one O–H bond.
- 2/14:
- Section on spectroscopic support for frontier orbital interactions.
 - Very confusing; electron transitions' colors in I_2 when it acts as a Lewis acid.
 - **Lewis basicity:** The thermodynamic tendency of a substance to act as a Lewis base.
 - It would be ideal to measure in the gas phase, but this is hard to do.
 - Determining a good reference acid is difficult.
 - When in solution, we want to make sure that we choose a solvent that will not react significantly as a Lewis acid with the solutes.
 - In other words, we need a solvent that primarily solvates the solutes through dispersion forces (i.e., nonpolar solvents).
 - Once we have a solvent, we quantify the Lewis basicity of a substance by finding the K_b and $\log K_b$ values for its complexation with a reference acid (such as I_2). The absolute basicity will vary based on the solvent, but the Lewis basicity ranking will be the same overall.
 - The standard scale for Lewis basicity is its BF_3 affinity, or $-\Delta H^\circ$ in the following reaction, corrected for the enthalpy of BF_3 dissolving in the solvent.



- **Halogen (bond):** A coordinate covalent bond formed by a halogen X_2 (e.g., I_2) or interhalogen XY (e.g., ICl) to a Lewis base.
 - Exhibit approximately 180° angles about the halogen donor atom, supporting use of σ^* LUMO.
 - I_2 affinity values are commonly determined in heptane, and while there is interest in creating an I_2 affinity scale analogous to the BF_3 one, as of yet, various experimental designs have not been reconciled.
 - Since the formation of a halogen bond involves donation into the halogen σ^* LUMO, the X–Y bond weakens and lengthens. This also decreases the stretching frequency. Measuring changes in stretching frequency induced in I_2 , ICN , and ICl via Raman spectroscopy has correlated reasonably well to basicity, but not perfectly since these stretching bands are affected by the presence of the Lewis base, too.
 - Inductive effects:
 - Adding more electronegative substituents decreases Lewis basicity (in C_{3v} Lewis bases).
 - Adding more alkyl substituents increases Lewis basicity (in C_{3v} Lewis bases).
 - Adding more electronegative substituents increases Lewis basicity (in D_{3h} Lewis bases) since the increase in bond lengths draws π -bonding electrons farther from the central boron.
 - Steric effects on Lewis acidity and basicity: **front strain**, **back strain**, and **internal strain**.
 - **Front strain:** Bulky groups interfering directly with the approach of an acid and a base to each other. *Also known as F strain.*
 - **Back strain:** Bulky groups interfering with each other when VSEPR effects force them to bend away from the other molecule forming the adduct. *Also known as B strain.*

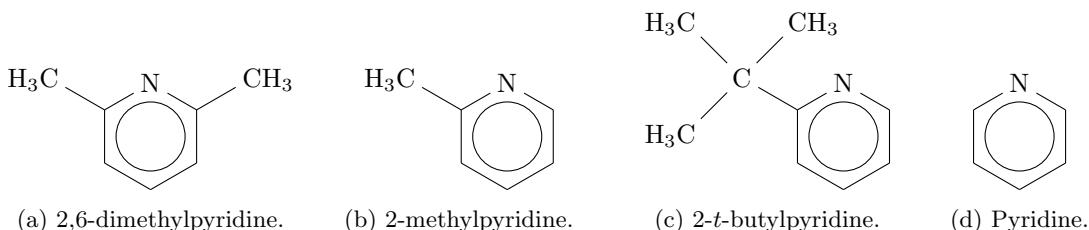


Figure IV.2: Four Lewis bases.

- **Internal strain:** Effects from electronic differences within similar molecules. *Also known as I strain.*
- An example of differences caused by F strain:
 - When reacting with H^+ , 2,6-dimethylpyridine > 2-methylpyridine > 2-*t*-butylpyridine > pyridine.
 - When reacting with BF_3 or BMe_3 , pyridine > 2-methylpyridine > 2,6-dimethylpyridine > 2-*t*-butylpyridine.
- An example of B strain: The fact that tri(*t*-butyl)boron vs. H^+ elicit a reversed order of the basicity of NH_3 , MeNH_2 , Me_2NH , and Me_3N .
- **Frustrated Lewis pair:** A lone pair on a Lewis base that does not form a traditional adduct with a Lewis acid because both molecules are hindered by excessive steric bulk. *Also known as FLP.*
- Example (refer to Figure IV.3 throughout the following discussion):

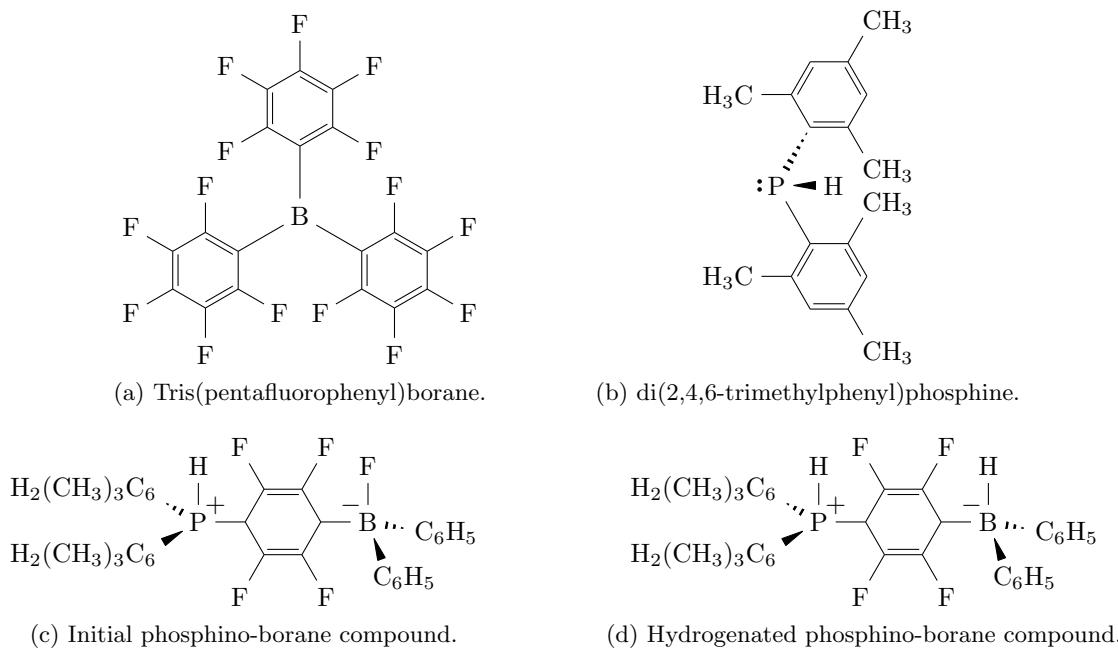


Figure IV.3: FLP phosphino-borane compound.

- Although the FLP in di(2,4,6-trimethylphenyl)phosphine does not attack the boron in tris(pentafluorophenyl)borane, it reacts with a para carbon of the borane to create a **zwitterionic** species after fluoride migration.
- This fluoride can be substituted for a hydrogen, creating a species that releases hydrogen gas upon heating and reacts with hydrogen gas at ambient temperature to reform the zwitterion.

- The phosphino-borane is the first non-transition metal species that can reversibly activate the H–H bond in H₂.
- **Zwitterion:** A species that contains at least one formal positive and negative charge.
- More small molecules have been activated with FLPs, including CO₂ and N₂O.
- **Hydrogen (bond):** An X–H…B bond formed from an attraction between an X–H unit (where the electronegativity of X is greater than that of H) and a donor atom B.
 - The components may be incorporated into larger molecular fragments.
 - Hydrogen bonds can be either intermolecular or intramolecular.
- Hydrogen bonds can be described on the basis of the varying relative contributions from three components:
 - Electrostatic contribution: The polarity of X–H.
 - Partial covalent character and charge transfer: The donor-acceptor nature of the interaction.
 - Dispersion forces.
- Experimental stipulations that must be satisfied for an interaction to be declared a hydrogen bond.
 - Bond angle close to 180°.
 - Red-shifted IR stretching frequency for the X–H bond (indicative of the bond weakening and, thus, donation into antibonding orbitals).
 - High deshielding of NMR chemical shift (extremely sensitive probe for hydrogen bonding).
 - Magnitude of ΔG for bond formation must exceed the thermal energy of the system.
- **H-bond puzzle:** Predicting the strength of hydrogen bonds solely on the basis of the structures of the participating molecules.
 - Based in the fact that structurally very similar hydrogen bonds can exhibit massive differences in strength.
 - For example, the O–H…O bond between hydronium and water is roughly six times stronger than the comparable bond between two water molecules.
- **pK_a equalization (paradigm):** An approach to predicting the strength of hydrogen bonds in aqueous solution by envisioning H-bond strength as associated with the 3-way proton transfer equilibrium between X and B.
 - The better matched the pK_a values, the stronger the predicted H-bond.
 - In effect, a strong H-bond minimizes $\Delta pK_a(X-H \cdots B) = pK_a(HX) - pK_a(BH^+)$.
- **Receptor-guest interaction:** An interaction between molecules with extended pi systems where their pi systems interact with each other to hold the molecules or portions of the molecules together. *Also known as receptor-substrate interaction, host-guest interaction.*
 - One example is a double-concave hydrocarbon buckycatcher that forms a ball-and-socket structure with C₆₀. This is an example of an **inclusion complex**.
 - The hard vs. soft distinction is largely a result of polarizability, with soft acids having high polarizability and hard acids having low polarizability.
 - Soft transition metals have *d* electrons available for π -bonding.
 - **Exchange reaction:** A reaction involving the exchange of a water and a base B.

- Is OH^- or S^{2-} more likely to form insoluble salts with 3+ transition-metal ions? Which is more likely to form insoluble salts with 2+ transition-metal ions?
 - OH^- and 3+ are hard, and thus will form insoluble salts.
 - S^{2-} and 2+ are soft, and thus will form insoluble salts.
- Both hardness/softness and strength must be considered to determine reactivity.
- MO diagrams for hard-hard and soft-soft interactions reveal that there is a lower covalent contribution in hard-hard interactions, but this is compensated for by a strong ionic contribution, so don't think that hard-hard interactions are weaker.
- Wrt. chemical hardness η (Greek “eta”), ionization energy is assumed to measure the energy of the HOMO ($E_{\text{HOMO}} = -I$) and electron affinity the energy of the LUMO ($E_{\text{LUMO}} = -A$); thus, a hard species is one with a high difference in energy between its HOMO and LUMO.
- **Chemical softness:** The inverse $\sigma = \frac{1}{\eta}$ of hardness.
- Reviews a quantitative system of acid-base parameters to account for reactivity by explicitly including electrostatic and covalent factors.

Topic V

Coordination Chemistry: Structures and Isomers of Metal Complexes

V.1 Module 28: Introduction to Coordination Compounds

2/12:

- Modern inorganic chemistry is heavily concerned with the transition metals, i.e., the *d*-block elements.
 - Most industrial catalysts utilize transition metal compounds.
- Transition metals vs. main-group elements:

Transition-Metal Compounds	Main-Group Elements
Multiple oxidation states (e.g., the 11 oxidation states of Mn from -3 to $+7$)	Single oxidation state
Brightly colored (thus a gap between HOMO and LUMO of a few electron volts)	Usually colorless
Usually have partially occupied valence <i>d</i> -orbitals that are often relatively close in energy	The valence <i>s</i> - or <i>p</i> -orbitals are either fully occupied or empty and are far apart energetically
Often paramagnetic	Usually diamagnetic
Often interact with small molecules such as CO, C ₂ H ₄ , and H ₂	Generally do not interact strongly with CO, C ₂ H ₄ , or H ₂

- The orbital energy gap matches the bond energies of small molecules pretty well.
 - This is exactly what is needed for activating those chemical bonds, i.e., making catalytic cycles!
- History:
 - Prussian blue ink is the first synthetic blue dye, and one of the first coordination compounds created (used in famous paintings such as *Starry Night*).
 - The structure of coordination complexes was not understood until 1907, however.
 - Blomstrand and Jorgenson tried to determine the structure of Co(NH₃)₄Cl₃, but their guess didn't explain isomers.
 - Alfred Werner (late 1800s) was the father of coordination chemistry.
 - He noticed that excess AgNO₃ could only liberate and precipitate as AgCl one chlorine from both the green and violet isomers of CoCl₃ · (NH₃)₄.
 - However, it could precipitate two chlorines from CoCl₃ · (NH₃)₅ and three from CoCl₃ · (NH₃)₆.
 - This observation plus a number of controls led him to **Werner's Conclusions**.

- **Werner's Conclusions:**

1. In this series of compounds, cobalt has a constant **coordination number** of 6.
2. As the NH_3 molecules are removed, they are replaced by Cl^- , which acts as if it is covalently bonded to cobalt.
3. Chloride and ammonia are now called ligands.
4. Ligands are a Lewis base/electron pair donors that can bind to a metal ion.
5. A metal complex is a metal ion combined with ligands.
6. Coordination complexes are neutral and counter ions are not bonded to the central metal ion but balance the charge.

– For example, in $[\text{Co}(\text{NH}_3)_6]^{+3}\text{Cl}_3^{-3}$, the three chloride ions are the counter ions.

- **Coordination number:** The number of groups that can bond directly to the metal.

- Werner also hypothesized an octahedral geometry for all cobalt complexes.
 - If it were hexagonal planar or trigonal antiprismatic, there would be three isomers of the coordination sphere for the compound with two chlorines (think ortho, meta, para isomers for the hexagonal planar example).
 - However, if it is octahedral, the compound with two chlorines will have two isomers (the chlorines can either be 180° to each other or 90° to each other).
 - Octahedral also reduces steric crowding.

- **Coordination compound:** A compound with a metal center, a coordination sphere, and counter ions. *Also known as coordination complex.*

- Werner also resolved hexol into optically active isomers. This was the first optically active chiral *inorganic* compound.

V.2 Module 29: Types and Classes of Ligands

- **Monodentate ligand:** A ligand that binds to a metal ion through a single donor site. *Etymology* one-toothed.
 - For example, NH_3 is a monodentate ligand.
- **Bridging ligand:** A ligand that binds to two or more metal ions simultaneously.
 - For example, O^{2-} is a bridging ligand.
- **Ambidentate ligand:** A ligand with two kinds of binding sites that can bind through one or the other but not both simultaneously.
 - For example, thiocyanide can bond through S or N but not both simultaneously.
- **Multidentate chelating ligand:** A ligand bound to a metal through several donor sites. *Etymology* multitooth crab claw (crabs grab their food with two claws in the same way a metal can be attracted to two lone pairs from different groups on the same ligand). *Also known as polydentate chelating ligand.*

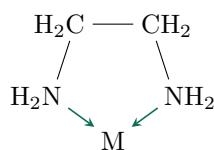


Figure V.1: An example of a bidentate chelating ligand.

- For example, ethylenediamine ($\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$; see Figure V.1) can bond to the same metal with both of its nitrogens' lone pairs at the same time.
 - Thus, ethylenediamine is bidentate, and it forms a 5-membered chelate ring.
- The chelate effect: For a given metal ion, the thermodynamic stability of a chelated complex involving bidentate or polydentate ligands is greater than that of a complex containing a corresponding number of comparable monodentate ligands.
 - Note that 5-membered rings are more stable than 6, and 4-membered rings (or smaller) are not stable due to angle strain.
 - For example, ethylenediaminetetraacetate (EDTA) is a hexadentate ligand has two nitrogens and four oxygens that wrap entirely around a metal atom and bond very strongly.
 - β -diketones and acetylacetone are also polydentate ligands.
 - Multidentate bonding is incredibly strong.
 - As one last example, hemoglobin and chlorophyll have extra stability because the iron/magnesium ion is attached to four nitrogens.
- Explanations of the chelate effect:
 - Effective concentration: If one bond breaks, the bridge between the two bonding sites in the ligands still holds the other site in close proximity to the metal, making it more likely that the bond will reform than if the metal and ligand were floating entirely independently.
 - Entropy considerations:
 - Imagine you have a coordination complex. If you substitute two monodentate ligands for two other monodentate ligands, you do not change the number of particles.
 - However, if you substitute one polydentate chelating ligand for two monodentate ligands, you increase the number of particles by 1, increasing disorder in the universe and favoring the forward reaction.
 - Looking at the temperature-dependence of equilibrium, we can calculate ΔG° . We can then use this to calculate ΔH and ΔS , and we find that the contributions are very similar. Thus, both explanations of the chelate effect contribute about equally.
- Covalent bond classification (CBC) method:
- **X-type** (ligand): A ligand that donates one electron to the metal and accepts one electron from the metal when using the neutral ligand method of electron counting, or donates two electrons to the metal when using the donor pair method of electron counting.
 - Examples: hydrogen, the halogens, hydroxide, cyanide, carbocation, and nitric oxide.
- **L-type** (ligand): A neutral ligand that donates two electrons to the metal center regardless of the electron counting method being used.
 - Examples: carbon monoxide, PR_3 , ammonia, water, carbenes ($=\text{CRR}'$), and alkenes.
- **Z-type** (ligand): A ligand that accepts two electrons from the metal center as opposed to the donation occurring with the other two types of ligands.
 - Examples: Lewis acids, such as BR_3 .
- LXZ notation:
 - Take the traditional formula and replace the central metal atom with M, each X-type ligand with X, each L-type ligand with L, and each Z-type ligand with Z. Then, if necessary, combine ligands of the same type using subscript arabic numerals, as per usual. Conserve the braces and charge.
 - For example, $[\text{Ir}(\text{CO})(\text{PPh}_3)_2(\text{Cl})(\text{NO})]^{2+}$ becomes $[\text{ML}_3\text{X}_2]^{2+}$.

V.3 Module 30: Nomenclature and Isomers of Metal Complexes

2/15:

- Midterm 1:
 - Average: 67.0.
 - Standard deviation: 14.8.
 - Max: 93.
 - 80+ scores: 10.
- Midterm 2 will cover modules 13-31 (with concepts from modules 1-12).
- Suggested reading: Miessler et al. (2014) Section 9.2.
- Naming metal complexes:
 1. For complex ions, write the cation first and anion last.
 - For example, $K_2[PtCl_4]$ is potassium tetrachloroplatinate.
 - Note that you do not use the prefix mono, di, tri, etc. here to indicate the number of cations.
 2. Name the ligands first in alphabetical order, the metals last.
 3. Prefixes to indicate numbers (di, tri, tetra, ...) for all monoatomic ligands, polyatomic ligands with short names and neutral ligands with special names.
 4. Prefixes bis-, tris-, tetrakis-, pentakis-, hexakis- for ligands whose names contain a prefix of the first type, neutral ligands without special names, and ionic ligands with particularly long names.
 5. If the anion is complex, add the suffix -ate to the name of the metal. If the symbol comes from Latin/Greek, then we go back to the Latin/Greek for the name of the anion.
 6. Put the oxidation state in Roman numerals in parentheses after the name of the central metal ion of the ligand.
- Examples:
 - $[CoCl_4]^{2-}$ is the tetrachlorocobaltate (II) ion.
 - $[Fe(CN)_6]^{4-}$ is the hexacyanoferrate (II) ion.
 - Notice the use of “ferr” for iron (this is going back to the Latin/Greek name of the central metal ion as in Step 5).
 - $[Cr(H_2O)_4Cl_2]^+$ is the tetraaquadicchlorochromium (III) ion.
 - No use of -ate because this is a cation (not an anion).
 - $[Cr(NH_2CH_2CH_2NH_2)_3]^{3+}$ is the tris(ethylenediammine)chromium (III) ion.
 - Note that ammine with two m's is correct — an **amine** is a functional group from ammonia, while an **ammine** is a group of coordination compounds where ammonia is a ligand.
- Neutral ligands have the same name as the molecule with 4 exceptions:
 - NH_3 is ammine.
 - H_2O is aquo.
 - Aquo or aqua?
 - CO is carbonyl.
 - NO is nitrosyl.
- Anionic ligands require replacing -ide in the ionic name (if present) with -o. For example:
 - Cl^- becomes chloro.
 - OH^- becomes hydroxo.

- SO_4^{2-} is still sulfate.

- Bridging ligands:

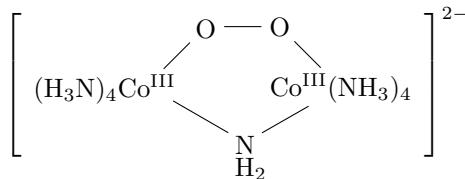


Figure V.2: A coordination compound with bridging ligands.

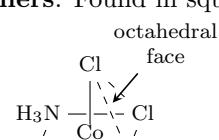
- Use μ to indicate a bridge.
- If there are more than one of a given bridging ligand, the prefix indicating the number of ligands is placed after the μ .
- If there are more than one different bridging ligands, they are given in alphabetical order.
- For example, the compound in Figure V.2 is the tetraaminecobaltate(III)- μ -amido- μ -peroxo-tetraaminecobaltate(III) ion.

Common Name	IUPAC Name	Formula
nitrido	nitrido	N^{3-}
azido	azido	N_3^-
oxo	oxido	O^{2-}
thiocyanato	thiocyanato-S	SCN^-
isothiocyanato	thiocyanato-N	NCS^-
thiocarbonyl	thiocarbonyl	CS
nitro	nitrito-N	NO_2^-
nitrito	nitrito-O	ONO^-
phosphine	phosphane	PR_3
pyridine	pyridine (abbrev. py)	$\text{C}_5\text{H}_5\text{N}$
amido	azanido	NH_2^-
imido	azanediido	NH^{2-}

Table V.1: Irregular and unfamiliar monodentate ligands.

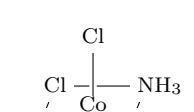
- Table V.1 lists some more common but irregular or unfamiliar ligands.
 - Note that when listed in Table V.1, ambidentate ligands bind through the atom listed most to the left in the formula.
 - Note also that the slides list a number of common bidentate and polydentate ligands.
- Suggested reading: Miessler et al. (2014) Section 9.3.
- Coordination complex isomers are divided into:
 - Structural isomers (different bonds).
 - Coordination isomerism.
 - Linkage isomerism.
 - Stereoisomers (same bonds, different spatial arrangements).

- Geometric isomerism (*cis-trans*).
 - Optical isomerism.
 - **Ionization isomers:** Ligands inside the coordination sphere exchange places with ligands outside the coordination sphere.
 - So named because these give different ions when dissolved in water.
 - Werner's four isomers of $\text{CrCl}_3 \cdot 6\text{H}_2\text{O}$ are ionization isomers.
 - **Linkage isomers:** Variations in at which site an ambidentate ligand bonds.
 - **Geometric isomers:** Found in square planar and octahedral complexes — *cis*- vs. *trans*-isomers.



octahedral
face

(a) fac-triamminetrichlorocobalt(III).



(b) mer-triamminetrichlorocobalt(III).

Figure V.3: Facial and meridional geometric isomers.



Figure V.4: Methods of sketching octahedral complexes.

- Which way you choose depends on what you are trying to show.
 - **Optical isomers:** Two compounds with non-superimposable mirror images (two chiral molecules).
 - C_1 , C_n , D_n , T , O , and I point groups are chiral.
 - Usually associated with tetrahedral/octahedral geometries.
 - Octahedral complexes have special chirality.
 - Octahedral complexes with two bidentate ligands:

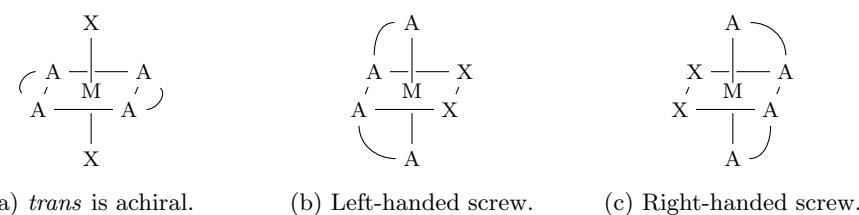


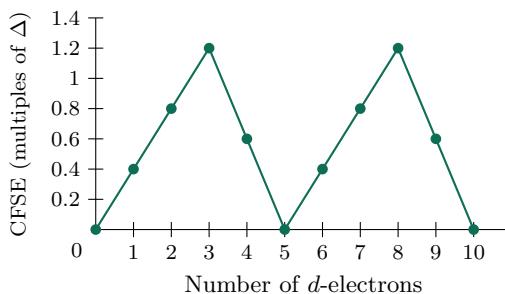
Figure V.5: Chirality of octahedral complexes with two bidentate ligands.

- As we can see in Figure V.5, *trans*-M(A–A)₂X₂ is not chiral, but *cis*-M(A–A)₂X₂ is chiral with two specially named forms.
- Octahedral complexes with three bidentate ligands:
 - *Also known as propellor chirality.*
 - Left-handed helices are Λ form, while right-handed helices are Δ form.

V.4 Module 31: Crystal Field Theory

- The first attempt to understand and rationalize the electronic structure of transition metal complexes.
 - Originally introduced to analyze crystals' electronic structure.
 - Since the coordination of a central atom in a crystalline closely mimics that of it in a coordination complex, the concepts of crystal field theory can easily be transferred to chemistry.
- Suggested reading: Miessler et al. (2014) Section 10.2.
- Refresher on *d*-orbitals.
 - There are four *d*-orbitals with four lobes and one strange d_{z^2} orbital.
 - Higher energy level ones have lobes corresponding to the change in sign of the radial component..
 - The d_{z^2} orbital is a linear combination of two four-lobed orbitals; it is created to reconcile mathematical theory with physical reality (the Pauli exclusion principle).
- Crystal field theory describes an electrostatic (ionic) approach to bonding — it is so named because it was first applied to crystalline substances.
 - Interactions between filled *d*-orbitals and ligands with excessive electrons are repulsive.
- Assumptions:
 1. Metal ion at the center.
 2. Ligands are treated as point charges.
 3. Bonding occurs through M⁺ and L[−] electrostatic attraction.
 4. Bonding is purely ionic.
 5. M and L electrons repel each other.
 6. *d*-orbital degeneracy is broken as ligands approach.
- Consider *d*-orbitals bonding to six ligands.
 - Keep Coulomb's Law ($E \propto \frac{q_1 q_2}{r}$) in mind: *d*-orbitals that overlap more with the ligands (smaller *r*) will be more destabilized (higher *E*).
 - d_{z^2} overlaps with the two axial ligands, and $d_{x^2-y^2}$ overlaps with the four equatorial ligands.
 - None of d_{xy} , d_{xz} , and d_{yz} overlap significantly with any ligands.
 - Therefore, the five degenerate *d*-orbitals split into the **t_{2g} set** and the **e_g set**.
 - When the orbitals split, they maintain an energetic “center of mass,” i.e., their combined energy as molecular orbitals must still be equal to their combined energy as degenerate, atomic orbitals. Thus, the stabilization energy of the three orbitals in the t_{2g} set is $-\frac{2}{5}\Delta$ while the destabilization energy of the two orbitals in the e_g set is $\frac{3}{5}\Delta$, where Δ is the **crystal field splitting parameter**.
- **t_{2g} set:** The three orbitals that lie between the ligand donor atoms.
- **e_g set:** The two orbitals that lie along the Cartesian coordinates, and so are adjacent to the donor atoms of the ligands, raising the set in energy.

- **Crystal field splitting parameter:** Different ligands produce different extents of splitting between the e_g and the t_{2g} levels. This energy difference is the crystal field splitting parameter. *Units* cm^{-1} . *Also known as* Δ , $10Dq$.
- Experimental verification of orbital splitting.
 - Consider a coordination complex with just one d -electron, i.e., electron configuration d^1 .
 - An example is $[\text{Ti}(\text{H}_2\text{O})_6]^{3+}$.
 - In such a complex, the electron will occupy the lowest energy orbital available, i.e., one of the three degenerate t_{2g} orbitals.
 - Shining light on the complex can promote the t_{2g} electron into the e_g energy level.
 - The UV-Vis absorption spectrum reveals that this transition occurs with a maximum at $20\,300 \text{ cm}^{-1}$, or $\Delta = 243 \text{ kJ/mol} \neq 0$.
- Note that $1000 \text{ cm}^{-1} = 11.96 \text{ kJ/mol} = 2.86 \text{ kcal/mol} = 0.124 \text{ eV}$.
- **Crystal field stabilization energy:** The overall change in energy when the d -subshell splits, which is given by $(0.4n(t_{2g}) - 0.6n(e_g))\Delta$ where $n(t_{2g})$ and $n(e_g)$ are the numbers of electrons in the t_{2g} and e_g levels respectively. *Also known as* CFSE.
 - When splitting of the d -subshell occurs, the occupation of the lower energy t_{2g} level by electrons causes a stabilization of the complex, whereas occupation of the e_g level causes a rise in energy. The t_{2g} level drops by 0.4Δ , whereas the e_g level is raised by 0.6Δ .
- High and low-spin complexes:
 - Whether a complex is **high-spin** or **low-spin** depends on Δ .
 - If $\Delta > P$ where P is the **spin-pairing energy**, then the complex is low-spin, and vice versa if $\Delta < P$.
- **High-spin** (complex): A complex with d^4 to d^8 electron configuration, where the electrons spread out and occupy the whole d -subshell.
 - High-spin complexes are often paramagnetic.
 - Electrons fill the whole d -subshell according to Hund's rule.
- **Low-spin** (complex): A complex with d^4 to d^8 electron configurations, where the t_{2g} energy level is filled first.
 - Low-spin complexes are often diamagnetic.
- **Spin-pairing energy:** The energy required to take pairs of electrons with the same spin orientation, and pair them up with the opposite spin. *Also known as* P .
- Calculating CFSE of d^0 to d^{10} high-spin M (II) ions.

Figure V.6: CFSE as a function of the number of d -electrons.

- Most first row and many second and third row transition elements will be prone to forming high-spin complexes.
 - The variation shown in Figure V.6 reveals that complexes of metal ions with high CFSE (such as Ni (II)) will undergo greater stabilization, and vice versa for metal ions with low CFSE (such as Ca (II)).
 - The predicted variation matches relatively well with formation constant values ($\log K_1$) obtained experimentally for these compounds.
- We can also look at orbital splitting in coordination compounds of other geometries.
 - In T_d compounds, for example, the splitting is flipped with the $d_{xy,xz,yz}$ orbitals destabilized and the $d_{x^2-y^2,z^2}$ orbitals stabilized.
 - We can also look analyze linear and square planar geometries.
 - Merits of crystal field theory:
 - Can be used to predict the most favorable geometry for the complex.
 - Can account for why some complexes are tetrahedral and others are square planar.
 - Useful in interpreting magnetic properties.
 - The colors of many transition metal complexes can be rationalized.
 - Limitations of crystal field theory:
 - Becomes less accurate as delocalization (covalent character) increases.
 - Point charge does not accurately represent complexes.
 - Does not account for π bonding interactions.
 - Does not account for the relative strengths of the ligands.

V.5 Module 32: Ligand Field Theory for the O_h σ -Only Case

2/17:

- Suggested reading: Miessler et al. (2014) Section 10.3.
- Ligand field theory:
 - Application of molecular orbital theory to transition metal complexes.
 - Ligands are not point charges.
 - Takes into account π bonding.
 - I.e., accounts for the fact that ligands can be σ -donors, π -donors, π -acceptors, or sometimes multiple simultaneously.
 - Can be used to explain spectrochemical series.
 - Better than valence-bond model or crystal field theory at explaining experimental data.
- Octahedral σ -only MO diagram workflow:
 1. Assign a point group.
 2. Choose basis function.
 3. Apply operations.
 4. Generate a reducible representation.
 5. Reduce to irreducible representations.
 6. Combine orbitals by their symmetry.
 7. Fill MOs with e^- .

8. Generate SALCs of peripheral atoms.
 9. Draw peripheral atom SALC with central atom orbital to generate bonding/antibonding MOs.
- MH_6 example.
 - Point group: O_h .
 - Basis functions: 1s-orbitals on the ligands (you can also choose the σ -bond vectors if you wish [if it's easier for you]), and s, p, d -orbitals on the metal center.
 - Apply operations & generate a reducible representation.

$$\Gamma_{\text{H}} = (6, 0, 0, 2, 2, 0, 0, 0, 4, 2) = A_{1g} + T_{1u} + E_g$$

$$\Gamma_{\text{M}_s} = A_{1g}$$

$$\Gamma_{\text{M}_p} = T_{1u}$$

$$\Gamma_{\text{M}_{d_{x^2-y^2,z^2}}} = E_g$$

$$\Gamma_{\text{M}_{d_{xy,xz,yz}}} = T_{2g}$$

- Combine orbitals by their symmetry.

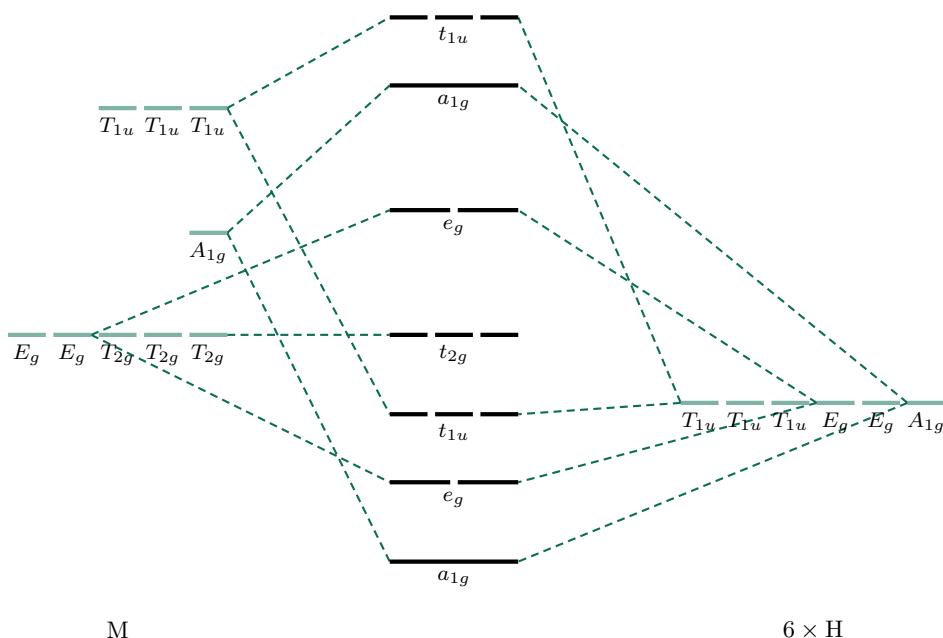


Figure V.7: MH_6 orbital diagram.

- In a transition metal compound/coordination complex, the t_{2g} MOs would be the HOMOs and the antibonding e_g MOs would be the LUMOs.
- Since the t_{2g} orbitals are nonbonding, they are reflective of the energy of the metal d -orbitals. Thus, the energy difference between them and the e_g orbitals is equal to the amount by which the e_g orbitals' energy changes during bonding, i.e., the splitting parameter. This energy change is also indicative of the antibonding character of the e_g orbitals, and consequently the strength of the bonding (bigger Δ implies higher energy e_g orbitals implies stronger bonding).
- In LFT, we call this quantity the **ligand field splitting parameter**, as opposed to the crystal field splitting parameter.
- Pauses to consider **weak field ligands** and **strong field ligands**.

- Generate SALCs of peripheral atoms.

$$\begin{aligned}\Psi(a_{1g}) &= c_1(4s) + c_2(\sigma_1 + \sigma_2 + \sigma_3 + \sigma_4 + \sigma_5 + \sigma_6) \\ \Psi_z(t_{1u}) &= c_3(4p_z) + c_4(\sigma_1 - \sigma_6) \\ \Psi_x(t_{1u}) &= c_3(4p_x) + c_4(\sigma_4 - \sigma_2) \\ \Psi_y(t_{1u}) &= c_3(4p_y) + c_4(\sigma_3 - \sigma_5) \\ \Psi_{z^2}(e_g) &= c_5(3d_{z^2}) + c_6(\sigma_1 + \sigma_6) + c_7(-\sigma_2 - \sigma_3 - \sigma_4 - \sigma_5) \\ \Psi_{x^2-y^2}(e_g) &= c_8(3d_{x^2-y^2}) + c_9(\sigma_2 - \sigma_3 + \sigma_4 - \sigma_5)\end{aligned}$$

- Recall that the coefficients reflect the degree of overlap.
- Note that for the $\Psi(t_{1u})$ wavefunctions, the in-plane σ -orbitals have zero overlap with the perpendicular $4p$ orbital, hence their coefficient of 0.
- Draw peripheral atom SALC with central atom orbital to generate bonding/antibonding MOs.

Figure V.8: MH_6 bonding and antibonding MOs.

- **Weak field ligand:** A ligand with high electronegativity, and thus low energy frontier orbitals.
 - Weak σ donor.
 - Weak Lewis base.
 - Weaker bonding interaction.
 - Smaller Δ_o .
- **Strong field ligand:** A ligand with comparable energy frontier orbitals to the d -orbitals of the metal cation.

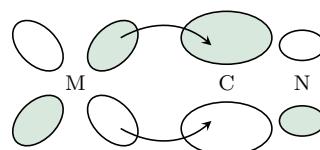
- Strong σ donor.
 - Strong Lewis base.
 - Stronger bonding interaction.
 - Larger Δ_o .
- Size matters more than Lewis basicity:
 - Stronger Lewis base \Rightarrow larger Δ_o .
 - Smaller ligand \Rightarrow larger Δ_o .
 - However, size overall is more important.
 - This is related to hardness and softness — with a hard Lewis acid metal center, harder ligands will bond with greater strengths.
 - For the halides' Δ_o : $I^- < Br^- < Cl^- < F^-$.
 - If we use p -orbitals pointing toward the center as bases for our peripheral atoms instead of s -orbitals, we get the exact same result.
 - Factors that influence $\Delta_{\text{oct}} = \Delta_o = 10Dq$.
 - Metal oxidation state: Increasing the charge increases Δ .
 - Principal quantum number: Increasing n (going down a period on the periodic table) increases Δ , too.
 - Nature of the ligand:

L	Ligand Type	z	$\Delta_o (\text{cm}^{-1})$
Cl^-	σ, π -donor	3-	13 700
H_2O	σ -donor	3+	17 400
NH_3	σ -donor	3+	21 500
CN^-	σ -donor/ π -acceptor	3-	26 600

Table V.2: How the nature of the ligand influences Δ .

V.6 Module 33: Ligand Field Theory for the O_h $\sigma + \pi$ Case

- Consider cyano-ligands (isoelectronic to CO and NO^+).
 - HOMO is perfect for σ bonding via donation.
 - LUMO is perfect for π acceptance from metal d -orbitals.

Figure V.9: LUMO π -acceptance.

- d_{xz} and π^* orbitals combine with stabilization energy e_π .
- This is how we determine what kinds of donors the ligands in Table V.2 are. For H_2O , for example, it is not a π -acceptor because its LUMO is a σ^* orbital, not a π^* one (see Figure III.11).

- Octahedral $\sigma + \pi$ bonding:

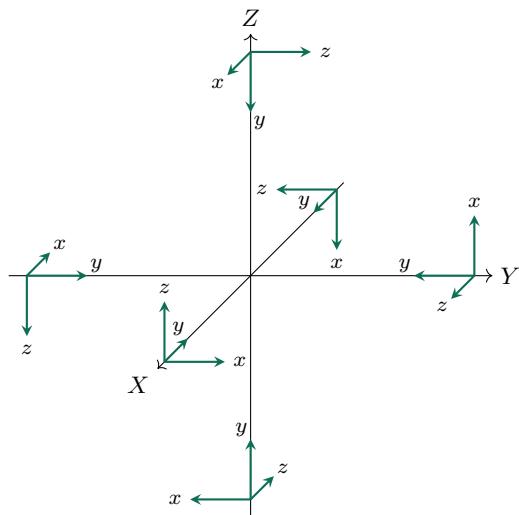


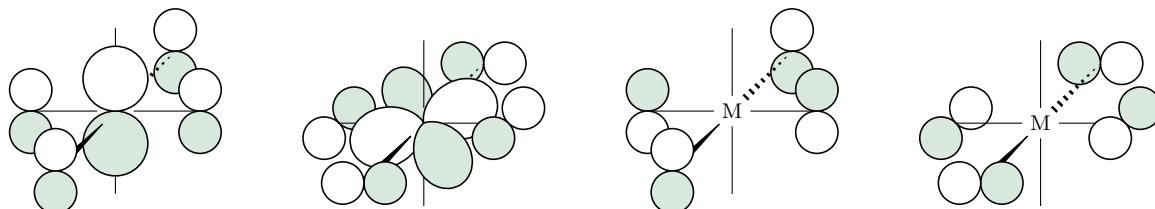
Figure V.10: Octahedral π -ligand coordinate systems.

- Generate a reducible representation: $\Gamma_{L\pi_{x,z}} = (12, 0, 0, -4, 0, 0, 0, 0, 0) = T_{1g} + T_{2g} + T_{1u} + T_{2u}$.
- Combine orbitals by their symmetry.



Figure V.11: ML_6 orbital diagram.

2/22:

(a) t_{1u} type (3 of these). (b) t_{2g} type (3 of these). (c) t_{2u} type (3 of these). (d) t_{1g} type (3 of these).Figure V.12: ML_6 bonding and antibonding MOs.

- There are three of each type of triply degenerate orbital, formed as in Figure V.12.
- If the ligand is an empty π -acceptor, π -acid that accepts electrons from M and thus has high energy T_{2g} orbitals, there is a stabilization energy.
 - If the ligand is a filled π -donor, π -base that donates electrons to M and thus has low energy T_{2g} orbitals, there is increased antibonding character.
- The following is a more formal treatment of the effects of π -bonding on Δ in an octahedral complex.
 - If the ligand possesses π -orbitals as well as σ -symmetry (as in Figure V.9), we need to take into account the overlap of these orbitals with the t_{2g} orbitals of the metal.
 - Two situations may arise: Either the ligand π -orbitals are empty and of higher energy than the metal t_{2g} orbitals, or the ligand π -orbitals are filled and of lower energy than the metal t_{2g} orbitals.
 - The former case may increase Δ and is the most important case, including ligands such as CO, NO^+ , and CN^- . This type of covalency, called π -bonding or back bonding, provides a plausible explanation of the position of these ligands in the spectrochemical series.
- Ligand field strength:
 - Characteristics of σ -bonding that make for a stronger field:
 - Stronger σ -donor.
 - Strong Lewis base.
 - Stronger bonding interaction.
 - Characteristics of π -bonding that make for a stronger field:
 - Empty π -acceptor.
 - π -acid.
 - Accepts from M .
 - And vice versa for what makes a weaker field.
- **Spectrochemical series:** An arrangement from greatest to least of ligands by ligand field strength.

$$\text{CO, CN}^- > \text{phen} > \text{NO}_2^- > \text{en} > \text{NH}_3 > \text{NCS}^- > \text{H}_2\text{O} > \text{F}^- > \text{RCO}_2^- > \text{OH}^- > \text{Cl}^- > \text{Br}^- > \text{I}^-$$
 - Note that en is ethylenediamine and phen is phenanthroline.
 - As in Table V.2, we see that π -acceptors are stronger than pure σ -donors are stronger than π -donors.
 - With different metal ions, the order of increasing Δ with different ligands is often the same. All metal ions produce high values of Δ in their hexacyano complex, while the hexafluoro complex always produces a low value of Δ . The increase in Δ with the changing of ligands can be placed in an order known as the spectrochemical series.

- Increasing Δ correlates with a color change (because of electron transitions because t_{2g} and e_g orbitals).

V.7 TA Review Session 2

2/19:

- σ bonds are stronger than π bonds.
- Steric number is also one-half the sum of the number of electrons around the central atom plus the number of bonding electrons each ligand contributes.
 - For example, in SF_4 , $\text{SN} = \frac{6+4}{2} = 5$.
 - A_1 orbitals mix more than B_1 , mix more than B_2 .
- Determining symmetry of SALCs:
 - Multiply unmoved atoms by each row in the character table; nonzero rows indicate symmetry under that SALC.
 - In the SF_4 example, let $F_{a,d}$ be the axial hydrogens and let $F_{b,c}$ be the equatorial ones. Then the projections of F_a are F_a, F_d, F_a, F_d . Multiplying by A_1 gives us $2F_a + 2F_d\dots$
- When numbering MOs, you're good to start with 1 at the bottom of the diagram (not taking into account the lower energy states).
- We're good to use arcs for bidentate ligand instead of drawing out the full structure.
- We can denote enantiomers with the dashed line (like a mirror plane) between them.
- For neutral coordination centers, put a 0 in parentheses after the compound (in place of a Roman numeral oxidation number).
 - Rubitium tetrafluoroaurate(III).
 - Learn Greek/Latin names!
- BH_2^+ is linear since it is isoelectronic to BeH_2 , i.e., the central atom only has two bonding electrons.
 - Reduce $D_{\infty h}$ to D_{2h} .
 - Empty orbitals can still mix occupied orbitals; empty p orbitals are still present in the orbital diagram.
 - BH_2^- would have a different structure (bent).
 - C_{2v} .
 - Know how to create a Walsh diagram!
 - An extra p orbital mixes, lowering energy, in the bent form.
- Compressing a 1D crystal:
 - a decreases, overlap increases, β (interaction integral) increases, band width ($|4\beta|$) increases.
 - Effective mass: $\frac{1}{m^*} = \frac{1}{\hbar^2}(\partial^2 E / \partial k^2)_{k=0}$.
- Review reference materials; skim chapter 9; read over notes.

V.8 Office Hours (Talapin)

- What is “d.I.” and “r.I.” with respect to band theory (on the slide with the Wigner-Seitz cell)?
 - Direct lattice (the lattice in physical space) and reciprocal lattice (the lattice in the k -space, the momentum space).
 - FCC is very common lattice in metals and semiconductors.
 - Its reciprocal lattice is the Brillouin zone, the Wigner-Seitz cell (the boronomic cell/boronomic tesselation in mathematics?).

- In CHEM 39000, he covers this more (after a more direct treatment of what a direct and reciprocal lattice is).
- Aquo or aqua?
- When Wooje went through the SF_4 example on the practice exam today, he applied the projection operator to two different atoms to generate 4 SALCs. Why?
 - SALCs generated from the same atom the same way are similar. However, linear combinations of two SALCs are distinct from their components.
 - We *do* have three SALCs for NH_3 , but we only have two states, so we have to eliminate one SALC by combining it with another to arrive at two solutions (one original and one linear combination) with physical meaning (we will actually see in NH_3 both of the final SALCs we arrive at).
- Does the d_{z^2} orbital actually exist (like can we get one of those electron density plots for an atom that shows its outline), or is it just a mathematical tool?
 - Yes — see explanation of previous question.
- It won't harm you to put the numbers on the MOs.
 - If we number, we can number from the bottom and put a note that we are not counting for more core electrons.

Topic VI

Bonding and Physical Properties of Metal Complexes

VI.1 Module 34: Magnetic Properties of Transition Metal Complexes

2/22:

- Electrons occupy the lowest energy triply degenerate orbitals in d^1 , d^2 , and d^3 configurations.
- However, in the d^4 configuration:
 - Low spin: The fourth electron will pair up in the lower t_{2g} energy level.
 - High spin: The fourth electron will occupy a higher energy e_g orbital.
- The pairing energy Π is made up of two parts (refer to Figure 0.9 and the associated discussion):
 1. Coulombic repulsion energy caused by having two electrons in the same orbital. Destabilization energy contribution of Π_c for each doubly occupied orbital. Has a positive sign because it increases the energy of the system.
 2. Exchange stabilization energy for each pair of electrons having the same spin and same energy. Stabilizing contribution of Π_e for each pair having same spin and same energy. Has a negative sign because it reduces the energy of the system.
- Deciding whether the fourth electron will go into the higher energy e_g orbital at an energy cost of Δ , or be paired at an energy cost of Π .
 - Strong field ligand has big Δ so $\Pi < \Delta$; this implies a low spin configuration.
 - Weak field ligand has small Δ so $\Pi > \Delta$; this implies a high spin configuration.
- We can experimentally discriminate between high- and low-spin compounds by measuring magnetic properties.
 - The Gouy balance can determine the magnetic susceptibility of materials.
 - A more modern way to measure magnetic properties uses a Superconducting Quantum Interference Device, or SQUID.
 - This device is just about the most sensitive machine humanity can build (can detect the magnetic field of the heart/brain).
- Main types of magnetic behavior:
 - Diamagnetism (from electron charge).
 - Paramagnetism (spin and orbital motion of electrons on individual atoms).

- Ferromagnetism and antiferromagnetism (cooperative interaction between magnetic moments of individual atoms).
- Paramagnetism is much stronger than diamagnetism and overpowers it.
 - Ferromagnetism overpowers both.
- Theoretical background for determining magnetic spins experimentally:
 - When we place a sample in a magnetic field of magnitude H , the sample will interact with the magnetic field and magnetize. This magnetization causes the magnetic flux B in the material to differ from the magnetic flux through the space the sample occupies (were the sample not there) by an amount determined by the magnetization parameter M , which is specific to each material. These three quantities are related via the equation

$$B = H + 4\pi M$$

- If we divide the flux by the magnetic field, we obtain the magnetic susceptibility per unit volume κ of the material:

$$\frac{B}{H} = 1 + 4\pi \cdot \frac{M}{H} = 1 + 4\pi\kappa$$

- This quantity can be normalized by the molecular weight and density of the substance to give the magnetic susceptibility per mole

$$\chi_M = \kappa \cdot \frac{\text{molecular weight}}{\text{density}}$$

- Dividing χ_M by Avogadro's number gives the magnetic susceptibility per molecule χ_M^{corr} .
- **Curie's law** relates χ_M^{corr} to the magnetic moment μ by the formula

$$\chi_M^{\text{corr}} = \frac{N\mu^2 k}{3T}$$

where N is Avogadro's number, $k = 1.381 \times 10^{-23}$ J/K is the Boltzmann constant, and T is the absolute temperature of the substance.

- Note that μ is measured in units of Bohr magnetons where $1 \text{ BM} = \frac{e\hbar}{4\pi m_e c}$. As per usual, we have $e = 1.602 \times 10^{-19}$ C is the charge of an electron, $\hbar = 6.626 \times 10^{-34}$ Js is Planck's constant, $m_e = 9.11 \times 10^{-31}$ kg is the mass of an electron, and $c = 2.998 \times 10^8$ m/s is the speed of light.
- We can rearrange Curie's law to express the magnetic moment in terms of χ_M^{corr} as follows.

$$\mu = \sqrt{3k/N} \cdot \sqrt{\chi_M^{\text{corr}} T}$$

- Magnetic moment μ and the spin-only formula: Materials that are diamagnetic are repelled by a magnetic field, whereas paramagnetic substances are attracted into a magnetic field, i.e., show magnetic susceptibility. The unpaired electrons in paramagnetic complexes of $3d$ -block metal ions create a magnetic field. The magnetic moment μ is then given by the spin-only formula

$$\mu_{\text{spin-only}} = \sqrt{n(n+2)}$$

where n is the number of unpaired electrons.

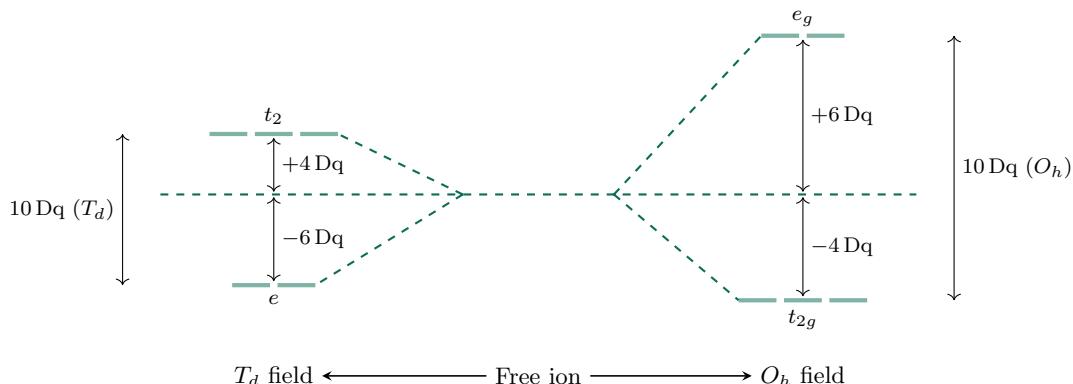
- In heavier transition metals, we need to account for not just the S quantum number but also L (which accounts for some ground state relativistic effects) by using the formula

$$\mu_{S+L} = \sqrt{4S(S+1) + L(L+1)}$$

VI.2 Module 35: Reflections on the Ligand Field Effects in O_h and T_d Complexes

2/24:

- Note that 2nd and 3rd row metals almost always low-spin, and 4th row transition metals often high-spin.
- T_d vs. O_h splitting.

Figure VI.1: T_d vs. O_h splitting.

- For T_d , 2 e -type orbitals are lower in energy and 3 t_2 orbitals are higher.
 - Note that we do not mark with gerade because T_d molecules lack an inversion center.
- For O_h , it's reversed.
- Note that the splitting energy of tetrahedral complexes is less than that of octahedral complexes. This is because there are fewer ligands acting on the d -orbitals of the metal center (4 vs. 6), and the angular overlap of the d -orbitals and the ligand group orbitals is less favorable when tetrahedral (there is a directional factor of $\frac{2}{3}$). Indeed, the tetrahedral splitting energy is generally $\frac{4}{6} \cdot \frac{2}{3} = \frac{4}{9}$ that of a relative octahedral splitting energy.
- Conclusion: T_d complexes are always weak-field and thus high spin.
- Rationalization of coordination geometries (factors that influence the geometry adopted):
 - Electronic factor (the number of bonds): Electrostatic and covalent model favor O_h (6 vs. 4).
 - Steric factor: Ligand-ligand repulsions favor T_d .
 - High charge on cation increases Δ — favors O_h coordination and O_h low spin. Indeed, O_h (l.s.) $>$ O_h (h.s.) $>$ T_d .
 - CFSE(O_h) \geq CFSE(T_d), always.
- In d^5 complexes (e.g., with Mn^{2+} and Fe^{3+}), there is no stabilization energy. Thus, L-L repulsions dominate and T_d complexes are formed.
 - Note however that the higher charge on Fe^{3+} favors low spin O_h complexes in more cases than Mn^{2+} .
- In d^6 complexes (e.g., with Fe^{2+} and Co^{3+}), the stabilization is much higher for O_h low spin than high spin than T_d . Thus, we form low spin complexes.
 - Note that there are exceptions — extremely weak field ligands such as F^- can still form high spin complexes as with CoF_6^{3-} .

VI.3 Module 36: Angular Overlap Model

- In many catalytic cycles, the coordination environment (molecular geometry) changes frequently throughout the cycle.
- To account for changes in the coordination environment, we use the **angular overlap model**.
 - We can also use this model to help account for complexes with different ligands in the coordination sphere.
- Recall that σ -bonding is stabilizing, but π -bonding is stabilizing only in the case of π -acceptance by a ligand, not π -donation.
- Angular overlap parameters:

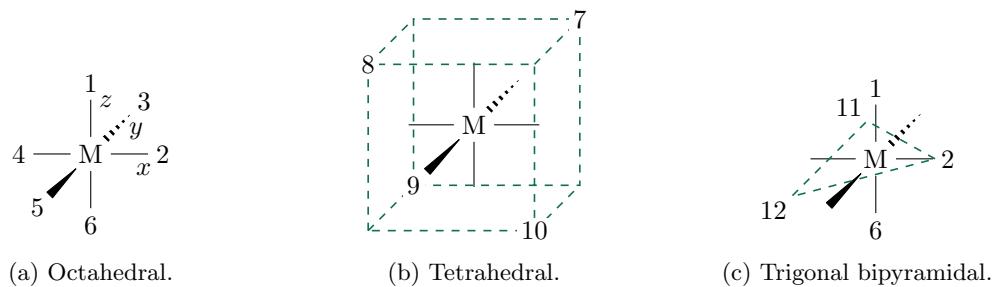


Figure VI.2: Angular overlap model: Labeling of positions.

- With multiple ligands and multiple orbitals, we add the angular overlap interaction coefficients.
- These are tabulated for each orbital of each ligand at each position in the coordination sphere with each metal-center d -orbital.
- Suggested reading (on coefficients' derivation): *TBD*.
- Trigonal planar coordination example:
 - From Figure VI.2c, the ligand positions are 2, 11, and 12.
 - Thus, we add the coefficients in these rows to get $e_\sigma = (\frac{3}{4}, \frac{9}{8}, \frac{9}{8}, 0, 0)$ and $e_\pi = (0, \frac{3}{2}, \frac{3}{2}, \frac{3}{2}, \frac{3}{2})$, where the respective d -orbitals are $z^2, x^2 - y^2, xy, xz, yz$.
 - With these energies, we can now sum $e_\sigma + e_\pi$ to determine that the energies of the orbitals are $(\frac{3}{4}, \frac{21}{8}, \frac{21}{8}, \frac{3}{2}, \frac{3}{2})$.
 - This gives us three sets of degenerate orbitals: Lowest energy ($d_{xz,yz}$), medium energy (d_{z^2}), and high energy (d_{xy,x^2-y^2}).
 - How did we get these energy rankings?
 - Assigning Mulliken symbols with the D_{3h} character table, we have from lowest to highest energy: $e'' < a'_1 < e'$.
- Note that e_σ is always positive (because ligands are σ -donors), but e_π can be negative (because ligands can be π -acceptors).
- Changing the metal and/or ligand affects the magnitudes of e_σ and e_π , thereby changing the value of Δ .
- $e_\sigma > e_\pi$ always.
- Values decrease with increasing size and decreasing electronegativity.
- Both positive and negative values for e_π .

VI.4 Module 37: Jahn-Teller Effect

- The **Jahn-Teller theorem** helps explain why the d^9 configuration is far more stable (far higher peak) than predicted by Figure V.6.
- Jahn-Teller theorem:** For nonlinear molecules/ions that have a degenerate ground-state, the molecule/ion will distort to remove the degeneracy. *Also known as J-T theorem.*
 - When orbitals in the same level are occupied by different numbers of electrons, this will lead to distortion of the molecule.
 - If the two orbitals of the e_g level have different numbers of electrons, this will lead to J-T distortion so as to stabilize the doubly occupied e_g orbital and destabilize the singly occupied e_g orbital.
 - Cu(II) with its d^9 configuration is degenerate and has J-T distortion.
- Consider the two degenerate e_g orbitals ($d_{x^2-y^2,z^2}$). *picture*
 - Elongating the z -axis in an O_h complex stabilizes the d_{z^2} orbital and destabilizes the $d_{x^2-y^2}$ orbital.
 - Vice versa for compressing the z -axis.
- Thus, we can see significant elongation of the z -axis bonds in $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$.
- History: Before the rigorous formulation and verification of the J-T theorem by Jahn and Teller, Landau proposed the **Landau statement** from his observations.
- Landau statement:** A molecule in an orbitally degenerate electronic state is unstable with respect to spontaneous distortion of the nuclear configuration that removes the degeneracy.
- Strength of the J-T effect in various configurations:

Number of electrons	1	2	3	4	5	6	7	8	9	10
High-spin Jahn-Teller	w	w		s		w	w		s	
Low-spin Jahn-Teller	w	w		w	w		s		s	

Table VI.1: Jahn-Teller effects in various configurations.

- Unequal occupation of t_{2g} orbitals leads to the J-T effect in principle, but only weakly in practice because t_{2g} orbitals are nonbonding in σ -bonded complexes, i.e., localized on the metal center, i.e., not strongly perturbed by ligand bonding.
- Unequal occupation of e_g orbitals leads to a strong J-T effect since they are antibonding.
- Structural effects of J-T distortion:
 - In $[\text{Cu}(\text{en})_2(\text{H}_2\text{O})_2]^{2+}$, we will have axial aqua groups and equatorial en groups for two reasons.
 - The aqua groups are weaker field ligands that interact less efficiently with the metal center, so it is easier for them to be farther from it.
 - Having the en groups in plane means they don't have to be structurally distorted.
 - In $[\text{Cu}(\text{en})_3]^{2+}$, we observe strong structural distortion from perfect octahedral forced by the J-T distortion; this angle strain is not energetically favorable.
- Jahn-Teller distortion of the excited state:
 - In high spin d^6 complexes such as $[\text{Fe}(\text{H}_2\text{O})_6]^{2+}$, there will only be weak J-T distortion of the ground state. Thus, we expect to see only one peak in the absorption spectrum, corresponding to the promotion of a t_{2g} electron to the e_g orbitals and its fall back down.

- However, we observe two bands.
- This is because promotion of an electron from the t_{2g} orbitals to the e_g orbitals leads to a much stronger J-T distortion (unequally occupied e_g orbitals).
- The resultant d -orbital splitting causes the two absorption peaks.
- Square planar complexes:
 - Jahn-Teller distortion leads to tetragonal distortion of the octahedron, with the extreme of tetragonal distortion being the complete loss of axial ligands, and formation of a square-planar complex. Tetragonal distortion is the stretching of the axial M–L bonds, and shortening of the in-plane bonds. Cu(II) is usually tetragonally distorted, while low-spin Ni(II) is usually square planar.
 - Since the axial bonds get weaker as they lengthen, eventually we can have enough thermal energy to break them entirely, resulting in a square planar complex.
 - This occurs in the case of Ni(II) bonded to strong field ligands, such as cyano ligands. Essentially, what happens is the splitting of the e_g orbitals exceeds the spin-pairing energy, causing the $d_{x^2-y^2}$ electron to pair with the d_{z^2} electron.
 - The filled d_{z^2} orbital now occupies two coordination sites, and the four donor atoms occupy the plane.
 - The structure is comparable to that of $[IF_4]^-$, where two lone pairs occupy the axial sites.
 - This is a particularly important special example because such compounds are very reactive, owing to their frontier d_{z^2} orbitals, and can be involved in nucleophilic attacks.
 - All high-spin d^8 metal ions are octahedral (or tetrahedral). Low-spin d^8 metal ions are usually square planar.
 - Both Wilkinson's catalyst and Crabtree's catalyst are square planar!

VI.5 Module 38: Applying MO Theory Beyond “Simple” ML_6 Complexes

- 2/26:
- Reviews Figure V.11 and the difference between π -acceptor and π -donor ligands.
 - Mixing different ligands within the same coordination sphere:
 - The MO diagram for O_h complexes proves to be a convenient starting point for deducing the electronic structure of many lower symmetry metal complexes.
 - To analyze lower symmetry compounds, we use the descent in symmetry technique.
 - $[Co(CN)_5Br]^{3-}$ example:
 - Five strong-field cyano ligands and one weak-field bromo ligand.
 - Choose to orient the Br^- ligand along the positive z -axis.
 - To build the MO, we could use DFT (first principles), but it's an unintuitive black box.
 - Alternatively, we can start with hexacyanocobaltate(III), remove one cyano ligand and see how the MOs are perturbed, and add one bromo ligand and see how the MOs are perturbed. This process correlates the electronic structure of this C_{4v} complex with its O_h parent complex.
 - To begin, consider perturbations to σ and π interactions upon substituting π -accepting CN^- with π -donating Br^- .
 - First, let's simply remove the CN^- ligand.
 - The $d_{x^2-y^2}$ orbital is not greatly perturbed by substitution along the z -axis.
 - The d_{z^2} orbital was $M-L\sigma^*$; thus, removal of 1 σ ligand from the z -axis will stabilize it.
 - The $d_{xz,yz}$ orbitals are destabilized owing to the removal of 1 $M-L\pi$ bonding interaction.

- The d_{xy} orbital is in-plane and nodal with respect to σ - and π -bonding along the z -axis; thus, it is not greatly perturbed by such substitution.
- The d_{z^2} orbital is more greatly stabilized than the $d_{xz,yz}$ orbitals are destabilized because the former is involved in stronger σ -interactions, as opposed to weaker π -interactions.
- Now let's add the Br^- ligand. Addition of Br^- to the C_{4v} fragment will give rise to new interactions.
 - σ : $\text{Br}(p_z)$ will interact with the d_{z^2} , s , and p_z orbitals of cobalt. All of these will be $M-\text{L}\sigma^*$ with respect to M orbitals, and $M-\text{L}\sigma$ with respect to the ligand.
 - π : $\text{Br}(p_x, p_y)$ will interact with $d_{xz,yz}$ in $M-\text{L}\pi^*$ interactions.
 - The formation of a bonding pair causes $\text{Br}(p_z)$ orbital stabilization, at the cost of the destabilization of the d_{z^2} orbital.
 - The metal p_z and d_{z^2} orbitals will be destabilized, but will remain below $\text{Co}-\text{CN}\sigma^*$ orbitals because the Co-Br interaction is not as great as the Co-CN interaction.

VI.6 Module 39: Metal-Metal Bonding

- First historical puzzle:
 - Copper (II) acetate has crystalizes with four O-CR-O bridges joining two square pyramidal copper atoms that are also bonded to one aqua group, each.
 - Copper is d^9 , so $\mu = 1.73$ in theory. However, we observe $\mu = 1.4$.
 - We resolve this conflict by noting that the copper atoms are not square pyramidal — they bond to each other, giving both an octahedral geometry.
 - This is important because it gives us a new pair of bonding and antibonding orbitals. We thus have low-spin and high-spin states, and the existence of a nonzero but not super high magnetic moment hints at a high-spin state with some preference for low-lyingness.
- Metal-metal bonding is common for metals in low oxidation state, and generally increases in strength along the series $3d << 4d < 5d$.
- Single metal-metal bonded complexes:
 - We will consider the paradiamagnetic single M-M complex $\text{M}_2(\text{CO})_{10}$.
 - Possible metals are Fe^{2+} , manganese, and uranium.
 - Strategy for group fragment approach: Correlate to $\text{M}(\text{CO})_6$ (O_h), remove a ligand to give $\text{M}(\text{CO})_5$, and dimerize.
 - Note that $\text{Mn}(\text{CO})_5$ is the inorganic analog to CH_3 , i.e., the $\text{Mn}(\text{CO})_5$ fragment is said to be isolobal with CH_3 .
 - The HOMO of $\text{Mn}(\text{CO})_5$ is a singly occupied d_{z^2} orbital.
 - The energetic stabilization of the electrons in d_{z^2} into a σ -bond by dimerization is the driving force for metal-metal bond formation.
 - Note that a diamagnetic complex is formed from the dimerization of two metallic radicals.
 - Also note that the d_{z^2} orbital is cylindrically symmetric, i.e., indicates no preference for the staggered vs. eclipsed conformation. Thus, the molecule will adopt the more torsionally favorable staggered D_{4d} configuration.
- Cluster formation:
 - As mentioned above, odd electron occupancy of the e_g orbitals (O_h) prompts ligand loss in order to destabilize d_{z^2} . Further stabilization occurs by metal-metal single bond formation. We can take limiting argument to explain cluster formation across the periodic table:

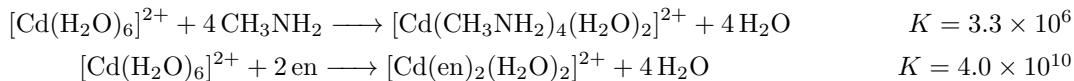
- In each case, the clusters assume an octahedral coordination as a result of burying 7 d -electrons in what are formally t_{2g} orbitals. The system loses a number of COs equivalent to the number of electrons in $M-L\sigma^*$; this permits maximum M–M bond formation and thus maximum stabilization.
- Clusters can trimerize or form higher polymers.
- Clusters can act as super atoms, combinations of atoms that you can add and remove electrons from with similar effects to changing the oxidation state in one atom.
- Multiple metal-metal bonded complexes:
 - Inorganic chemists synthesized a compound with a quadruple bond in 1964.
 - The MO strategy will be to correlate $\text{Re}_2\text{Cl}_8^{2-}$ with ReCl_6^{3-} (O_h), remove two axial Cl^- ligands to give a square planar ReCl_4^- fragment, and then dimerize.
 - Here, we see stabilization of the $d_{xz,yz}$ (the Cl^- ligand had been giving antibonding character) orbitals and massive stabilization of the d_{z^2} orbital (same reason).
 - The d_{z^2} orbital reacts the most to form an a_{1g} σ -bond.
 - The $d_{xz,yz}$ orbitals react the second most to form two e_u π -bonds.
 - The d_{xy} orbitals react the third most to form one b_{2g} δ -bond.
 - The eclipsed D_{4h} structure is a result of the δ bond.
 - Adding up our 8 bonding and 0 antibonding electrons and diving by two gives us our first quadruple bond.
- Metal-metal bonding vs. configuration: Increases single to quadruple bond as $d^1 \rightarrow d^4$; decreases triple to no bond as $d^5 \rightarrow d^8$.
- f orbitals are deep within the core of the atom and not generally available for bonding, so we will not see higher bonding because of f orbitals in lanthanides and actinides.
- The d^8 case:
 - No bonding predicted, but...
 - We actually see one-dimensional crystals, as in band theory.
 - The orbitals responsible are the d_{z^2} orbitals, which mix with the higher energy p_z orbitals to stabilize the d_{z^2} σ -bonding MO (and destabilize the p_z bonding MO, but this orbital is unfilled).
- The $d_{x^2-y^2}$ orbitals cannot participate in ML_4 coupling since they are used for ligand bonding.
 - Thus, the only orbitals available for coupling are $d_{z^2,xz,yz,xy}$.
- Inorganic chemists (including Phil Power, who is probably the greatest currently living main-group inorganic chemist) synthesized a compound with a quintuple bond in 2005.
 - They had to change the coordination environment around the metal center to free up the $d_{x^2-y^2}$ orbital. They did so with sterically bulky ligands that hindered more bonding.
 - The quintuple bonding was verified by observing a very short bond length (not super short on the scheme of things, but shorter than ever reported before, indicating the presence of a weak fifth bond).
 - C_{2h} symmetry.
 - Suggested reading: The full paper.

VI.7 Chapter 10: Coordination Chemistry II (Bonding)

From Miessler et al. (2014).

3/3:

- We seek to review experimental observations of coordination complexes and show how these observations have been reconciled with electronic structure and bonding theories.
- **Stability constant:** An equilibrium constant in a reaction that forms a coordination complex, typically by replacing one or more aqua ligands with another type. *Also known as formation constant.*
 - A measure of bond strength frequently used by inorganic chemists.
- Stronger field ligands preferably bond to the metal center.
- Polydentate ligands bond more readily than one split into chunks each with fewer attachment sites because bonding once makes bonding a second time easier (proximity), and ligand dissociation is more difficult.
 - Thus, for example, $[\text{Ni}(\text{en})_3]^{2+}$ is more stable than $[\text{Ni}(\text{CH}_3\text{NH}_2)_6]^{2+}$.
- The chelate effect has the largest impact on formation constants when the ring is 5-6 atoms long (shorter causes strain; longer means less proximity, hence tougher bond formation).
- “Thermodynamic parameters such as ΔH° , ΔS° , and the dependence of K with T are useful for comparing reactions of different metal ions reacting with the same ligand or a series of different ligands reacting with the same metal ion” (Miessler et al., 2014, p. 358).
 - Sometimes, properties of coordination complexes can be predicted directly off of this information, but it’s rare.
- An example of the thermodynamics of the chelate effect:



- Differences in ΔS° play the dominant role in the chelate effect. ΔH° is similar for the above two reactions, but for the en substitution, we dramatically increase entropy by binding two en particles and liberate four H_2O particles, as opposed to conserving the number of particles in the methylamine reaction.
- **Magnetic susceptibility:** A molecule’s attraction to or repulsion from a magnetic field. *Also known as χ .*
- $\mu = 2.828\sqrt{\chi T}$, where $2.828 \approx \sqrt{3k/N}$.
- “The total spin magnetic moment for a configuration of electrons is characterized by the spin quantum number S , which is equal to the maximum total spin, the sum of the m_s values” (Miessler et al., 2014, p. 360).
- “The orbital angular momentum [is] characterized by the quantum number L , where L is equal to the maximum possible sum of the m_l values for an electronic configuration” (Miessler et al., 2014, p. 360).
 - For a p^4 ground state electron configuration, L is maximized when two electrons occupy the orbital with $m_l = +1$, and the other two occupy the $m_l = 0, -1$ orbitals, respectively. Thus, $L_{p^4} = 1 + 1 + 0 - 1 = 1$.

- Technically, $\mu_S = 2.00023\sqrt{S(S+1)}$ where 2.00023 is the gyromagnetic ratio.
 - However, if we round the gyromagnetic ratio to 2 and note that $S = \frac{n}{2}$ where n is the number of unpaired electrons, we obtain $\mu_S = \sqrt{n(n+2)}$.

- Recall from Chapter 2 that in transition metals, ns electrons are lost before $(n - 1)d$ electrons.
- Number of unpaired electrons examples:
 - Fe is $4s^23d^6$, so 4 unpaired d -electrons.
 - Fe^{2+} is $3d^6$, so 4 unpaired d -electrons.
 - Cr is $4s^13d^5$, so 6 unpaired electrons (1 unpaired s -electron and 5 unpaired d -electrons).
 - Cu is $4s^13d^{10}$, so 1 unpaired s -electron.
 - Cu is $3d^9$, so 1 unpaired d -electron.
- The unit symbol for Bohr magnetons is μ_B , so we say, for instance, that the magnetic moment of copper is $1.73\ \mu_B$.
- Miessler et al. (2014) gives a full explanation of the Gouy method.
 - Modern magnetic susceptibility measurements are taken with a magnetic susceptibility balance (for solids), via the Evans NMR method (for solutes), or with a SQUID.
- **Mass susceptibility:** The magnetic susceptibility of a substance per gram.
- “A complex with one unpaired electron exhibits ideal **Curie paramagnetism** if the inverse of the molar susceptibility (for a given applied external field) increases linearly with temperature and has a y -intercept of 0” (Miessler et al., 2014, p. 362).
- “In a paramagnetic complex, the magnetic moments of individual species do not effectively couple, but act more or less independently of each other” (Miessler et al., 2014, p. 362).
 - This contrasts with ferromagnetism.
- **Antiferromagnetism:** An alternate long-range arrangement of magnetic moments where adjacent moments line up in opposite directions.
- Electronic spectra and molecular shapes provide further evidence for electronic structures.
- There are three main bonding models for coordination complexes.
- **Crystal field theory:** This is an electrostatic approach, used to describe the split in metal d -orbital energies within an octahedral environment. It provides an approximate description of the electronic energy levels often responsible for the ultraviolet and visible spectra of coordination complexes, but it does not describe metal-ligand bonding.
- **Ligand field theory:** This is a description of bonding in terms of the interactions between metal and ligand frontier orbitals to form molecular orbitals. It uses some crystal field theory terminology but focuses on orbital interactions rather than attractions between ions.
- **Angular overlap method:** This is a method of estimating the relative magnitudes of molecular orbital energies within coordination complexes. It explicitly takes into account the orbitals responsible for ligand binding as well as the relative orientation of the frontier orbitals.
- **Born-Oppenheimer approximation:** Atomic nuclei are in fixed positions in comparison with rapidly moving electrons.
- **Extended Hückel theory:** The simplest approach to approximating the many-body problem of electronic movements.

- Thoughts on CFT:

– “When the d orbitals of a metal ion are placed in an octahedral field of ligand electron pairs, any electrons in these orbitals are repelled by the field. As a result, the $d_{x^2-y^2}$ and d_{z^2} orbitals, which have e_g symmetry, are directed at the surrounding ligands and are raised in energy. The d_{xy} , d_{xz} , and d_{yz} orbitals (t_{2g} symmetry), directed between the ligands, are relatively unaffected by the field” (Miessler et al., 2014, p. 364).

– In Δ_o , the o stands for octahedral.

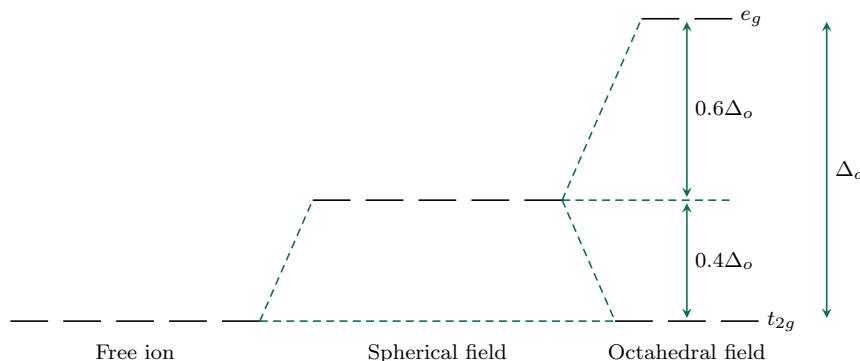


Figure VI.3: Crystal field splitting from the free ion.

- “The average energy of the five d orbitals [in a coordination complex] is above that of the free ion orbitals [i.e., those in the metal ion before it is surrounded by ligands], because the electrostatic field of the ligands raises their energy” (Miessler et al., 2014, p. 364).
- CFT is flawed because as we can see from Figure VI.3, there is no energetically favorable bonding; at best, we could have zero energy difference between the free ion and bonded octahedral field state (this occurs if only the t_{2g} orbitals are occupied).

3/4:

- Ligand field theory was synthesized from CFT and MO theory by Griffith and Orgel.
- The ligand HOMO often serves as the basis for group orbitals.
- Based on their shapes, it makes sense that the symmetry of the $d_{xy,xz,yz}$ orbitals doesn't match that of the ligand σ orbitals.
- The colors of transition metal complexes are often due to the energy difference between the t_{2g} and e_g^* orbitals, which is often equal to the energy of photons of visible light.
 - $[\text{Ni}(\text{H}_2\text{O})_6]^{2+}$, for example, is green because Δ_o matches well with the energy of red light, so when white light passes through a solution of this compound, red light is absorbed and the complement (green light) passes through to be perceived.
- “ Δ_o is also used in ligand field theory as a measure of the magnitude of metal-ligand interactions” (Miessler et al., 2014, p. 367).
- Perhaps the orbital labeling scheme in Figure V.10 isn't actually procedural; it's just a consistent reference frame.
- The interaction of ligand $\pi(t_{1u})$ orbitals with metal $p(T_{1u})$ orbitals is relatively weak since metal p orbitals largely interact with ligand $\sigma(t_{1u})$ orbitals.
- “Strong π acceptor ligands have the ability to increase the magnitude of Δ_o by lowering the energy of the t_{2g} orbitals” (Miessler et al., 2014, p. 368).

- “Electrons in the lower bonding orbitals are largely concentrated on the ligands. It is the stabilization of these ligand electrons that is primarily responsible for why these ligands bind to the metal center” (Miessler et al., 2014, p. 368).
- Electrons in higher energy orbitals are affected by ligand field effects and determine “structural details, magnetic properties, electronic spectrum absorptions, and coordination complex reactivity” (Miessler et al., 2014, p. 368).
- Metal-to-ligand ($M \longrightarrow L$) π bonding:** Electron donation from metal $d_{xy,xz,yz}$ orbitals to ligand π^* orbitals. *Also known as π back-bonding.*
- Δ_o is the splitting between the *highest occupied* t_{2g} (which can be either t_{2g} or t_{2g}^*) orbital and the e_g^* orbital.

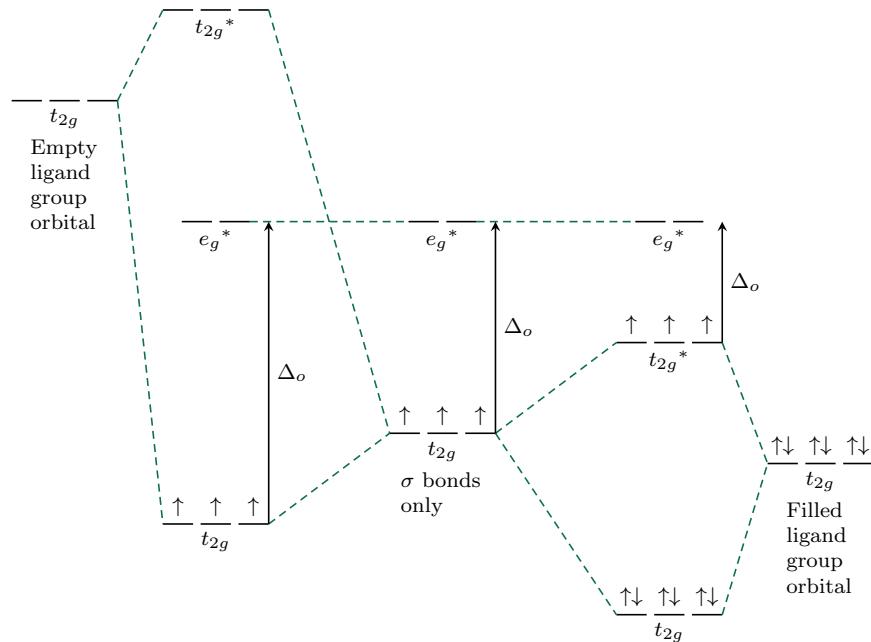


Figure VI.4: Effects of π bonding on Δ_o using a d^3 ion.

- This is why we consider Δ_o to be the energy difference between the t_{2g} and e_g^* orbitals in complexes with π -accepting ligands (where the t_{2g} orbitals are filled and lie beneath the e_g^* orbitals), but the difference between the t_{2g}^* and e_g^* orbitals in complexes with π -donating ligands (where the t_{2g} and t_{2g}^* are filled and both lie beneath the e_g^* orbitals).
- Empty ligand group orbitals engage in $M \longrightarrow L$ bonding while filled ligand group orbitals engage in $L \longrightarrow M$ bonding.
- Ligand-to-metal ($L \longrightarrow M$) π bonding:** Electron donation from the ligand π orbitals to metal $d_{xy,xz,yz}$ orbitals.
- $L \longrightarrow M$ bonding favors high-spin configurations while $M \longrightarrow L$ bonding favors low-spin configurations.
- Part of the stabilization of $M \longrightarrow L$ bonding comes from donation of excess electron density on the metal cation resulting from σ acceptance back to the ligands (metals prefer to be relatively positive because of their low electronegativity).
- Note that because backbonding orbitals are composed of π^* orbitals, filling of them results in weakening of the ligand π bond.

- “In octahedral coordination complexes, electrons from the ligands fill all six bonding molecular orbitals, and the metal valence electrons occupy the t_{2g} and e_g^* orbitals” (Miessler et al., 2014, p. 372).
- Strong field ligands interact more strongly with metal orbitals, leading to greater stabilization of the t_{2g} set, leading to a larger Δ_o and low-spin configurations.
 - Weak field ligands interact more weakly with metal orbitals, leading to lesser stabilization of the t_{2g} set, leading to a smaller Δ_o and low-spin configurations.
- Recall that increases in Π_c *destabilize* the complex, whereas increases in Π_e *stabilize* the complex.
- Is there any case in which we would have a d^5 configuration with four t_{2g} electrons and one e_g^* electron?
- Co^{3+} is the third row cation that most easily adopts a low-spin configuration.
- Magnitudes of Δ_o values:
 - Values for d^5 ions are smaller than those of d^4 and d^6 ions.
 - Increases in the principal quantum number n increase Δ_o because of two effects:
 - There is greater overlap between the larger $4d$ and $5d$ orbitals and the ligand orbitals.
 - There is a decreased pairing energy due to a larger available volume for electrons in the $4d$ and $5d$ orbitals relative to with $3d$ orbitals.
- **Ligand field stabilization energy:** The difference^[1] between the energy of the t_{2g} / e_g electronic configuration resulting from the ligand field splitting and the hypothetical energy of the t_{2g} / e_g electronic configuration with all five orbitals degenerate and equally populated. *Also known as LFSE.*
- Note that LFSE can only be calculated using the $\frac{2}{5}$ / $\frac{3}{5}$ method for σ -bonded octahedral complexes.
- Although we might expect transition metal ions to have linearly^[2] increasingly exothermic ΔH_s for hydration reactions (see below) across a period (as decreasing atomic radius makes them more concentrated [hence more unstable] sources of positive charge), we observe two loops similar to Figure V.6 but inverted.



- This is because the LFSE provides extra stabilization.
- It is also because of spin-orbit coupling, a relaxation effect caused by contraction of the metal-ligand distance, and an interelectronic repulsion energy (see Chapter 11 for a more detailed description of the first and third phenomena).
- Small corrections must be made in cases where the J-T distortion is present.
- **Nephelauxetic effect:** The magnitude of the reduction in the interelectron repulsion energy between the free ion and coordination complex.
 - Is this due to increases in delocalization?
 - Factors that increase the nephelauxetic effect:
 - Softness of the ligand.
 - Increase in metal oxidation state magnitude.

¹This is equivalent to the CFSE defined in the lecture.

²Errata: Miessler et al. (2014) refers to the linear line in the graph as “blue,” even though it should be green (a relic from the previous edition).

Topic VII

Band Theory in Solids

VII.1 Module 21: Electronic Structure of Solids (1D Solids)

2/5:

- Solid silicon's symmetry space group would be $Fd\bar{3}m$.
- Suggested reading: Hoffmann (1987).
 - A rare less mathematical article on band theory that has taught generations of chemists; comes at it from a physics perspective.
- To consider solids, let's first consider an infinite chain of hydrogen atoms.
 - This should separate into H_2 molecules (**Peierls's instability**); considering it to exist is a consequence of looking at chemistry through a physics perspective, where this is a simple model.
 - However, other substances can have chains of p_z orbitals, such as platinum atoms.
- An imaginary zoo of hydrogen molecules (we use the limit of a cycle of hydrogen atoms to approximate an infinitely long chain):
 - H_2 has a bonding and antibonding MO.
 - Cyclic H_3^+ is the most abundant ion in the universe (recently discovered by UChicago). One bonding and two antibonding orbitals.
 - We can keep adding hydrogen atoms to our rings.
 - For an infinitely long cycle of hydrogen atoms, we will have an infinite number of states close together that resembles a band in solids.
- Back to the chain of H atoms:
 - The basis function on each lattice point is a H_{1s} orbital; there are countably many.
 - The appropriate SALCs ψ_k are based in translating every orbital by a finite number of units:

$$\psi_k = \sum_n e^{ikna} \phi_n$$

- a is the distance between neighboring hydrogen atoms.
- Since there are infinitely many translations, there should be infinitely many translational symmetry elements, so infinitely many irreducible representations, too.
- The coefficients e^{ikna} come from **Bloch's theorem**.
 - In this formalism, k is an index labeling irreducible representations of the translation group. ψ transforms just like a , e_1 , and e_2 (e.g., in the C_5 point symmetry group).
 - This process of symmetry adaptation is called “forming Bloch functions.”

- Elementary band theory for extended solids:
 - Energy bands in solids arise from overlapping atomic orbitals, which become the **crystal orbitals** that make up the bands.
 - Recipe: Use LCAO (tight binding) approach.
 - A crystal is a regular periodic array with translational symmetry.
 - Periodic boundary conditions require $\psi(x+Na) = \psi(x)$, i.e., each wavefunction must be symmetry equivalent to the one in the neighboring cells.
 - For a 1D solid with lattice constant a and atom index n , Bloch's theorem tells us that the above SALC ψ_k is a solution to the Schrödinger equation.
 - If we calculate ψ_0 and $\psi_{\pi/a}$, we get the most and least bonding states possible, respectively (the least bonding state is the most antibonding state and has the highest energy).



Figure VII.1: *s* orbital bonding states.

$$\psi_0 = \phi_0 + \phi_1 + \phi_2 + \phi_3 + \dots$$

$$\psi_{\pi/a} = \phi_0 - \phi_1 + \phi_2 - \phi_3 + \dots$$

- At this point, we can construct a band between these two states.
 - The band is *almost* infinite; it's on the order of Avogadro's number.
 - We have as many k values as translations in the crystal or as many unit cells in a crystal.
 - **First Brillouin zone:** The region that covers all possible energy states that the crystal can have.
 - It is $-\frac{\pi}{a} < k < \frac{\pi}{a}$; which is the range of all possible values that the sine function will give.
 - There is one energy level for each value of k , but $E(k) = E(-k)$.
 - k is proportional to the electron momentum, or electron velocity.
 - Calculation of 1D band structure:
 - We have N atoms such that $\psi_k = \sum_{n=0}^N e^{inka} \phi_n$.
 - The crystal Schrödinger equation is $\hat{H}\Psi(k) = E(k)\Psi(k)$.

$$E(k) = \langle \psi | \hat{H} | \psi \rangle$$

- Recall that in Dirac’s bra-ket notation, $\langle \psi | \hat{H} | \psi \rangle \equiv \int \psi^* \hat{H} \psi d\tau$; for normalized atomic orbitals

$$\langle \phi_m | \phi_n \rangle = \begin{cases} 1 & m = n \\ 0 & m \neq n \end{cases}$$

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$$\langle \psi | \psi \rangle = \sum e^{i(n-m)ka} \langle \phi_m | \phi_n \rangle = N$$

- Thus, we can calculate for on-site ($m = n$):

$$\langle \psi(k) | \hat{H} | \psi(k) \rangle = \sum_n \langle \phi_n | \hat{H} | \phi_n \rangle = N\alpha$$

And for resonance ($m \neq n$), where we need only consider the two nearest neighbors:

$$\left\langle e^{-inka} \phi_n | \hat{H} | e^{i(n\pm 1)ka} \phi_{n\pm 1} \right\rangle = \beta e^{\pm ika}$$

- Putting everything together, we have

$$E(k) = \frac{\langle \psi | \hat{H} | \psi \rangle}{\langle \psi | \psi \rangle} = \frac{N\alpha + N\beta(e^{ika} + e^{-ika})}{N} = \alpha + 2\beta \cos(ka)$$

- **Zone center:** The state where all atomic orbitals are in phase (all bonding σ). *Also known as Γ .*
- **Zone border:** The state where all atomic orbitals are out of phase (all antibonding σ^*). *Also known as X .*
- Large numbers of MOs form bands of states.
- **Band structure:** The plot of E as a function of k .
 - The one we've derived so far is an s-shape curve.
- The p -orbitals are opposite — they form a bonding state with inverted phases.



Figure VII.2: p orbital bonding states.

- The analysis of Figures VII.1 and VII.2 can be done for many more types of orbitals, including p_z , d_{z^2} , and d_{xz} .
- Bonding orbital bands run uphill (concave upwards $E(k)$) at $k = 0$ and antibonding orbital bands run downhill (concave downwards $E(k)$) at $k = 0$.
- Energy bands run from $\alpha + 2\beta$ to $\alpha - 2\beta$ since β is negative for s orbitals.
- **Density of states:** The number of energy levels in the energy interval ΔE . *Also known as DOS.*

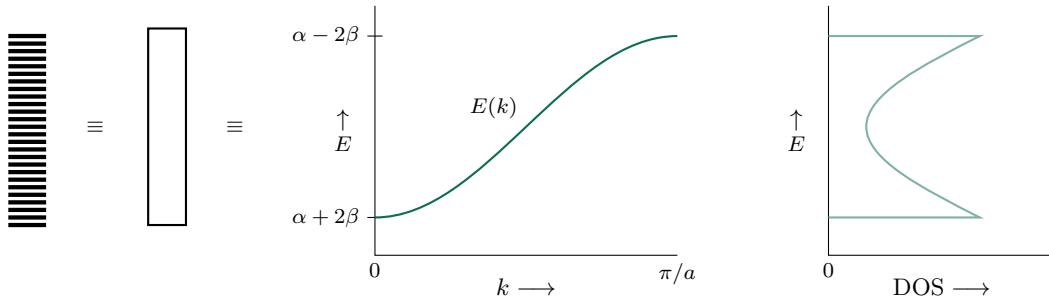


Figure VII.3: Density of states.

- Proportional to the inverse slope of the band; steep bands with large overlap yield a small DOS, and vice versa for flat bands.
- Reality check: PES for a long-chain alkane ($C_{36}H_{74}$) shows this inverse DOS relationship for a little while.

VII.2 Module 22: Electronic Structure of Solids (2D and 3D solids)

- 2D band structure:

- Simple Hückel: A two-dimensional square net (s orbitals only (or p_z)).

$$\psi(k) = \sum_{m,n} e^{ik_x m a + ik_y n a} \cdot \phi_{m,n}$$

- Consider the **crystal orbitals** at special k points (high symmetry).
- The **Brillouin zone** is 2D here (we have a **wave vector**).

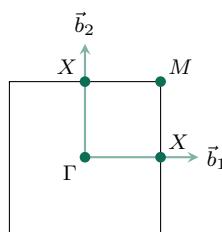


Figure VII.4: 2D Brillouin zone.

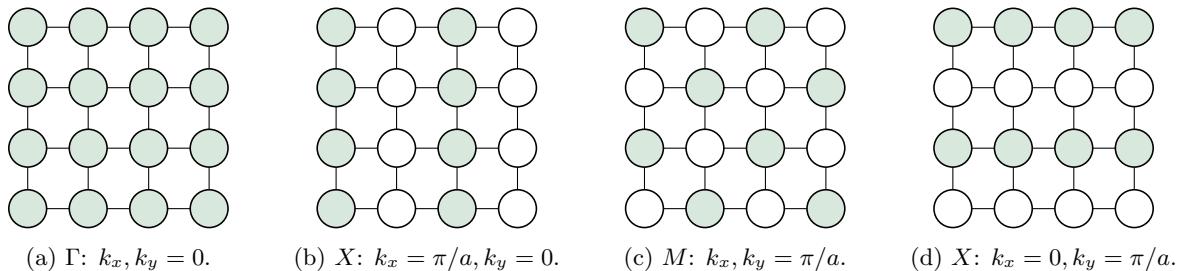


Figure VII.5: Special k points.

- The center is the Γ point ($k_x = k_y = 0$; every orbital is surrounded by 4 orbitals of matching phase). The midpoint of the lines are called X points ($k_x = \frac{\pi}{a}, k_y = 0$, and vice versa; every orbital is surrounded by 2 orbitals of matching phase and 2 orbitals of unlike phase). The maximum point is the M point ($k_x = k_y = \frac{\pi}{a}$; every orbital is surrounded by 4 orbitals of unlike phase).
- Calculating $E(k)$ in two dimensions.

$$E(k) = \alpha + 2\beta(\cos(k_x a) + \cos(k_y a))$$

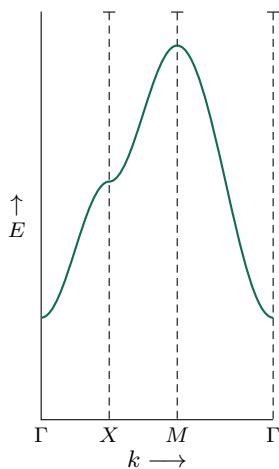


Figure VII.6: Schematic band structure (2D).

- Our schematic band structure (Figure VII.6) traces values along a 1D path in two-space from $\Gamma \rightarrow X \rightarrow M \rightarrow \Gamma$.
- 2/8:
- The bandwidth $4|\beta|$ is proportional to the degree of interaction between neighboring orbitals.
 - Since β is the interaction integral and $E(k)$ varies from $\alpha - 2\beta$ to $\alpha + 2\beta$.
 - For p_σ orbitals, $\beta > 0$.
 - Deriving the density of states formula:
 - We often simplify $E(k)$ with the first term of the Taylor series expansion; this gives us
$$E = \frac{\hbar}{2m}k^2$$
 - This implies that $E \propto k^2$ and, hence, $k \propto \sqrt{E}$.
 - Thus, the one-dimensional density $D_{1d}(k)$ of states as a function of k is $dN(k)/dk = 1$ since the number of states is evenly distributed along the k axis (i.e., in Figure VII.3).
 - It follows that the one-dimensional density $D_{1d}(E)$ of states as a function of E is
$$D_{1d}(E) = \frac{dN(E)}{dE} = \frac{dN(k)}{dk} \frac{dk}{dE} \propto 1 \cdot \frac{1}{\sqrt{E}} = \frac{1}{\sqrt{E}}$$
 - For each orbital, there is a unique path akin to Figure VII.6. The combination of all of these **bands** in one graph characterizes a material.
 - **Wigner-Seitz cell** (of the reciprocal lattice): The first Brillouin zone, or FBZ.
 - A primitive cell with a lattice point at its center.
 - A 3D discrete Fourier transform of the lattice.
 - Has $k_{x,y,z}$.
 - What is “d.I.” and “r.I.”?
 - We once again can find high symmetry points and directions akin to those in Figure VII.4.
 - Electronic band structure is calculated within the first Brillouin zone to give us the electronic band structure of a solid.

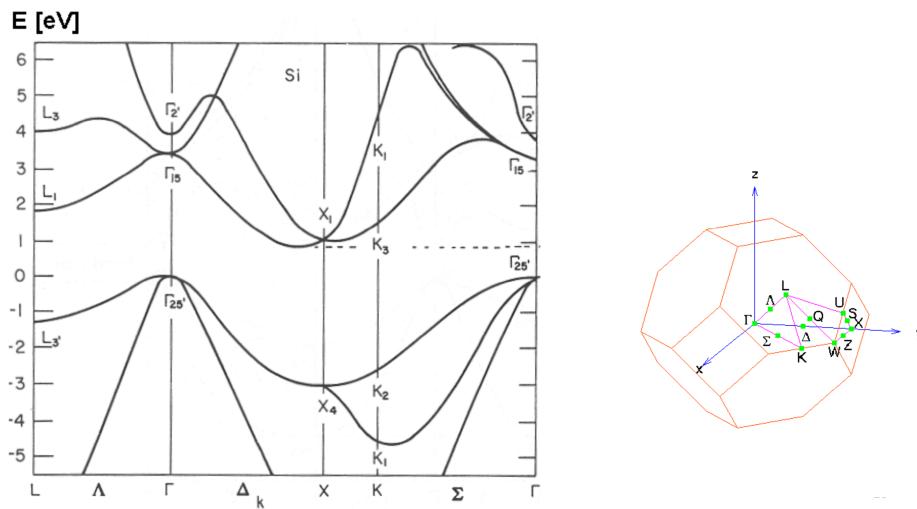


Figure VII.7: Electronic band structure of Si.

- **Angle-resolved photoemission spectroscopy^[1]:** If the incoming photon's energy is greater than the electron's binding energy, the electron will eventually be emitted with a characteristic kinetic energy and angle relative to the surface normal. This angle is related to the electron's crystal momentum. The Bloch wave vector is linked to the measured electron's momentum. *Also known as ARPES.*

– Indeed, ARPES can be used to reconstruct the band structure of a solid. The bands are real!

VII.3 Module 23: Filling Bands With Electrons

- The Fermi-Dirac statistics and Fermi Energy:
 - At $T = 0$, we expect all of the atoms in a solid to be in the ground state. The distribution of electrons (fermions) at the various energy levels is governed by the Fermi-Dirac distribution:

$$F(E) = \frac{1}{1 + \exp\left(\frac{E - E_F}{kT}\right)}$$

where E_F is the Fermi energy.

- $F(E)$ is the probability to fill the states with a given energy E .
- When $T = 0$ K, the Fermi energy is the energy of the last occupied state. Moreover,
- The Fermi energy is the energy of the last occupied state at $T = 0$ K; it is proportional to the square of the Fermi state k_F , i.e., $E_F \propto k_F^2$.
- If $T > 0$, then:
 - We fill the states from bottom to top.
 - Instead of having a sharp shift from occupied to unoccupied states, we have a sort-of washed-out step function.
 - Far below E_F , $F(E) = 1$; far above E_F , $F(E) = 0$. In the small **Fermi window** (aka. **Fermi level**) at the border (where the washed-out step function is), $0 < F(E) < 1$.
 - The Fermi window is $4kbt$.
 - It is in the Fermi level that all of the important stuff happens (i.e., electrons flowing in metals).

¹Figure 7.20 in Labalme (2020a) actually refers to this kind of photoelectron spectroscopy!

- If the gap (range where $\text{DOS}(E) = 0$) in the density of states at the Fermi level is smaller than 3 eV, then we have a semiconductor. If larger, we have an insulator. If 0, we have a conductor. Magnitude of DOS at Fermi level correlates with conductivity (e.g., Al has a higher DOS at the Fermi level than Ag, and we observe that Al is more conductive than Ag).
 - In metals and insulators, the Fermi level is within bands.
 - In semiconductors, it is between bands.
 - In physics, everything is a metal or insulator; semiconductors are a constructed perspective.
- Fermi sphere:
 - The surface of the Fermi sphere separates occupied and unoccupied states in k -space.
 - Bounded by Fermi surface.
 - Radius is Fermi wave vector $k_F = \sqrt[3]{3\pi^2 n}$ where n is a parameter related to the density of electrons.
 - Fermi energy: $E_F = \hbar^2 k_F^2 / 2m$.
 - Fermi momentum: $p_F = \hbar k_F$.
 - Fermi velocity: $v_F = \hbar k_F / m$.
 - Fermi temperature: $T_F = E_F / k_B$.
- Electrons at $T = 0$ K still move very quickly (approximately 0.06 the speed of light) since they're quantum particles (not classical ones).
- 3D density of states:
 - The density of states $g(E)$ is the number of one-electron states (including spin multiplicity) per unit energy and volume:

$$g(E)_{3D} \equiv \frac{1}{V} \frac{dN}{dE}$$
 - N is twice the product of the Fermi sphere volume and the number of levels per unit volume.
 - Thus,
$$N = 2 \times \frac{4}{3}\pi k^3 \times \frac{V}{8\pi^3} = \frac{V}{3\pi^2 \hbar^3} (2m^* E)^{3/2}$$

so

$$g(E)_{3D} = \frac{1}{V} \frac{dN}{dE} = \frac{1}{2\pi^2} \left(\frac{2m^*}{\hbar^2} \right)^{3/2} \sqrt{E}$$
 - Take CHEM 39000 Solids, Materials, and Surfaces to learn more.

Topic VIII

Electronic Spectra of Coordination Compounds

VIII.1 Module 40: Electronic Transitions

3/1:

- Suggested reading: Chapter 11.1.
- Transition metal complexes are known to show rich photophysics and optical properties.
 - For example, $[\text{Ni}(\text{NH}_3)_6]^{2+}$ has peaks in the infrared, visible, and UV spectra.
- How electronic transitions occur:
 - Take a solid or aqueous sample and illuminate it with photons of a particular power/intensity P_0 through $l\text{cm}$ of it.
 - Some will be absorbed and some will pass through. Measure the power P that comes out on the other side.
- Transmittance $T = P/P_0$.
- Absorbance $A = -\log T = \log \frac{P_0}{P}$.
- The Beer-Lambert Law: $A = \varepsilon Cl$, where C is the concentration of the sample in solution, l is the path length (the length through solution), and ε is the molar absorption coefficient.
- Plotting the wavelength of the impinging photons vs. ε gives us a graph with peaks, where each peak corresponds to an electron transition.
- Spectral features:
 - Number of transitions.
 - Energy of the transitions.
 - Intensity of the transitions.
 - Shape of the transition.
- **Transition probability:** The probability of a particular transition taking place.
- The transition probability depends on:
 - Energy of the transition vs. incident light.
 - Orientation of the molecule/material.
 - Symmetry of the initial and final states.

- Angular momentum (spin).
- The absorption spectra of various hexaaqua complexes of the first-row transition metals give us a zoo of spectra.
- We usually have $\varepsilon < 10$, which means faint colors.
- Types of molecular transitions:
 - Metal Centered (MC): Transitions between the d -orbitals on the metal center.
 - Ligand to Metal Charge Transfer (LMCT): For example, MnO_4^- has $\varepsilon \approx 10\,000$.
 - Metal to Ligand Charge Transfer (MLCT) and Metal to Metal Charge Transfer (MMCT), too.
- The transition probability of one molecule from one state Ψ_1 to another state Ψ_2 is given by $|\vec{M}_{21}|$, the transition dipole moment or transition moment from Ψ_1 to Ψ_2 .
 - The transition matrix element $\vec{M}_{21} = \int \Psi_2 \vec{\mu} \Psi_1 d\tau$, where $\vec{\mu}$ is the electric dipole moment operator $\vec{\mu} = \sum_n Q_n \vec{x}_n$, where Q_n is charge and \vec{x}_n is the position vector operator.
 - Derived with time-dependent perturbation theory.
 - For an electronic transition to be allowed, the transition moment integral must be nonzero.
 - Note that $\varepsilon \approx \vec{M}_{21}$.
- How the HOMO moves about the molecule depends on the type of incoming light.
 - If $\vec{M}_{21} = 0$, then the transition probability is 0 and the transition from Ψ_1 to Ψ_2 is forbidden or electric-dipole forbidden ($\varepsilon = 0$).
 - If $\vec{M}_{21} \neq 0$, then the transition probability is not 0 and the transition from Ψ_1 to Ψ_2 is not forbidden ($\varepsilon \geq 0$).
 - If $\vec{M}_{21} \neq 0$, we do not definitively know that there will be an electron transition or know how intense it will be; we just know that it is not electric-dipole forbidden.
- Calculating \vec{M}_{21} :
 - Use the same procedure with $\Gamma_2 \otimes \Gamma_\mu \otimes \Gamma_1$ as in Module 12.
 - If the direct product does not contain the totally symmetric representation, then the transition is forbidden by symmetry arguments.
 - If the direct product does contain the totally symmetric representation, then the transition is allowed by symmetry arguments.

C_{3v}	A₁	A₂	E
A₁	A_1	A_2	E
A₂		A_1	E
E			$A_1 + [A_2] + E$

Table VIII.1: Direct product table for the C_{3v} point group.

- Be aware of direct product tables, such as the above example, which we may readily obtain from Table III.1.
- Example: In a D_{2h} complex, can we excite a d_{z^2} electron to the p_z orbital?
 - From the D_{2h} character table, we have that $\Gamma_1 = A_g$ and $\Gamma_2 = B_{1u}$. We also have that Γ_μ for an x -, y -, and z -basis is B_{3u} , B_{2u} , and B_{1u} , respectively.

- Taking direct products under each basis gives us

$$\begin{aligned} B_{1u} \otimes B_{3u} \otimes A_g &= B_{2g} && (x\text{-basis}) \\ B_{1u} \otimes B_{2u} \otimes A_g &= B_{3g} && (y\text{-basis}) \\ B_{1u} \otimes B_{1u} \otimes A_g &= A_g && (z\text{-basis}) \end{aligned}$$

- Thus, the x - and y -components are forbidden while the z one is not.
- What this means is that z -plane polarized light will be able to cause the desired electron transition, but x - and y -plane polarized light will not.

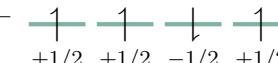
- We can use the same procedure to prove that we can never promote an electron from d_{xy} to p_z .
- We can use the same procedure for octahedral complexes, except the calculations of the direct products are just a bit more difficult.
 - For a d^1 complex, we calculate $E_g \otimes T_{1u} \otimes T_{2g}$.
 - For a d^6 complex, we calculate $(T_{2g} \otimes E_g) \otimes T_{1u} \otimes A_{1g} = (T_{1g} + T_{2g}) \otimes T_{1u} \otimes A_{1g}$.
 - A low spin d^6 complex has A_{1g} symmetry by taking the direct product of T_{2g} times itself six times.
 - The excited state has T_{2g} times itself five times, and then times E_g .
 - Basically, we take the direct product of the orbital that each electron occupies.

VIII.2 Module 41: Many Electron States

- Suggested reading: Chapter 11.2.
- For octahedral d^3 , we have multiple excited states (six, to be exact).
 - Fortunately, there is an easier way to describe transitions between states (we will talk about this next time).
- A single electron is completely described by the principal quantum number n , its angular momentum ℓ , its magnetic quantum number m_ℓ , and its spin m_s .
- Multielectron states are described by **Russell-Saunders coupling**, also known as **LS coupling**, **L-S coupling**.
- For example, consider the d^2 configured V^{3+} ion.
 - There are 45 different possible microstates. Some will have the same energy, some will not.
 - There are five states (denoted by **term symbols**) with distinct energy in total.
- To find the term symbol, we need:
 - $L = \text{total orbital angular momentum} = \sum m_\ell$.
 - $S = \text{total spin angular momentum} = \sum m_s$.
- Term symbols then are of the form $^{2S+1}L_J$

where $2S + 1$ is the spin multiplicity, L is the subshell letter corresponding to the angular momentum quantum number ($0 \leftrightarrow s$, $1 \leftrightarrow p$, $2 \leftrightarrow d$, $3 \leftrightarrow f$, \dots), and J is the total angular momentum ($J = L + S, L + S - 1, L + S - 2, \dots, |L - S|$, the spin-orbit coupling).

- Some examples:

-  : $S = \frac{1}{2}$, so our term symbol will be of the form $^2L_J^{[1]}$.
-  : $S = 1$, so our term symbol will be of the form 3L_J .
-  : $S = 1$ and $L = 3$, so our term symbol will be of the form $^3F_J^{[2]}$.
-  : $S = 1$ and $L = 0$, so our term symbol will be of the form $^3S_J^{[3]}$.

- Spin-orbit coupling is weak for a carbon atom (we can essentially just disregard it).
 - For lanthanides, it becomes very significant.
- 3/3: • In term symbols, J is typically a small correction to the energy for light elements.
 - It becomes significant with the lanthanides and actinides.
- We group electrons into terms because each state has a characteristic energy in the absence of external electric and magnetic fields.
 - If we do apply an external magnetic fields, states will split into $2S + 1$ substates.
 - Each term includes multiple microstates, which are atomic states produced by interactions of the atom with a magnetic field.
- Consider the examples from above. The third and fourth examples are very similar but give rise to two different terms. This is because if we have ions with these states, we will observe differing energy levels spectroscopically.
- What if we have one electron in the orbital with $m_\ell = -2$? Then how do we create the term symbol?
 - It is part of one of the microstates of the 1D term.
 - We fill electrons starting with the lowest energy states, and the lowest energy state has the greatest spin multiplicity (e.g., +2; this is one of Hund's rules).
- Correspondences between d electron configurations and term symbols:

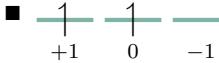
$$\begin{array}{ll}
 d^1 = ^2D & d^6 = ^5D \\
 d^2 = ^3F & d^7 = ^4F \\
 d^3 = ^4F & d^8 = ^3F \\
 d^4 = ^5D & d^9 = ^2D \\
 d^5 = ^6S & d^{10} = ^1S
 \end{array}$$

- Note that a fully occupied or unoccupied subshell is always 1S .
- Now if we take a direct product of a singlet S state with any other state, we will have just that state left over. This implies that we can ignore all fully occupied shells, and the term will be determined entirely by partially filled ones.
- If we want to build the full picture and see all possible states, we need a **microstate table**.
- Microstate table:

¹Pronounced “doublet state.”

²Pronounced “triplet eff state.”

³Pronounced “triplet ess state.”

- A microstate table contains all possible combinations of m_ℓ and m_s .
- Each microstate represents a possible electron configuration.
- It includes both ground and excited states.
- It must obey the Pauli exclusion principle.
- Example (p^2 electron configuration):
 - Microstate notation:
 - For each of the n electrons in the configuration, list a special symbol in an n -tuple.
 - The special symbol will be a number (m_ℓ) with either a + or – sign as an exponent (for positive and negative spin, respectively).
 - There are three ground state configurations.
 -  : The microstate is $(1^+, 0^+, -1^-)$.
 -  : The microstate is $(0^+, -1^+, 1^-)$.
 -  : The microstate is $(1^+, -1^+, 0^-)$.
 - We will not show all excited state configurations, but we will show a few.
 -  : The microstate is $(1^+, 1^-, 0^-)$.
 -  : The microstate is $(0^+, 0^-, 1^-)$.
 -  : The microstate is $(-1^+, -1^-, 0^-)$.
 - We can now generate the microstate table.

		M_S		
		-1	0	$+1$
M_L	$+2$		$1^+ 1^-$	
	$+1$	$1^- 0^-$	$1^+ 0^-$ $1^- 0^+$	$1^+ 0^+$
	0	$-1^- 1^-$	$-1^+ 1^-$ $0^+ 0^-$ $-1^- 1^+$	$-1^+ 1^+$
	-1	$-1^- 0^-$	$-1^+ 0^-$ $-1^- 0^+$	$-1^+ 0^+$
	-2		$-1^+ - 1^-$	

Table VIII.2: Microstate table for a p^2 electron configuration.

- Each column represents a state with a given spin angular momentum.
- Each row represents the sum of the angular momentum of two electrons.
- To analyze the microstate label, rename each microstate with X .
 - Each term consists of multiple microstates of equivalent energies (similar to how we can have multiple degenerate orbitals of a given energy).

- We can lift degeneracy by applying an external electric (**Stark effect**) or magnetic field (**Ziemann effect**).
- Now each term consists of $(2L + 1)(2S + 1)$ states (this is the **double multiplicity formula**).
- Group energetically equivalent states:
 - First, focus on the term containing the states with the largest possible L (i.e., those with $+2$). We know that it contains $(2(2) + 1)(2(0) + 1) = 5$ microstates. Choose 5 microstates from the $M_S = 0$ column, one per M_L .
 - Next, focus on the term containing states with the next largest possible L (i.e., those with $+1$ and $M_S = +1$). We know that it contains $(2(1) + 1)(2(1) + 1) = 9$ microstates. Choose 9 microstates, one from each box in the square bounded by $(-1, -1)$ and $(+1, +1)$.
 - The remaining microstate in $M_S = M_L = 0$ forms its own term.
- Thus, our microstate table can be decomposed into three terms (1D , 3P , and 1S).
- Identifying relative energies with Hund's rules:
 - For a given electron configuration, the term with the greatest spin multiplicity lies lowest in energy (Hund's rule).
 - For a term with a given multiplicity, the greater the value of L , the lower the energy.
 - Note that the rules for predicting the ground state always work, but they may fail in predicting the order of energies for excited states.
- Thus, going back to our example, we have that energetically, $^3P < ^1D < ^1S$.
- Example (d^2 electron configuration):
 - In this example, Hund's rules do not provide accurate energy predictions (they would predict $^3F < ^3P < ^1G < ^1D < ^1S$, but in reality, $^3F < ^1D < ^3P < ^1G < ^1S$).
- **Electron-hole formalism:** The Russel-Saunders terms for d^n and d^{10-n} configurations are identical for $n = 0, \dots, 5$.
 - We can rationalize this by thinking of n electrons and $10 - n$ holes (or positrons) as related to $10 - n$ electrons and n holes.

VIII.3 Module 42: Many Electron States and Transitions in Co-ordination Compounds

- Suggested reading: Chapter 11.3.
- Ligand field dependence (d^1 system):
 - Degenerate symmetric field:
 - Absence of ligand field.
 - Free-ion term.
 - All d -orbitals are energetically equal.
 - If all d orbitals are degenerate, then we can put a single electron in any orbital and we will have a microstate of 2D .
 - Infinite O_h field:
 - Strong ligand field.
 - Coordination complexes.
 - d -orbitals are not degenerate (d_{z^2,x^2-y^2} have higher energy; $d_{xy,xz,yz}$ have lower energy).
 - In this case, it matters in which d orbital we put the electron.

- Real molecules:
 - We use a correlation diagram or Orgel diagram.
 - If ligand field strength is zero, we have the degenerate symmetric field. If it is at maximum strength, we have two distinct states (t_{2g} and e_g). Anywhere in between, the states are in between in energy, too (as we apply the ligand field, we split the state).
- Ligand field dependence (d^2 system):

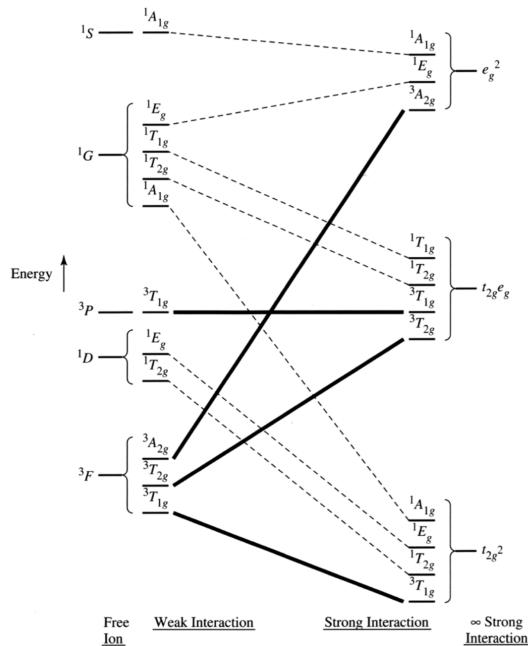


Figure VIII.1: Correlation diagram for a d^2 system.

- Under a weak ligand field, the multielectron states split.
 - From the O_h character table, we can determine how each term transforms similarly to determining how each orbital transforms (however, we may need a character table with cubic, quartic, and beyond functions).
 - For example, S terms transform as A_{1g} , P as T_{1g} , D as $T_{2g} + E_g$, F as $T_{1g} + T_{2g} + A_{2g}$, and G as $A_{1g} + E_g + T_{1g} + T_{2g}$.
- Under an infinitely strong field, we will return to our t_{2g} / e_g formalism with a low energy t_{2g}^2 state, a midlevel $t_{2g}e_g$ state (with one of the d^2 electrons in each level), and a high level e_g^2 state.
 - From the O_h group multiplication table, we can determine how each state splits under strong but finite ligand fields.
 - For example, $E_g \otimes E_g = A_{1g} + E_g + A_{2g}$, so that's why the e_g^2 state coalesces from these three states.
- The bold lines in Figure VIII.1 correspond to triplet states, and the dashed lines to singlet states.
- The direct product of two irreducible representations of dimension 2 or higher is reducible to a sum of symmetric and antisymmetric irreducible representations, so named because their respective symmetry-adapted basis functions are either symmetric or antisymmetric with respect to exchange of the components. It is important to understand these properties when dealing with the electronic states produced by incompletely filled degenerate orbitals.

- The following expressions may be applied to determine the characters (χ) of the symmetric and antisymmetric components of a direct product. The characters of the symmetric irreducible representation(s) (χ^+) are given by

$$\chi^+(R) = \frac{1}{2} ([\chi(R)]^2 + \chi(R^2))$$

The characters of the antisymmetric irreducible representation(s) are given by

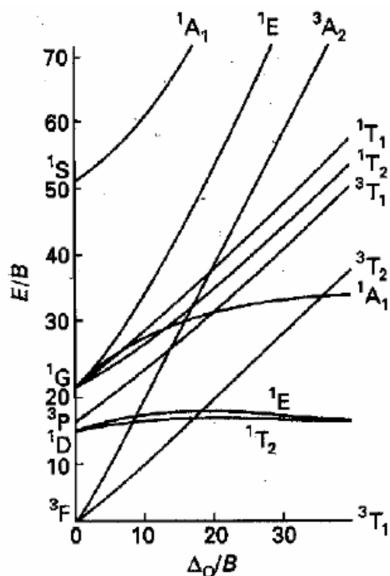
$$\chi^-(R) = \frac{1}{2} ([\chi(R)]^2 - \chi(R^2))$$

In these expressions, $\chi(R)$ is the character under symmetry operation R and $\chi(R^2)$ is the character associated with the operation R^2 .

- We do not need to know these formulas since deriving them is outside the scope of this course.
- For example, in the C_{4v} point group, the direct product $E \otimes E = A_1 + A_2 + B_1 + B_2$.
- These results are typically written showing the antisymmetric component in square brackets, i.e., $E \otimes E = A_1 + [A_2] + B_1 + B_2$.
- The electron wave function, which is a product of the orbital and spin wave function components, must be antisymmetric (see Pauli exclusion principle). Therefore, if the orbital component is symmetric, the spin one should be antisymmetric, i.e., singlet state. And vice versa.
- Back to our d^2 example:
 - Symmetric components have electrons with opposite spins; antisymmetric components have electrons with like spin.
 - This demonstrates how $T_2 \otimes T_2 = A_1 + E + [T_1] + T_2$ rationalizes the existence of three singlet states and one triplet state under t_{2g}^2 .
 - Correlation diagrams are immensely useful, and we can tediously construct them or find them in the textbook. However, we can nicely simplify them into Tanabe-Sugano diagrams:

Correlation Diagram	Tanabe-Sugano Diagram
Number of states.	Number of states.
General sense of field effects.	Field effects.
Only qualitative.	Quantitative.

- Tanabe-Sugano diagrams are designed to intuitively interpret optical spectra and electron transitions in transition metal complexes.
- In a Tanabe-Sugano diagram, we make the lowest line in the corresponding correlation diagram the x -axis/ground state and calculate energies of every state above with respect to the gap between the ground state and the upper state.
- Tanabe-Sugano diagrams show:
 - Relative energies of the states vs. ligand field strength.
 - Electronic states with the same symmetry never cross (**non-crossing rule**).
 - Curvature (1E and 3E).
 - Ground state on the x -axis; all other states are excited.
 - Transitions between states.
 - A Tanabe-Sugano diagram is a graph which plots the energy of different spectroscopic terms (on the y -axis) against the strength of the ligand field (on the x -axis). The units used for each are given in terms of the **Racah parameter B** .

Figure VIII.2: Tanabe-Sugano diagram for a d^2 system.

- The choice of unit means that the diagram takes account of electron-electron repulsion effects.
- The lowest energy state is usually placed along the x -axis.
- The electrostatic repulsion between electrons varies from atom to atom, depending upon the number and spin of the electrons and the orbitals they occupy. The total repulsion can be expressed in terms of three parameters A , B , and C , which are known as the Racah parameters. They are generally obtained empirically from gas-phase spectroscopic studies of atoms.
- Each particular coordination compound will be a vertical slice through its electron configuration's Tanabe-Sugano diagram.

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