For almost every covalent molecule that exists, we can now draw the Lewis structure, predict the electron-pair geometry, predict the molecular geometry, and come close to predicting bond angles. However, one of the most important molecules we know, the oxygen molecule $\text{O}_2$, presents a problem with respect to its Lewis structure. We would write the following Lewis structure for $\text{O}_2$:

$$\text{O} = \text{O}$$

This electronic structure adheres to all the rules governing Lewis theory. There is an $\text{O}=\text{O}$ double bond, and each oxygen atom has eight electrons around it. However, this picture is at odds with the magnetic behavior of oxygen. By itself, $\text{O}_2$ is not magnetic, but it is attracted to magnetic fields. Thus, when we pour liquid oxygen past a strong magnet, it collects between the poles of the magnet and defies gravity. Such attraction to a magnetic field is called paramagnetism, and it arises in molecules that have unpaired electrons. And yet, the Lewis structure of $\text{O}_2$ indicates that all electrons are paired. How do we account for this discrepancy?

Magnetic susceptibility measures the force experienced by a substance in a magnetic field. When we compare the weight of a sample to the weight measured in a magnetic field (Figure \((\text{PageIndex}1)\)), paramagnetic samples that are attracted to the magnet will appear heavier because of the force exerted by the magnetic field. We can calculate the number of unpaired electrons based on the increase in weight.

Experiments show that each $\text{O}_2$ molecule has two unpaired electrons. The Lewis-structure model does not predict the presence of these two unpaired electrons. Unlike oxygen, the apparent weight of most molecules decreases slightly in the presence of an inhomogeneous magnetic field. Materials in which all of the electrons are paired are diamagnetic and weakly repel a magnetic field. Paramagnetic and diamagnetic materials do not act as permanent magnets. Only in the presence of an applied magnetic field do they demonstrate attraction or repulsion.
Video \(\PageIndex{1}\): Water, like most molecules, contains all paired electrons. Living things contain a large percentage of water, so they demonstrate diamagnetic behavior. If you place a frog near a sufficiently large magnet, it will levitate. You can see videos of diamagnetic floating frogs, strawberries, and more ([https://www.youtube.com/watch?v=A1vyB-O5i6E](https://www.youtube.com/watch?v=A1vyB-O5i6E)).

Molecular orbital theory (MO theory) provides an explanation of chemical bonding that accounts for the paramagnetism of the oxygen molecule. It also explains the bonding in a number of other molecules, such as violations of the octet rule and more molecules with more complicated bonding (beyond the scope of this text) that are difficult to describe with Lewis structures. Additionally, it provides a model for describing the energies of electrons in a molecule and the probable location of these electrons. Unlike valence bond theory, which uses hybrid orbitals that are assigned to one specific atom, MO theory uses the combination of atomic orbitals to yield molecular orbitals that are delocalized over the entire molecule rather than being localized on its constituent atoms. MO theory also helps us understand why some substances are electrical conductors, others are semiconductors, and still others are insulators. Table \(\PageIndex{1}\) summarizes the main points of the two complementary bonding theories. Both theories provide different, useful ways of describing molecular structure.

<table>
<thead>
<tr>
<th>Valence Bond Theory</th>
<th>Molecular Orbital Theory</th>
</tr>
</thead>
<tbody>
<tr>
<td>considers bonds as localized between one pair of atoms</td>
<td>considers electrons delocalized throughout the entire molecule</td>
</tr>
<tr>
<td>creates bonds from overlap of atomic orbitals ((s, p, d\ldots)) and hybrid orbitals ((sp, sp^2, sp^3\ldots))</td>
<td>combines atomic orbitals to form molecular orbitals ((\sigma, \sigma^<em>, \pi, \pi^</em>))</td>
</tr>
</tbody>
</table>
Valence Bond Theory
forms σ or π bonds
predicts molecular shape based on the number of regions of electron density
needs multiple structures to describe resonance

Molecular Orbital Theory
creates bonding and antibonding interactions based on which orbitals are filled
predicts the arrangement of electrons in molecules

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<td></td>
</tr>
</tbody>
</table>

**Molecular Orbitals Formed from ns and np Atomic Orbitals**

Atomic orbitals other than ns orbitals can also interact to form molecular orbitals. Because individual p, d, and f orbitals are not spherically symmetrical, however, we need to define a coordinate system so we know which lobes are interacting in three-dimensional space. Recall that for each np subshell, for example, there are npₓ, npᵧ, and np𝑧 orbitals. All have the same energy and are therefore degenerate, but they have different spatial orientations.

\[
\sigma_{np_{z}} = np_{z}(A) - np_{z}(B) \quad (\text{Eq 6})
\]

Just as with ns orbitals, we can form molecular orbitals from np orbitals by taking their mathematical sum and difference. When two positive lobes with the appropriate spatial orientation overlap, as illustrated for two np𝑧 atomic orbitals in part (a) in Figure 1, it is the mathematical difference of their wave functions that results in constructive interference, which in turn increases the electron probability density between the two atoms. The difference therefore corresponds to a molecular orbital called a \(\sigma_{np_{z}}\) bonding molecular orbital because, just as with the σ orbitals discussed previously, it is symmetrical about the internuclear axis (in this case, the z-axis):

\[
\sigma_{np_{z}} = np_{z}(A) - np_{z}(B) \quad (\text{Eq 6a})
\]

The other possible combination of the two np𝑧 orbitals is the mathematical sum:

\[
\sigma_{np_{z}} = np_{z}(A) + np_{z}(B) \quad (\text{Eq 7})
\]

In this combination, shown in part (b) in Figure 1, the positive lobe of one np𝑧 atomic orbital overlaps the negative lobe of the other, leading to destructive interference of the two waves and creating a node between the two atoms. Hence this is an antibonding molecular orbital. Because it, too, is symmetrical about the internuclear axis, this molecular orbital is called a \(\sigma_{np_{z}}\) antibonding molecular orbital.

Whenever orbitals combine, the bonding combination is always lower in energy (more stable) than the atomic orbitals from which it was derived, and the antibonding combination is higher in energy (less stable).
Figure \(\PageIndex{2}\): Formation of Molecular Orbitals from \(np_z\) Atomic Orbitals on Adjacent Atoms. (a) By convention, in a linear molecule or ion, the z-axis always corresponds to the internuclear axis, with +z to the right. As a result, the signs of the lobes of the \(np_z\) atomic orbitals on the two atoms alternate − + − +, from left to right. In this case, the σ (bonding) molecular orbital corresponds to the mathematical difference, in which the overlap of lobes with the same sign results in increased probability density between the nuclei. (b) In contrast, the σ* (antibonding) molecular orbital corresponds to the mathematical sum, in which the overlap of lobes with opposite signs results in a nodal plane of zero probability density perpendicular to the internuclear axis.

Overlap of atomic orbital lobes with the same sign produces a bonding molecular orbital, regardless of whether it corresponds to the sum or the difference of the atomic orbitals.

The remaining \(p\) orbitals on each of the two atoms, \(np_x\) and \(np_y\), do not point directly toward each other. Instead, they are perpendicular to the internuclear axis. If we arbitrarily label the axes as shown in Figure \(\PageIndex{3}\), we see that we have two pairs of \(np\) orbitals: the two \(np_x\) orbitals lying in the plane of the page, and two \(np_y\) orbitals perpendicular to the plane. Although these two pairs are equivalent in energy, the \(np_x\) orbital on one atom can interact with only the \(np_x\) orbital on the other, and the \(np_y\) orbital on one atom can interact with only the \(np_y\) on the other. These interactions are side-to-side rather than the head-to-head interactions characteristic of σ orbitals. Each pair of overlapping atomic orbitals again forms two molecular orbitals: one corresponds to the arithmetic sum of the two atomic orbitals and one to the difference. The sum of these side-to-side interactions increases the electron probability in the region above and below a line connecting the nuclei, so it is a bonding molecular orbital that is called a pi (\(\pi\)) orbital (a bonding molecular orbital formed from the side-to-side interactions of two or more parallel \(np\) atomic orbitals). The difference results in the overlap of orbital lobes with opposite signs, which produces a nodal plane perpendicular to the internuclear axis; hence it is an antibonding molecular orbital, called a pi star (\(\pi^*\)) orbital. An antibonding molecular orbital formed from the difference of the side-to-side interactions of two or more parallel \(np\) atomic orbitals, creating a nodal plane perpendicular to the internuclear axis.

\[
\pi_{np_x} = np_x(A) + np_x(B) \quad \text{(Eq 8)}
\]
\[
\pi^{\star}_{np_x} = np_x(A) - np_x(B) \quad \text{(Eq 9)}
\]

The two \(np_y\) orbitals can also combine using side-to-side interactions to produce a bonding \(\pi_{np_y}\) molecular orbital and an antibonding \(\pi^{\star}_{np_y}\) molecular orbital. Because the \(np_x\) and \(np_y\) atomic orbitals interact in the same way (side-to-side) and have the same energy, the \(\pi_{np_y}\) and \(\pi^{\star}_{np_y}\) molecular orbitals are a degenerate pair, as are the \(\pi^{\star}_{np_x}\) and \(\pi^{\star}_{np_y}\) molecular orbitals.

Figure \(\PageIndex{3}\): Formation of \(\pi\) Molecular Orbitals from \(np_x\) and \(np_y\) Atomic Orbitals on Adjacent Atoms. (a)
Because the signs of the lobes of both the \( np_x \) and the \( np_y \) atomic orbitals on adjacent atoms are the same, in both cases the mathematical sum corresponds to a \( \pi \) (bonding) molecular orbital. (b) In contrast, in both cases, the mathematical difference corresponds to a \( \pi^* \) (antibonding) molecular orbital, with a nodal plane of zero probability density perpendicular to the internuclear axis.

Figure \( \PageIndex{4} \) is an energy-level diagram that can be applied to two identical interacting atoms that have three \( np \) atomic orbitals each. There are six degenerate \( p \) atomic orbitals (three from each atom) that combine to form six molecular orbitals, three bonding and three antibonding. The bonding molecular orbitals are lower in energy than the atomic orbitals because of the increased stability associated with the formation of a bond. Conversely, the antibonding molecular orbitals are higher in energy, as shown. The energy difference between the \( \sigma \) and \( \sigma^* \) molecular orbitals is significantly greater than the difference between the two \( \pi \) and \( \pi^* \) sets. The reason for this is that the atomic orbital overlap and thus the strength of the interaction are greater for a \( \sigma \) bond than a \( \pi \) bond, which means that the \( \sigma \) molecular orbital is more stable (lower in energy) than the \( \pi \) molecular orbitals.

![figure](image)

Figure \( \PageIndex{4} \): The Relative Energies of the \( \sigma \) and \( \pi \) Molecular Orbitals Derived from \( np_x, np_y, \) and \( np_z \) Orbitals on Identical Adjacent Atoms. Because the two \( np_z \) orbitals point directly at each other, their orbital overlap is greater, so the difference in energy between the \( \sigma \) and \( \sigma^* \) molecular orbitals is greater than the energy difference between the \( \pi \) and \( \pi^* \) orbitals.

Although many combinations of atomic orbitals form molecular orbitals, we will discuss only one other interaction: an \( ns \) atomic orbital on one atom with an \( np_z \) atomic orbital on another. As shown in Figure \( \PageIndex{5} \), the sum of the two atomic wave functions \( (ns + np_z) \) produces a \( \sigma \) bonding molecular orbital. Their difference \( (ns - np_z) \) produces a \( \sigma^* \) antibonding molecular orbital, which has a nodal plane of zero probability density perpendicular to the internuclear axis.

![figure](image)

Figure \( \PageIndex{5} \): Formation of Molecular Orbitals from an \( ns \) Atomic Orbital on One Atom and an \( np_z \) Atomic Orbital on an Adjacent Atom. (a) The mathematical sum results in a \( \sigma \) (bonding) molecular orbital, with increased probability density between the nuclei. (b) The mathematical difference results in a \( \sigma^* \) (antibonding) molecular orbital, with a nodal plane of zero probability density perpendicular to the internuclear axis.

Second Row Diatomic Molecules

If we combine the splitting schemes for the 2s and 2p orbitals, we can predict bond order in all of the diatomic molecules and ions composed of elements in the first complete row of the periodic table. Remember that only the valence orbitals of the atoms need be considered; as we saw in the cases of lithium hydride and dilithium, the inner orbitals remain tightly
bound and retain their localized atomic character.

We now describe examples of systems involving period 2 homonuclear diatomic molecules, such as N$_2$, O$_2$, and F$_2$. When we draw a molecular orbital diagram for a molecule, there are four key points to remember:

1. The number of molecular orbitals produced is the same as the number of atomic orbitals used to create them (the law of conservation of orbitals).
2. As the overlap between two atomic orbitals increases, the difference in energy between the resulting bonding and antibonding molecular orbitals increases.
3. When two atomic orbitals combine to form a pair of molecular orbitals, the bonding molecular orbital is stabilized about as much as the antibonding molecular orbital is destabilized.
4. The interaction between atomic orbitals is greatest when they have the same energy.

The number of molecular orbitals is always equal to the total number of atomic orbitals we started with.

We illustrate how to use these points by constructing a molecular orbital energy-level diagram for F$_2$. We use the diagram in part (a) in Figure 6: Molecular Orbital Energy-Level Diagrams for Homonuclear Diatomic Molecules. For F$_2$, with 14 valence electrons (7 from each F atom), all of the energy levels except the highest, $\sigma^\star_{2p_z}$, are filled. This diagram shows 8 electrons in bonding orbitals and 6 in antibonding orbitals, resulting in a bond order of $8 - 6 = 2$. Thus F$_2$ is predicted to have a stable F–F single bond, in agreement with experimental data.

Figure 6: Molecular Orbital Energy-Level Diagrams for Homonuclear Diatomic Molecules. (a) For F$_2$, with 14 valence electrons (7 from each F atom), all of the energy levels except the highest, $\sigma^\star_{2p_z}$, are filled. This diagram shows 8 electrons in bonding orbitals and 6 in antibonding orbitals, resulting in a bond order of 1. (b) For O$_2$, with 12 valence electrons (6 from each O atom), there are only 2 electrons to place in the $\pi^\star$ orbital, giving a bond order of $(8 - 6) / 2 = 1$. Thus O$_2$ is predicted to have a stable O=O double bond, in agreement with experimental data.
\( \pi^\star \ \) pair of orbitals. Hund’s rule dictates that one electron occupies each orbital, and their spins are parallel, giving the \( \text{O}_2 \) molecule two unpaired electrons. This diagram shows 8 electrons in bonding orbitals and 4 in antibonding orbitals, resulting in a predicted bond order of 2.

We now turn to a molecular orbital description of the bonding in \( \text{O}_2 \). It so happens that the molecular orbital description of this molecule provided an explanation for a long-standing puzzle that could not be explained using other bonding models. To obtain the molecular orbital energy-level diagram for \( \text{O}_2 \), we need to place 12 valence electrons (6 from each O atom) in the energy-level diagram shown in part (b) in Figure \( \PageIndex{6} \). We again fill the orbitals according to Hund’s rule and the Pauli principle, beginning with the orbital that is lowest in energy. Two electrons each are needed to fill the \( \sigma_{2s} \) and \( \sigma_{2s}^\star \) orbitals, 2 more to fill the \( \sigma_{2p_z} \) orbital, and 4 to fill the degenerate \( \pi_{2p_x}^\star \) and \( \pi_{2p_y}^\star \) orbitals. According to Hund’s rule, the last 2 electrons must be placed in separate \( \pi^\star \) orbitals with their spins parallel, giving two unpaired electrons. This leads to a predicted bond order of \( (8 - 4) \div 2 = 2 \), which corresponds to a double bond, in agreement with experimental data (Table 4.5): the O–O bond length is 120.7 pm, and the bond energy is 498.4 kJ/mol at 298 K.

None of the other bonding models can predict the presence of two unpaired electrons in \( \text{O}_2 \). Chemists had long wondered why, unlike most other substances, liquid \( \text{O}_2 \) is attracted into a magnetic field. As shown in Video \( \PageIndex{2} \), it actually remains suspended between the poles of a magnet until the liquid boils away. The only way to explain this behavior was for \( \text{O}_2 \) to have unpaired electrons, making it paramagnetic, exactly as predicted by molecular orbital theory. This result was one of the earliest triumphs of molecular orbital theory over the other bonding approaches we have discussed.

Video \( \PageIndex{2} \): Liquid \( \text{O}_2 \) Suspended between the Poles of a Magnet. Because the \( \text{O}_2 \) molecule has two unpaired electrons, it is paramagnetic. Consequently, it is attracted into a magnetic field, which allows it to remain suspended between the poles of a powerful magnet until it evaporates. Full video can be found at...
The magnetic properties of O\(_2\) are not just a laboratory curiosity; they are absolutely crucial to the existence of life. Because Earth’s atmosphere contains 20% oxygen, all organic compounds, including those that compose our body tissues, should react rapidly with air to form H\(_2\)O, CO\(_2\), and N\(_2\) in an exothermic reaction. Fortunately for us, however, this reaction is very, very slow. The reason for the unexpected stability of organic compounds in an oxygen atmosphere is that virtually all organic compounds, as well as H\(_2\)O, CO\(_2\), and N\(_2\), have only paired electrons, whereas oxygen has two unpaired electrons. Thus the reaction of O\(_2\) with organic compounds to give H\(_2\)O, CO\(_2\), and N\(_2\) would require that at least one of the electrons on O\(_2\) change its spin during the reaction. This would require a large input of energy, an obstacle that chemists call a *spin barrier*. Consequently, reactions of this type are usually exceedingly slow. If they were not so slow, all organic substances, including this book and you, would disappear in a puff of smoke!

For period 2 diatomic molecules to the left of N\(_2\) in the periodic table, a slightly different molecular orbital energy-level diagram is needed because the \( \sigma \_\{2p\_\{z\}\} \) molecular orbital is slightly higher in energy than the degenerate \( \pi \_\{(\pi^\star\_{np\_x})\} \) and \( \pi \_\{(\pi^\star\_{np\_y})\} \) orbitals. The difference in energy between the 2s and 2p atomic orbitals increases from Li\(_2\) to F\(_2\) due to increasing nuclear charge and poor screening of the 2s electrons by electrons in the 2p subshell. The bonding interaction between the 2s orbital on one atom and the 2p\(_z\) orbital on the other is most important when the two orbitals have similar energies. This interaction decreases the energy of the \( \sigma \_\{2s\} \) orbital and increases the energy of the \( \sigma \_\{2p\_\{z\}\} \) orbital. Thus for Li\(_2\), Be\(_2\), B\(_2\), C\(_2\), and N\(_2\), the \( \sigma \_\{2p\_\{z\}\} \) orbital is higher in energy than the \( \sigma \_\{3p\_\{z\}\} \) orbitals, as shown in Figure \( \PageIndex{7} \). Experimentally, it is found that the energy gap between the \( ns \) and \( np \) atomic orbitals increases as the nuclear charge increases (Figure \( \PageIndex{7} \)). Thus for example, the \( \sigma \_\{2p\_\{z\}\} \) molecular orbital is at a lower energy than the \( \pi \_\{(\pi^\star\_{2p\_\{x\}})\} \) pair.

For Li\(_2\) through N\(_2\), the \( \sigma \_\{2p\_\{z\}\} \) orbital is higher in energy than the \( \pi \_\{(\pi^\star\_{2p\_\{x\}})\} \) orbitals. In contrast, the \( \pi \_\{(\pi^\star\_{2p\_\{y\}})\} \) orbital is lower in energy than the \( \pi \_\{(\pi^\star\_{2p\_\{x\}})\} \) orbitals for O\(_2\) and F\(_2\) due to the increase in the energy difference between the 2s and 2p atomic orbitals as the nuclear charge increases across the row.

Completing the diagram for N\(_2\) in the same manner as demonstrated previously, we find that the 10 valence electrons result in 8 bonding electrons and 2 antibonding electrons, for a predicted bond order of 3, a triple bond. Experimental data show that the N–N bond is significantly shorter than the F–F bond (109.8 pm in N\(_2\) versus 141.2 pm in F\(_2\)), and the bond...
energy is much greater for \( \text{N}_2 \) than for \( \text{F}_2 \) (945.3 kJ/mol versus 158.8 kJ/mol, respectively). Thus the \( \text{N}_2 \) bond is much shorter and stronger than the \( \text{F}_2 \) bond, consistent with what we would expect when comparing a triple bond with a single bond.

Example \( \PageIndex{1} \))

Use a qualitative molecular orbital energy-level diagram to predict the electron configuration, the bond order, and the number of unpaired electrons in \( \text{S}_2 \), a bright blue gas at high temperatures.

**Given:** chemical species

**Asked for:** molecular orbital energy-level diagram, bond order, and number of unpaired electrons

**Strategy:**

A. Write the valence electron configuration of sulfur and determine the type of molecular orbitals formed in \( \text{S}_2 \). Predict the relative energies of the molecular orbitals based on how close in energy the valence atomic orbitals are to one another.

B. Draw the molecular orbital energy-level diagram for this system and determine the total number of valence electrons in \( \text{S}_2 \).

C. Fill the molecular orbitals in order of increasing energy, being sure to obey the Pauli principle and Hund’s rule.

D. Calculate the bond order and describe the bonding.

**Solution:**

A Sulfur has a \([\text{Ne}]3s^23p^4\) valence electron configuration. To create a molecular orbital energy-level diagram similar to those in Figure \( \PageIndex{6} \) and Figure \( \PageIndex{7} \), we need to know how close in energy the 3s and 3p atomic orbitals are because their energy separation will determine whether the \( \pi_{3p_{x,y}} \) or the \( \sigma_{3p_z} \) molecular orbital is higher in energy. Because the \( ns–np \) energy gap increases as the nuclear charge increases (Figure \( \PageIndex{7} \)), the \( \sigma_{3p_z} \) molecular orbital will be lower in energy than the \( \pi_{3p_{x,y}} \) pair.

B The molecular orbital energy-level diagram is as follows:

![Molecular Orbital Diagram](image)

Each sulfur atom contributes 6 valence electrons, for a total of 12 valence electrons.
Ten valence electrons are used to fill the orbitals through $\pi_{3p_{x}}$ and $\pi_{3p_{y}}$, leaving 2 electrons to occupy the degenerate $\pi^*_{3p_{x}}$ and $\pi^*_{3p_{y}}$ pair. From Hund’s rule, the remaining 2 electrons must occupy these orbitals separately with their spins aligned. With the numbers of electrons written as superscripts, the electron configuration of $S_2$ is $\left( \sigma_{3s} \right)^2 \left( \sigma^*_{3s} \right)^2 \left( \sigma_{3p_{z}} \right)^2 \left( \pi_{3p_{x,y}} \right)^4 \left( \pi^*_{3p_{x,y}} \right)^2$ with 2 unpaired electrons. The bond order is $(8 - 4) ÷ 2 = 2$, so we predict an S=S double bond.

Exercise

Use a qualitative molecular orbital energy-level diagram to predict the electron configuration, the bond order, and the number of unpaired electrons in the peroxide ion ($O_2^{2-}$).

Answer

$\left( \sigma_{2s} \right)^2 \left( \sigma^*_{2s} \right)^2 \left( \sigma_{2p_{z}} \right)^2 \left( \pi_{2p_{x,y}} \right)^4 \left( \pi^*_{2p_{x,y}} \right)^4$ bond order of 1; no unpaired electrons

**Molecular Orbitals for Heteronuclear Diatomic Molecules**

Diatomic molecules with two different atoms are called heteronuclear diatomic molecules. When two nonidentical atoms interact to form a chemical bond, the interacting atomic orbitals do not have the same energy. If, for example, element B is more electronegative than element A ($\chi_B > \chi_A$), the net result is a “skewed” molecular orbital energy-level diagram, such as the one shown for a hypothetical A–B molecule in Figure. The atomic orbitals of element B are uniformly lower in energy than the corresponding atomic orbitals of element A because of the enhanced stability of the electrons in element B. The molecular orbitals are no longer symmetrical, and the energies of the bonding molecular orbitals are more similar to those of the atomic orbitals of B. Hence the electron density of bonding electrons is likely to be closer to the more electronegative atom. In this way, molecular orbital theory can describe a polar covalent bond.

![Figure](image.png): Molecular Orbital Energy-Level Diagram for a Heteronuclear Diatomic Molecule AB, Where $\chi_B > \chi_A$. The bonding molecular orbitals are closer in energy to the atomic orbitals of the more electronegative B atom. Consequently, the electrons in the bonding orbitals are not shared equally between the two atoms. On average, they are
closer to the B atom, resulting in a polar covalent bond.

A molecular orbital energy-level diagram is always skewed toward the more electronegative atom.

### An Odd Number of Valence Electrons: NO

Nitric oxide (NO) is an example of a heteronuclear diatomic molecule. The reaction of O$_2$ with N$_2$ at high temperatures in internal combustion engines forms nitric oxide, which undergoes a complex reaction with O$_2$ to produce NO$_2$, which in turn is responsible for the brown color we associate with air pollution. Recently, however, nitric oxide has also been recognized to be a vital biological messenger involved in regulating blood pressure and long-term memory in mammals.

Because NO has an odd number of valence electrons (5 from nitrogen and 6 from oxygen, for a total of 11), its bonding and properties cannot be successfully explained by either the Lewis electron-pair approach or valence bond theory. The molecular orbital energy-level diagram for NO (Figure \( \PageIndex{9} \)) shows that the general pattern is similar to that for the O$_2$ molecule (Figure \( \PageIndex{7} \)). Because 10 electrons are sufficient to fill all the bonding molecular orbitals derived from 2$p$ atomic orbitals, the 11th electron must occupy one of the degenerate $\pi^*$ orbitals. The predicted bond order for NO is therefore $\frac{(8-3)}{2} = 2\ 1/2$. Experimental data, showing an N–O bond length of 115 pm and N–O bond energy of 631 kJ/mol, are consistent with this description. These values lie between those of the N$_2$ and O$_2$ molecules, which have triple and double bonds, respectively. As we stated earlier, molecular orbital theory can therefore explain the bonding in molecules with an odd number of electrons, such as NO, whereas Lewis electron structures cannot.

![Molecular Orbital Energy-Level Diagram for NO](image)

Figure \( \PageIndex{9} \): Molecular Orbital Energy-Level Diagram for NO. Because NO has 11 valence electrons, it is paramagnetic, with a single electron occupying the $\{ \pi^\star \}_{2p_x}, \pi^\star_{2p_y}\$ pair of orbitals.

Molecular orbital theory can also tell us something about the chemistry of \( \text{NO} \). As indicated in the energy-level diagram in Figure \( \PageIndex{9} \), NO has a single electron in a relatively high-energy molecular orbital. We might therefore expect it to have similar reactivity as alkali metals such as Li and Na with their single valence electrons. In fact, \( \text{NO}^+ \) is easily oxidized to the \( \text{NO}^{2+} \) cation, which is isoelectronic with \( \text{N}_2^+ \) and has a bond order of 3, corresponding to an
Nonbonding Molecular Orbitals

Molecular orbital theory is also able to explain the presence of lone pairs of electrons. Consider, for example, the HCl molecule, whose Lewis electron structure has three lone pairs of electrons on the chlorine atom. Using the molecular orbital approach to describe the bonding in HCl, we can see from Figure (PageIndex{10}) that the 1s orbital of atomic hydrogen is closest in energy to the 3p orbitals of chlorine. Consequently, the filled Cl 3s atomic orbital is not involved in bonding to any appreciable extent, and the only important interactions are those between the H 1s and Cl 3p orbitals. Of the three p orbitals, only one, designated as 3pz, can interact with the H 1s orbital. The 3px and 3py atomic orbitals have no net overlap with the 1s orbital on hydrogen, so they are not involved in bonding. Because the energies of the Cl 3s, 3px, and 3py orbitals do not change when HCl forms, they are called nonbonding molecular orbitals. A nonbonding molecular orbital occupied by a pair of electrons is the molecular orbital equivalent of a lone pair of electrons. By definition, electrons in nonbonding orbitals have no effect on bond order, so they are not counted in the calculation of bond order. Thus the predicted bond order of HCl is \((2 - 0) ÷ 2 = 1\). Because the σ bonding molecular orbital is closer in energy to the Cl 3pz than to the H 1s atomic orbital, the electrons in the σ orbital are concentrated closer to the chlorine atom than to hydrogen. A molecular orbital approach to bonding can therefore be used to describe the polarization of the H–Cl bond to give \( H^\delta^+ - Cl^\delta^- \).

Electrons in nonbonding molecular orbitals have no effect on bond order.

Example (PageIndex{2})

Use a "skewed" molecular orbital energy-level diagram like the one in Figure (PageIndex{8}) to describe the bonding in the cyanide ion \((CN^-)\). What is the bond order?

**Given:** chemical species
**Asked for:** "skewed" molecular orbital energy-level diagram, bonding description, and bond order

**Strategy:**

A. Calculate the total number of valence electrons in CN\(^{-}\). Then place these electrons in a molecular orbital energy-level diagram like Figure \(\PageIndex{8}\) in order of increasing energy. Be sure to obey the Pauli principle and Hund’s rule while doing so.

B. Calculate the bond order and describe the bonding in CN\(^{-}\).

**Solution:**

A The CN\(^{-}\) ion has a total of 10 valence electrons: 4 from C, 5 from N, and 1 for the −1 charge. Placing these electrons in an energy-level diagram like Figure \(\PageIndex{8}\) fills the five lowest-energy orbitals, as shown here:

Because \(\chi_{N} > \chi_{C}\), the atomic orbitals of N (on the right) are lower in energy than those of C. B The resulting valence electron configuration gives a predicted bond order of \((8 - 2) / 2 = 3\), indicating that the CN\(^{-}\) ion has a triple bond, analogous to that in N\(_2\).

**Exercise \(\PageIndex{2}\)**

Use a qualitative molecular orbital energy-level diagram to describe the bonding in the hypochlorite ion (OCl\(^{-}\)). What is the bond order?

**Answer**

All molecular orbitals except the highest-energy \(\sigma^{*}\) are filled, giving a bond order of 1.

Although the molecular orbital approach reveals a great deal about the bonding in a given molecule, the procedure quickly becomes computationally intensive for molecules of even moderate complexity. Furthermore, because the computed molecular orbitals extend over the entire molecule, they are often difficult to represent in a way that is easy to visualize. Therefore we do not use a pure molecular orbital approach to describe the bonding in molecules or ions with more than two atoms. Instead, we use a valence bond approach and a molecular orbital approach to explain, among other things, the concept of resonance, which cannot adequately be explained using other methods.