

## 9.12: Molecular-Orbital Theory Also Applies to 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 9.12.4. 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.

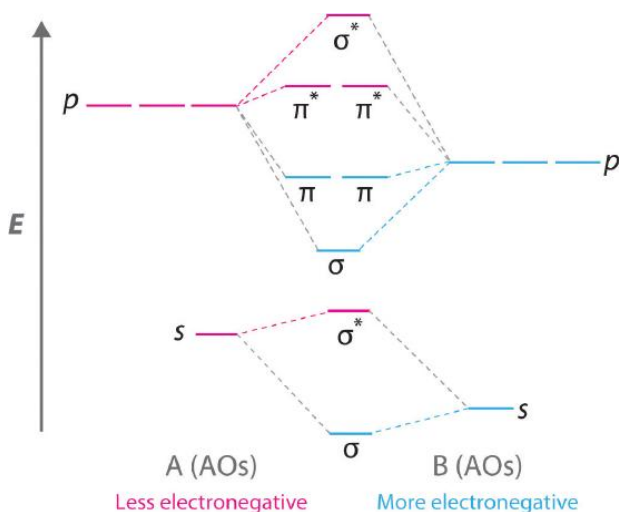


Figure 9.12.4: Molecular Orbital Energy-Level Diagram for a Heteronuclear Diatomic Molecule AB, Where  $\chi_B > \chi_A$ . The bonding molecular orbitals are closer in energy to the atomic orbitals of the more electronegative B atom. Consequently, the electrons in the bonding orbitals are not shared 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 9.12.13) shows that the general pattern is similar to that for the  $O_2$  molecule (Figure 9.12.11). Because 10 electrons are sufficient to fill all the bonding molecular orbitals derived from  $2p$  atomic orbitals, the 11th electron must occupy one of the degenerate  $\pi^*$  orbitals. The predicted bond order for NO is therefore  $(8-3) \div 2 = 2 \frac{1}{2}$ . 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.

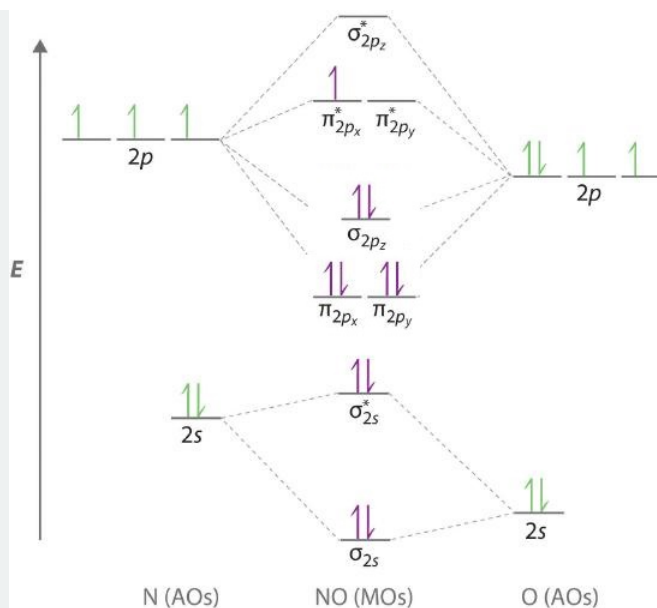


Figure 9.12.13: Molecular Orbital Energy-Level Diagram for NO. Because NO has 11 valence electrons, it is paramagnetic, with a single electron occupying the  $(\pi_{2p_x}^*, \pi_{2p_y}^*)$  pair of orbitals.

Note that electronic structure studies show the ground state configuration of NO to be  $(\sigma_{2s})^2(\sigma_{2s}^*)^2(\pi_{2p_{x,y}})^4(\sigma_{2p_z})^2(\pi_{2p_{x,y}}^*)^1$  in order of increasing energy. Hence, the  $\pi_{2p_{x,y}}$  orbitals are lower in energy than the  $\sigma_{2p_z}$  orbital. This is because the NO molecule is near the transition of flipping energies levels observed in homonuclear diatomics where the sigma bond drops below the pi bond (Figure 9.12.11).

Molecular orbital theory can also tell us something about the *chemistry* of NO. As indicated in the energy-level diagram in Figure 9.12.13 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, NO is easily oxidized to the  $NO^+$  cation, which is isoelectronic with  $N_2$  and has a bond order of 3, corresponding to an  $N \equiv O$  triple bond.

### 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 9.12.6 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  $3p_z$ , can interact with the H 1s orbital. The  $3p_x$  and  $3p_y$  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,  $3p_x$ , and  $3p_y$  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  $(2 - 0) \div 2 = 1$ . Because the  $\sigma$  bonding molecular orbital is closer in energy to the Cl  $3p_z$  than to the H 1s atomic orbital, the electrons in the  $\sigma$  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  $H^{\delta+} - Cl^{\delta-}$ .

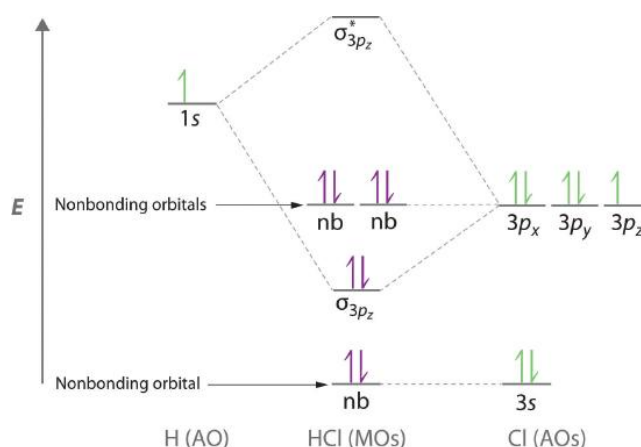


Figure 9.12.6: Molecular Orbital Energy-Level Diagram for HCl. The hydrogen 1s atomic orbital interacts most strongly with the  $3p_z$  orbital on chlorine, producing a bonding/antibonding pair of molecular orbitals. The other electrons on Cl are best viewed as nonbonding. As a result, only the bonding  $\sigma$  orbital is occupied by electrons, giving a bond order of 1.

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

#### ✓ Example 9.12.4: The Cyanide Ion

Use a “skewed” molecular orbital energy-level diagram like the one in Figure 9.12.4 to describe the bonding in the cyanide ion ( $\text{CN}^-$ ). What is the bond order?

**Given:** chemical species

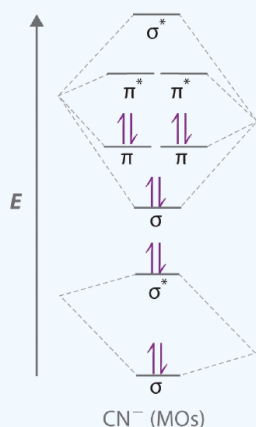
**Asked for:** “skewed” molecular orbital energy-level diagram, bonding description, and bond order

**Strategy:**

- Calculate the total number of valence electrons in  $\text{CN}^-$ . Then place these electrons in a molecular orbital energy-level diagram like Figure 9.12.4 in order of increasing energy. Be sure to obey the [Pauli principle](#) and [Hund's rule](#) while doing so.
- Calculate the bond order and describe the bonding in  $\text{CN}^-$ .

**Solution:**

**A** The  $\text{CN}^-$  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 9.12.4 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  $(8 - 2) \div 2 = 3$ , indicating that the  $\text{CN}^-$  ion has a triple bond, analogous to that in  $\text{N}_2$ .

### ? Exercise 9.12.4: The Hypochlorite Ion

Use a qualitative molecular orbital energy-level diagram to describe the bonding in the hypochlorite ion ( $\text{OCl}^-$ ). 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.



### Summary

Molecular orbital energy-level diagrams for diatomic molecules can be created if the electron configuration of the parent atoms is known, following a few simple rules. Most important, the number of molecular orbitals in a molecule is the same as the number of atomic orbitals that interact. The difference between bonding and antibonding molecular orbital combinations is proportional to the overlap of the parent orbitals and decreases as the energy difference between the parent atomic orbitals increases. With such an approach, the electronic structures of virtually all commonly encountered **homonuclear diatomic molecules**, molecules with two identical atoms, can be understood. The molecular orbital approach correctly predicts that the  $\text{O}_2$  molecule has two unpaired electrons and hence is attracted into a magnetic field. In contrast, most substances have only paired electrons. A similar procedure can be applied to molecules with two dissimilar atoms, called **heteronuclear diatomic molecules**, using a molecular orbital energy-level diagram that is skewed or tilted toward the more electronegative element. Molecular orbital theory is able to describe the bonding in a molecule with an odd number of electrons such as NO and even to predict something about its chemistry.

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