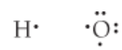
Pearson

4.7: Covalent Bonding: Simple Lewis Structures

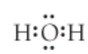
In the Lewis model, we represent covalent bonding with a *Lewis structure*, which depicts neighboring atoms as sharing some (or all) of their valence electrons in order to attain octets (or duets for hydrogen).

Single Covalent Bonds

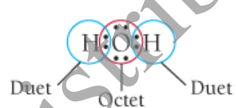
Consider hydrogen and oxygen, which have the following Lewis symbols:



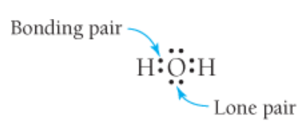
In water, these atoms share their unpaired valence electrons so that each hydrogen atom gets a duet and the oxygen atom gets an octet as represented with this Lewis structure:



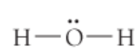
The shared electrons—those that appear in the space between the two atoms—count toward the octets (or duets) of *both of the atoms.*



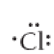
A shared pair of electrons is called a **bonding pair**, while a pair that is associated with only one atom—and therefore not involved in bonding—is called a **lone pair**. Lone pair electrons are also called **nonbonding electrons**.



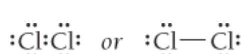
We often represent a bonding pair of electrons by a dash to emphasize that the pair constitutes a chemical bond.



The Lewis model helps explain the tendency of some elements to form diatomic molecules. For example, consider the Lewis symbol for chlorine:



If two Cl atoms pair together, they each get an octet.



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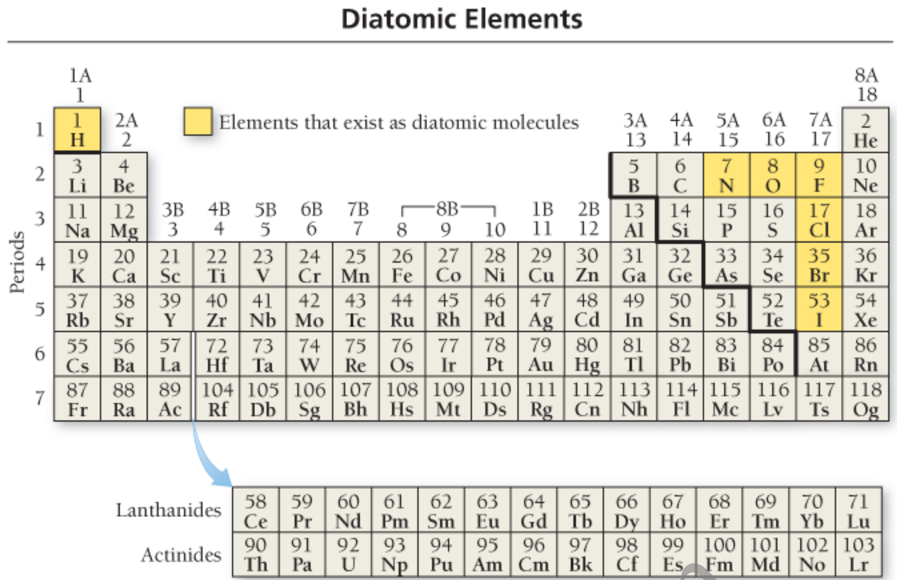
Elemental chlorine does indeed exist as a diatomic molecule in nature, just as the Lewis model predicts. The same is true for the other halogens and several other elements as shown in Figure 4.9.

Figure 4.9 Diatomic Elements

The highlighted elements exist primarily as diatomic molecules.



Notice from Figure 4.9 that hydrogen also exists as a diatomic element. Similar to chlorine, the Lewis symbol for hydrogen has one unpaired electron.

H·

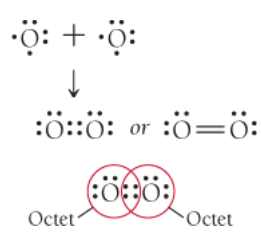
When two hydrogen atoms share their unpaired electron, each gets a duet, a stable configuration for hydrogen.

H:H *or* H–H

Again, the Lewis model helps us explain what we see in nature.

Double and Triple Covalent Bonds

In the Lewis model, two atoms may share more than one electron pair to get octets. For example, if we pair two oxygen atoms together, they share two electron pairs.



> One dash always stands for *two* electrons (a single bonding pair).

Each oxygen atom now has an octet because *the additional bonding pair counts toward the octet of both oxygen atoms.* When two atoms share two electron pairs, the resulting bond is a **double bond**. In general, double bonds are shorter and stronger than single bonds. The double bond that forms between two oxygen atoms explains why oxygen exists as a diatomic molecule.

> We will explore the characteristics of multiple bonds more fully in Section 6.3.

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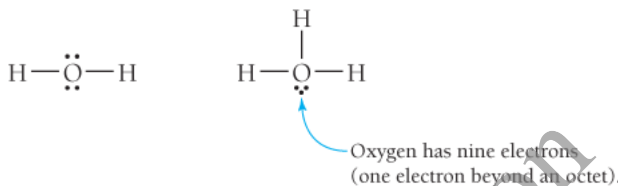
Atoms can also share three electron pairs. Consider the Lewis structure of another diatomic molecule, N_2. Because each N atom has five valence electrons, the Lewis structure for N_2 has 10 electrons. Both nitrogen atoms attain octets by sharing three electron pairs:

$$:N:::N: \quad or \quad :N\equiv N:$$

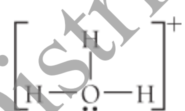
In this case, the bond is a **triple bond**. Triple bonds are even shorter and stronger than double bonds. When we examine nitrogen in nature, we find that it exists as a diatomic molecule with a very strong bond between the two nitrogen atoms. The bond is so strong that it is difficult to break, making N_2 a relatively unreactive molecule.

Covalent Bonding: Models and Reality

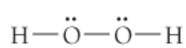
The Lewis model predicts the properties of molecular compounds in many ways. We have already seen how it explains the existence of several diatomic elements. The Lewis model also accounts for why particular combinations of atoms form molecules and others do not. For example, why is water H_2O and not H_3O? We can write a good Lewis structure for H_2O but not for H_3O.



In this way, the Lewis model predicts that H_2O should be stable, while H_3O should not be, and that is in fact the case. However, if we remove an electron from H_3O, we get H_3O^+, which should be stable (according to the Lewis model) because, when we remove the extra electron, oxygen gets an octet.



This ion, called the hydronium ion, is in fact stable in aqueous solutions (see Section 8.7). The Lewis model predicts other possible combinations for hydrogen and oxygen as well. For example, we can write a Lewis structure for H_2O_2 as follows:



Indeed, H_2O_2, or hydrogen peroxide, exists and is often used as a disinfectant and a bleach.

The Lewis model also accounts for why covalent bonds are highly *directional*. The attraction between two covalently bonded atoms is due to the sharing of one or more electron pairs in the space between them. Thus, each bond links just one specific pair of atoms—*in contrast to ionic bonds, which are nondirectional and hold together an entire array of ions.* As a result, the fundamental units of covalently bonded compounds are individual molecules. These molecules can interact with one another in a number of different ways that we will cover in Chapter 11. However, in covalently bonded molecular compounds, the interactions *between* molecules (intermolecular forces) are generally much weaker than the bonding interactions within a molecule (intramolecular forces), as shown in Figure 4.10. When a molecular compound melts or boils, the molecules themselves remain intact—only the relatively weak interactions between molecules must be overcome. Consequently, molecular compounds tend to have lower melting and boiling points than ionic compounds.

Figure 4.10 Intermolecular and Intramolecular Forces

The covalent bonds between atoms of a molecule are much stronger than the interactions between

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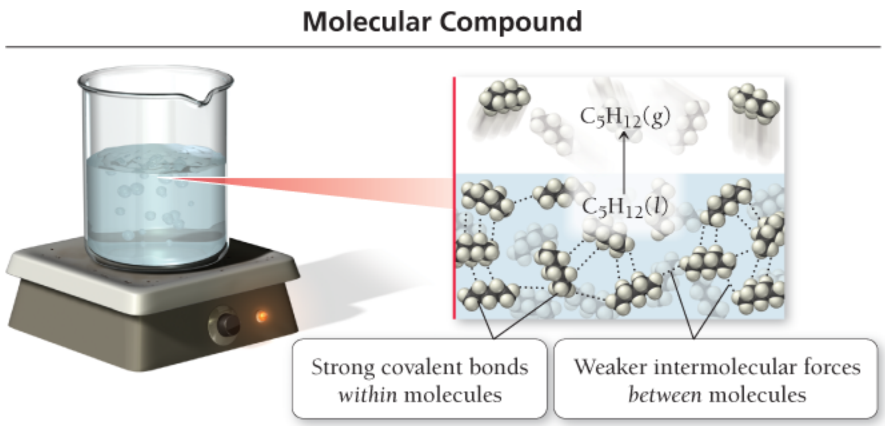
molecules. To bring a molecular substance to a boil, only the relatively weak intermolecular forces have to be overcome, so molecular compounds often have low boiling points.



Conceptual Connection 4.5 **Energy and the Octet Rule**

Conceptual Connection 4.6 **Ionic and Molecular Compounds**

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