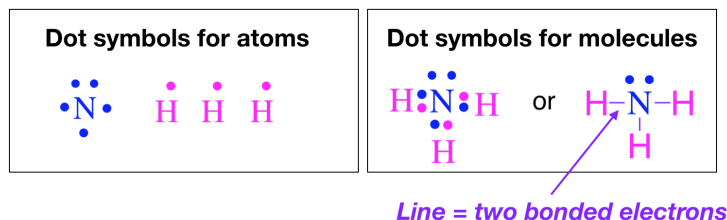


2.2.1: Lewis Electron-Dot Diagrams

In 1916, Gilbert Lewis Newton introduced a simple way to show the bonding between atoms in a molecule through Lewis electron dot diagrams. Creating Lewis diagrams is rather simple and requires only a few steps and some accounting of the valence electrons on each atom. Valence electrons are represented as dots. When two electrons are paired (lone pairs), they are represented by two adjacent dots located on an atom, and when two paired electrons are shared between atoms (bonds), they are shown as lines. For example, below are the electron dot structures of atoms and the Lewis electron dot structures of the molecules. These diagrams are helpful because they allow us to show how atoms are connected, and when coupled with [Valence Shell Electron Repulsion Theory \(VSEPR\)](#), we can use Lewis structures to predict the shape of the molecule.



The drawing of Lewis electron-dot structures is guided largely by the **octet rule**: that atoms form bonds to achieve eight electrons in their valence shell. For many elements, a full valence shell has an electron configuration of s^2p^6 , or eight electrons. A common exception to this rule is the first row elements, H and He. These two elements have $n = 1$ as their valence shell, and so they have only two electrons in a full valence shell ($1s^2$ electron configuration). Although H and He are exceptions to the "octet rule", they still form bonds to achieve a full valence shell. It may be better to think of this as the "full valence" rule of bonding. We will see many more violations to the "octet" rule as we progress through this course. In the case of metals and metalloids, breaking of the rules is particularly common. (CC-BY-NC-SA; Kathryn Haas)

Common Violations of the Octet Rule

- **Less than eight electrons (hypovalency):** H and He are examples of elements that cannot have more than two electrons in their full valence shell. Additionally, there are cases where a valid Lewis structure contains atoms with **hypovalency**: *partially-filled valence shells*. This is common in atoms with few valence electrons, such as beryllium and boron
- **More than eight electrons (hypervalency):** This is a case where an element has more than eight electrons in its valence shell. It is common for larger atoms ($n \geq 3$), and it is discussed further in [Section 3.3.3](#).

There are different rules for counting electrons depending on the purpose of the counting. These are the rules for counting for "octets". The rules for calculation of an atom's formal charge are **different** and are described in [Section 3.3.1](#). When an atom is part of a molecule, all electrons that are associated with the atom are counted as contributing to the atom's valence. This includes electrons that are lone pairs on the atoms, and *all* electrons that are shared in bonds. If four electrons are shared between two atoms, it is a double bond. If six electrons are shared between atoms, it is a triple bond.

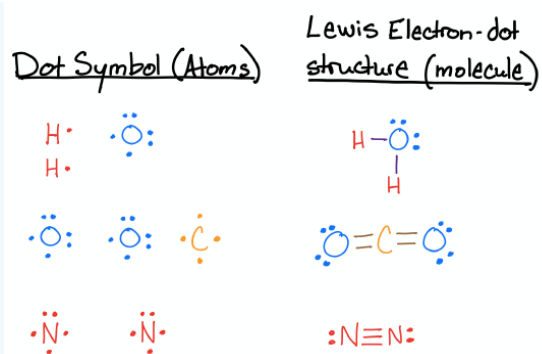
Electrons in valence (octet) = total unbonded electrons on the atom + total bonded electrons (2 electrons per bond)

Even if you're an old pro at drawing Lewis structures, it's a good idea to polish up. Please complete the practice exercises below. You should get out an actual piece of paper (I know...just do it. It's good for you.) and a writing tool and try to complete each problem before checking the answer.

Exercise 2.2.1.1

Draw the Lewis structures for H_2O , CO_2 , and N_2 .

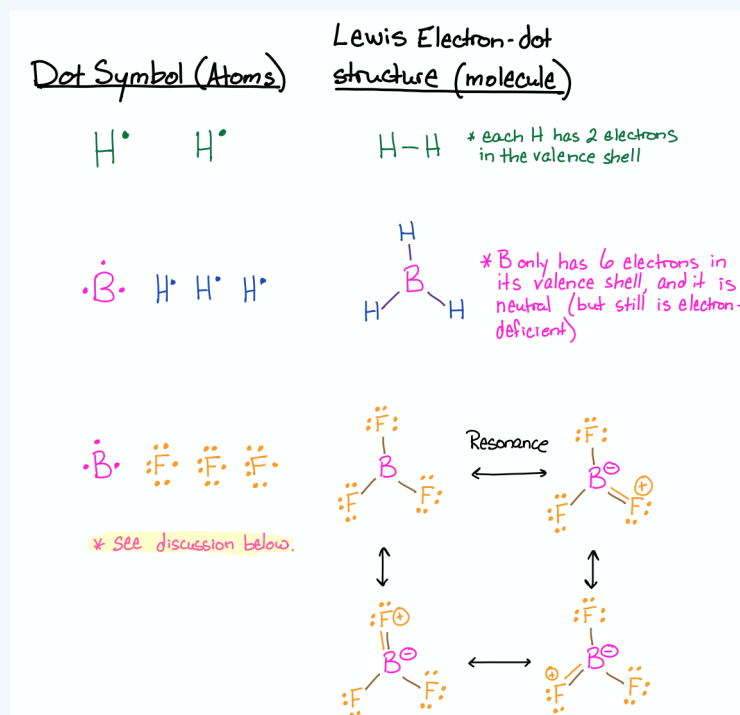
Answer



Exercise 2.2.1.2

Draw the Lewis structures for H_2 , BH_3 and BF_3 .

Answer



These three examples include atoms that have less than eight electrons in their valence shell. In the case of H, it is satisfied with only two electrons in its valence, as was discussed earlier in this section. The case of BH_3 was also discussed above. In the Lewis structure of BH_3 , the boron can only have six electrons in its octet and it is neutral in charge. The boron is electron deficient even though it has neutral charge.

The case of BF_3 deserves a discussion: If you are unfamiliar with resonance and formal charge, see Sections 3.3.1 (Formal Charge) and 3.3.2 (Resonance) first and come back to this afterward. You might have drawn the BF_3 structure similar to the one that is drawn for BH_3 , where boron has three bonds and only six electrons in its valence shell. If you did that, then you are correct; but if you only gave this structure, then your answer is not complete. There are three other correct ways to draw structure. Since F has lone pairs of electrons, other valid Lewis structures would each have a double bond to one of the fluorines (three total). All of these structures are called resonance structures, and based on the four of them, we could predict that BF_3 would have B-F bonds that have some double-bond character. In fact, this is the case for BF_3 ; it has bond lengths that are shorter than single bonds, but longer than double bonds. Read more about it on its [Wikipedia page here \(click\)](https://en.wikipedia.org/wiki/Boron_trifluoride).

Outside Resources

- [en.Wikipedia.org/wiki/Lewis_structures](https://en.wikipedia.org/wiki/Lewis_structures)

Contributors and Attributions

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