

## 1.7: COMMON BONDING PATTERNS FOR ORGANIC CHEMISTRY

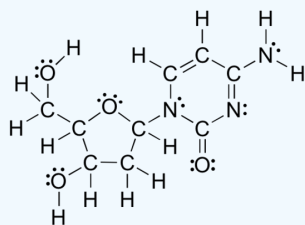
### Learning Objective

Draw, interpret, and convert between Lewis (Kekule), Condensed, and Bond-line Structures

Note: The review of general chemistry in sections 1.3 - 1.6 is integrated into the above Learning Objective for organic chemistry in sections 1.7 and 1.8.

### Common bonding patterns in organic structures

The Lewis structure below (of one of the four nucleoside building blocks that make up DNA) can appear complex and confusing at first glance. Fortunately, common bonding patterns occur that can allow for simplifications when drawing structures. The rules for the simplifies structures rely on the neutral bonding patterns for carbon, oxygen, nitrogen, phosphorus, and sulfur primarily. Since organic compounds have a hydrocarbon backbone, the atoms that are NOT carbon and hydrogen are called heteroatoms. 2'-deoxycytidine contains seven heteroatoms: four oxygen atoms and three nitrogen atoms. Heteroatoms are a primary source of chemical reactivity for organic chemistry.

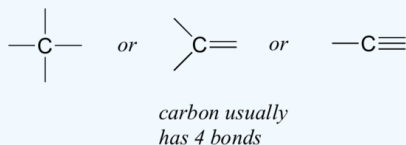


2'-deoxycytidine

**Heteroatoms:** atoms in an organic compound that are NOT carbon or hydrogen, typically oxygen, nitrogen, phosphorus, and sulfur

The ability to quickly and efficiently draw large structures and determine formal charges is not terribly hard to come by - all it takes is a few shortcuts and some practice at recognizing common bonding patterns.

Let's start with carbon, the most important element for organic chemists. Carbon is said to be tetravalent, meaning that it tends to form four bonds. If you look at the simple structures of methane, methanol, ethane, ethene, and ethyne in the figures from the previous section, you should quickly recognize that in each molecule, the carbon atom has four bonds, and a formal charge of zero.



This is a pattern that holds throughout most of the organic molecules we will see, but there are also exceptions.

In carbon dioxide, the carbon atom has double bonds to oxygen on both sides ( $O=C=O$ ). Later on in this chapter and throughout this book we will see examples of organic ions called 'carbocations' and 'carbanions', in which a carbon atom bears a positive or negative formal charge, respectively. If a carbon has only three bonds and an unfilled valence shell (in other words, if it does not fulfill the octet rule), it will have a positive formal charge.



3 bonds, no lone pair:  
carbocation



3 bonds + lone pair:  
carbanion



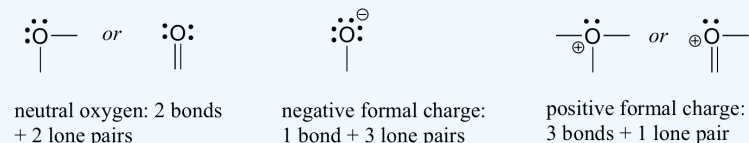
3 bonds + unpaired electron:  
carbon radical

If, on the other hand, it has three bonds plus a lone pair of electrons, it will have a formal charge of -1. Another possibility is a carbon with three bonds and a single, unpaired (free radical) electron: in this case, the carbon has a formal charge of zero. (One last possibility is a highly reactive species called a 'carbene', in which a carbon has two bonds and one lone pair of electrons, giving it a formal charge of zero. You may encounter carbenes in more advanced chemistry courses, but they will not be discussed any further in this book).

You should certainly use the methods you have learned to check that these formal charges are correct for the examples given above. More importantly, you will need, before you progress much further in your study of organic chemistry, to simply recognize these patterns (and the patterns described below for other atoms) and be able to identify carbons that bear positive and negative formal charges by a quick inspection.

The pattern for hydrogens is easy: hydrogen atoms have only one bond, and no formal charge. The exceptions to this rule are the proton,  $H^+$ , and the hydride ion,  $H^-$ , which is a proton plus two electrons. Because we are concentrating in this book on organic chemistry as applied to living things, however, we will not be seeing 'naked' protons and hydrides as such, because they are too reactive to be present in that form in aqueous solution. Nonetheless, the *idea* of a proton will be very important when we discuss acid-base chemistry, and the *idea* of a hydride ion will become very important much later in the book when we discuss organic oxidation and reduction reactions. As a rule, though, all hydrogen atoms in organic molecules have one bond, and no formal charge.

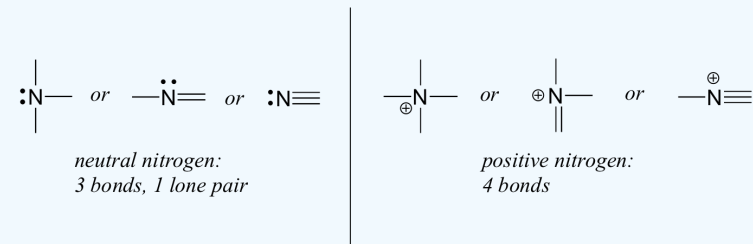
Let us next turn to oxygen atoms. Typically, you will see an oxygen bonding in three ways, all of which fulfill the octet rule.



If it has two bonds and two lone pairs, as in water, it will have a formal charge of zero. If it has one bond and three lone pairs, as in hydroxide ion, it will have a formal charge of -1. If it has three bonds and one lone pair, as in hydronium ion, it will have a formal charge of +1.

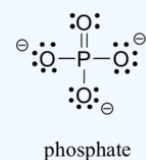
When we get to our discussion of free radical chemistry in chapter 17, we will see other possibilities, such as where an oxygen atom has one bond, one lone pair, and one unpaired (free radical) electron, giving it a formal charge of zero. For now, however, concentrate on the three main non-radical examples, as these will account for virtually everything we see until chapter 17.

Nitrogen has two major bonding patterns, both of which fulfill the octet rule:

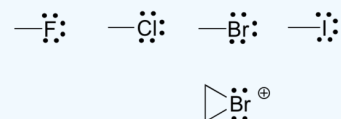


If a nitrogen has three bonds and a lone pair, it has a formal charge of zero. If it has four bonds (and no lone pair), it has a formal charge of +1. In a fairly uncommon bonding pattern, negatively charged nitrogen has two bonds and two lone pairs.

Two third row elements are commonly found in biological organic molecules: sulfur and phosphorus. Although both of these elements have other bonding patterns that are relevant in laboratory chemistry, in a biological context sulfur almost always follows the same bonding/formal charge pattern as oxygen, while phosphorus is present in the form of phosphate ion ( $\text{PO}_4^{3-}$ ), where it has five bonds (almost always to oxygen), no lone pairs, and a formal charge of zero. Remember that atoms of elements in the third row and below in the periodic table have 'expanded valence shells' with *d* orbitals available for bonding, and the octet rule does not apply.



Finally, the halogens (fluorine, chlorine, bromine, and iodine) are very important in laboratory and medicinal organic chemistry, but less common in naturally occurring organic molecules. Halogens in organic compounds usually are seen with one bond, three lone pairs, and a formal charge of zero. Sometimes, especially in the case of bromine, we will encounter reactive species in which the halogen has two bonds (usually in a three-membered ring), two lone pairs, and a formal charge of +1.



These rules, if learned and internalized so that you don't even need to think about them, will allow you to draw large organic structures, complete with formal charges, quite quickly.

Once you have gotten the hang of drawing Lewis structures, it is not always necessary to draw lone pairs on heteroatoms, as you can assume that the proper number of electrons are present around each atom to match the indicated formal charge (or lack thereof). Occasionally, though, lone pairs are drawn if doing so helps to make an explanation more clear.

**Exercise 1:** Draw one structure that corresponds to each of the following molecular formulas, using the common bonding patterns covered above. Be sure to include all lone pairs and formal charges where applicable, and assume that all atoms have a full valence shell of electrons. More than one correct answer is possible for each, so you will want to check your answers with your instructor or tutor.

a)  $\text{C}_5\text{H}_{10}\text{O}$  b)  $\text{C}_5\text{H}_8\text{O}$  c)  $\text{C}_6\text{H}_8\text{NO}^+$  d)  $\text{C}_4\text{H}_3\text{O}_2^-$

Solutions to exercises

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