

Drawing simple Lewis structures

One of the most important tools that chemists have in understanding what's going on in a chemical reaction is the Lewis structure. Though you won't see a ton of this in your general chemistry class, even this course will require Lewis structures for determining the polarity of a solvent, which is really, *really* important. Simply put, you really can't get by in chemistry without understanding Lewis structures.

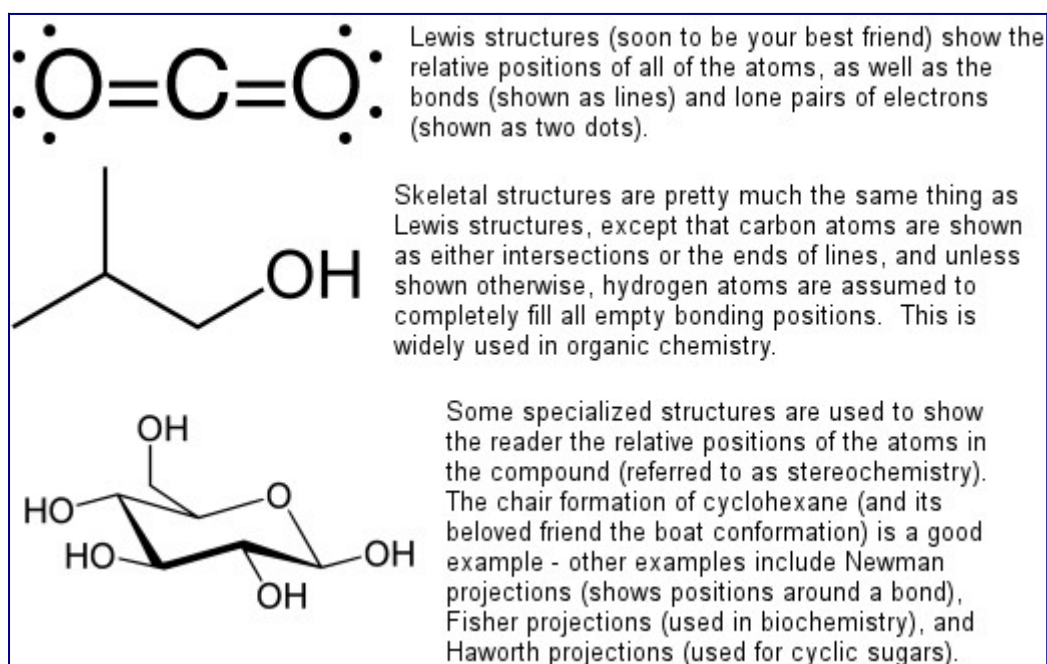
So, let's learn about them!

Types of formulas

Before we talk about Lewis structures, we really should talk about the various types of representing chemical compounds. After all, if we can't tell people the important things about a compound, there's not really any way that they'll be able to figure out what the heck we're talking about.

There are three big classes of chemical formula:

- **Molecular formulas** are the formulas that you're used to seeing. For example, even knowing nothing about ethylene, you can still undoubtedly tell somebody that the formula C_2H_4 means that a molecule of ethylene contains two atoms of carbon and four atoms of hydrogen. If you want to put these in fancy chemical terms, a molecular formula is "a list of how many atoms of each element are present in a compound."
- **Empirical formulas** are reduced molecular formulas. For example, if ethylene has a formula of C_2H_4 , the empirical formula of ethylene is a reduced version where both subscripts are divided by two: CH_2 . The official chemistry definition of an empirical formula is that it "shows the ratios of the numbers of atoms of each element in a compound."¹
- **Structural formulas** are pictures that show where all of the atoms in a compound are. Because different types of chemists need to know different types of stuff, there are lots of different types of structural formula:



Lewis structures: Your new chemical buddies!

Now that I've gotten all of that out of the way, it's time to talk about how to draw Lewis structures. As mentioned earlier, Lewis structures are incredibly important in this class, and way, way more important for actual chemists. If you ever wonder whether something is just being taught to keep you busy, I can assure you that this is not the case here.

To make life easy (?) for you, I've got a series of steps for drawing most Lewis structures. I won't lie to you and tell you that the steps are fun, but I *can* tell you that if you follow these steps, you'll be a Lewis structures pro.

Our example: Methane (CH₄)

Step 1: Find the number of valence electrons in the molecule

For those of you who forgot what that means, that's just the number of outer electrons that the molecule contains in the s- and p-orbitals. If you still don't know what that means, use the handy chart below:

1 2 3 4 5 6 7 8

H

C

These are d-block elements. For the purposes of figuring out the number of valence electrons, they might as well not exist at all. So, in your mind, pretend that they've just wandered off for a refreshing Fanta.

This is the f-block, so we're going to ignore it here.

The number at the top of each group that we're not busy ignoring refers to the number of valence electrons that elements in that group have. (The only exception is helium, which has 2, but since it doesn't bond, we don't care).

Because carbon is in the fourth row (i.e. the one with the four over it), it has four valence electrons. Because hydrogen is in the first row, it has one valence electron. Let's add 'em up:

- Total carbon valence electrons: 4 valence electrons x 1 atom C = 4 valence electrons
- Total hydrogen valence electrons: 1 valence electron x 4 atoms H = 4 valence electrons

Adding them up, the big total = $4 + 4 = 8$ **valence electrons**.

Step 2: Find the number of octet electrons in the molecule

Whereas the valence electrons refers to the number of electrons that the atoms in the molecule have, the octet electrons we're finding here refers to the number that each wants in order to obey the octet rule. Let's see what they are:

- hydrogen – wants 2 octet electrons
- beryllium – wants 4 octet electrons
- boron – wants 6 octet electrons
- everybody else wants 8 octet electrons

Let's figure it out:

- Total carbon octet electrons: 8 octet electrons x 1 atom = 8 octet electrons
- Total hydrogen octet electrons: 2 octet electrons x 4 atoms = 8 octet electrons

Adding them up, the atoms in the molecule want a total of **16 octet electrons**.

Step 3: Find out how many electrons the atoms need (bonding electrons)

Basically, subtract the valence electrons from the octet electrons. Or if you can't remember the name of each type of electron, just subtract the little one from the big one. Either way, in our example:

$$16 \text{ octet electrons} - 8 \text{ valence electrons} = \mathbf{8 \text{ bonding electrons}}$$

These are the electrons that will be making chemical bonds in this compound.

Step 4: Figure out how many bonds will be formed

If you remember that there are two electrons in a covalent bond (and I hope you do), then it stands to reason that if you divide the number of bonding electrons by 2, you'll get the number of bonds in the compound. Which is what we'll do:

$$8 \text{ bonding electrons} / 2 \text{ electrons per bond} = \mathbf{4 \text{ bonds}}$$

Step 5: Draw this thing

It's *finally* time for you to draw the molecule at hand. However, I'm guessing that you have no particular idea how to do this, and that you're wishing you'd taken earth science instead of chemistry right about now. In order to help you out, here's a handy hint:

Write the letter of the least abundant element on your paper

In this case, there's one carbon and four hydrogen atoms, so draw a big C. If you guessed the wrong element and put an H instead, it really doesn't make any difference. I just figured that putting a letter on the paper would make you feel like you're making progress. Which you are.

To figure out what bonds go where, here are some more handy rules:

- carbon's family bonds four times
- nitrogen's family and boron bond three times

- oxygen's family and beryllium bond twice
- hydrogen and the halogens bond once

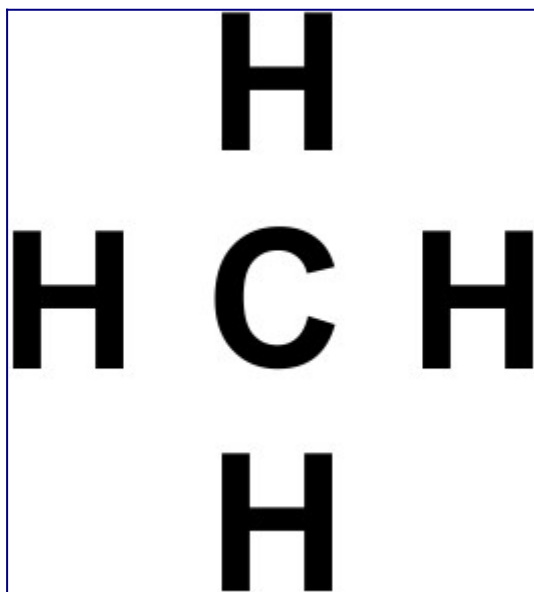
Don't break these rules, OK?

OK... let's go back and have a look at the C on your paper:

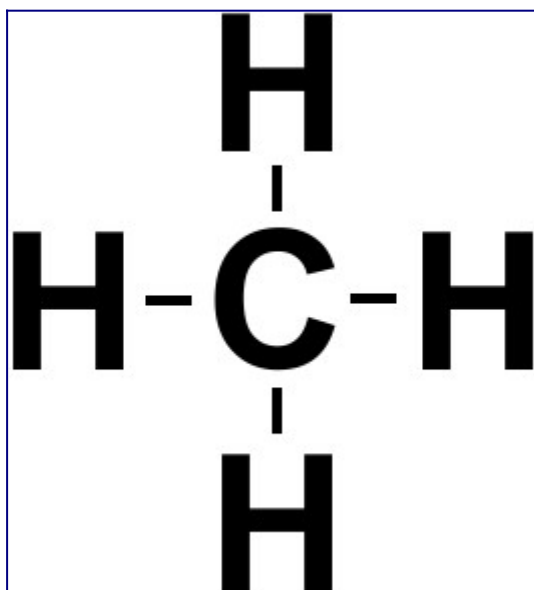
C

That's a C, all right

Let's start drawing some other letters around it. Why not put one on each side?



Now, ask yourself an important question: Do these atoms just hang around in space for no particular reason? Of course not. They need bonds. So let's just draw bonds between the carbon and each hydrogen. It's not like you've got anything better to do:



Let's check out if what we have is OK so far:

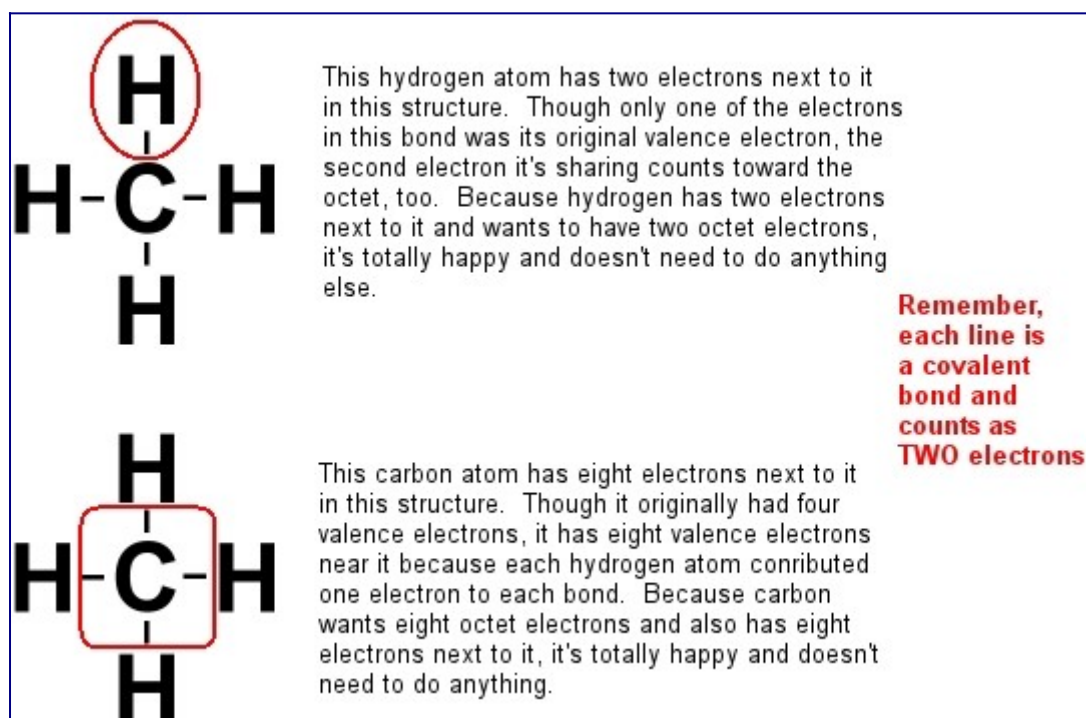
- Is the number of bonds right? Yep, we wanted four and we drew four lines.
- Does hydrogen bond the right number of times? Yep, it wants to bond once each and it does.
- Does carbon bond the right number of times? Yep, it wants to bond four times and it does.

Looks OK so far!

Step 6: Adding lone pairs

The last step in drawing this Lewis structure is to figure out if we need to add any lone pairs to any of the atoms present. In case you've forgotten (or I forgot to tell you), a lone pair of electrons consists of two electrons that are connected solely to one of the atoms in a molecule. They are drawn as two dots, which isn't really surprising.

To do this, what we first have to do is to figure out how many electrons each of the atoms in the structure we drew has around it. This is different than the number of valence electrons, because both of the electrons in each covalent bond count toward filling an atom's octet electrons. Heck, instead of talking about it, let's just have a look:



Because all of the electrons are accounted for, and because all of the atoms have the right number of electrons to satisfy the octet rule, we've come up with the right answer!

Another example: Ammonia, NH₃

Rather than do a lot of blathering on, let's just go through the steps:

1. Valence electrons:

Nitrogen: 5 valence electrons x 1 atom = 5 valence electrons

Hydrogen: 1 valence electron x 3 atoms = 3 valence electrons

Total number of valence electrons = 5 + 3 = 8

2. Octet electrons:

Nitrogen: 8 octet electrons x 1 atom = 8 octet electrons

Hydrogen: 2 octet electrons x 3 atoms = 6 octet electrons

Total number of octet electrons = 8 + 6 = 14.

3. Bonding electrons:

Octet electrons – valence electrons = bonding electrons

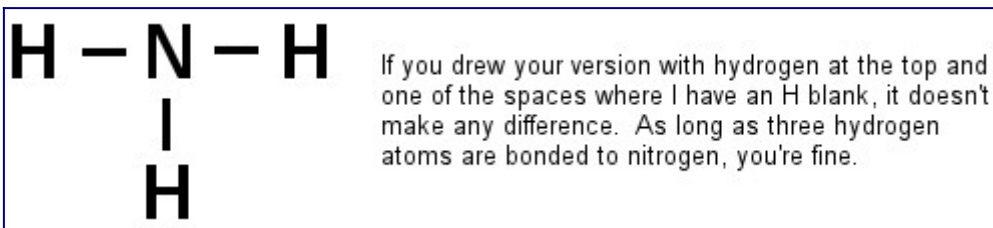
14 electrons – 8 electrons = 6 bonding electrons

4. Number of bonds:

6 bonding electrons / 2 = 3 bonds

5. Draw it:

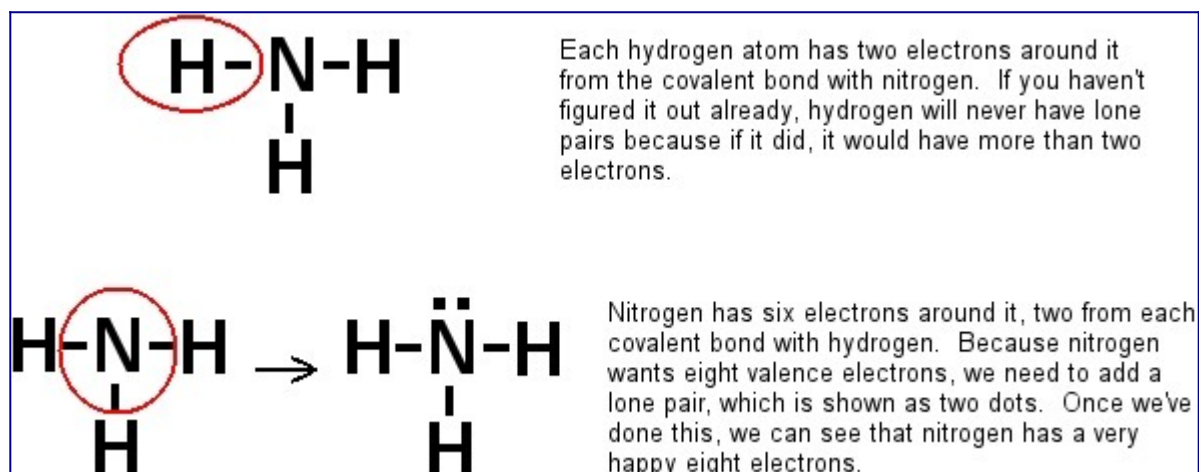
There are fewer nitrogen atoms than other atoms, so just write an “N” on your paper somewhere. When you're done with that, draw three “H” atoms around it and connect them to the N with bonds. You should have something that looks like this:



Given that hydrogen wants one bond each and that nitrogen wants three bonds, we're doing OK so far.

6. Adding lone pairs:

Let's take a look at each atom in this molecule and see if any of them need lone pairs:



And just like that, we're done!

Yet another example: carbon dioxide, CO₂

I'm not going to do the math for you, but will show the answers to each step so you can follow along:

1. 16 valence electrons
2. 24 octet electrons
3. 8 bonding electrons
4. 4 bonds
5. To draw this thing, let's put carbon in the middle and put a couple of oxygen atoms next to it:



Piece of cake! Except... uh... we have a problem.

The math we did in steps 1-4 tell us that we need to have four bonds in this molecule, whereas we only have two shown. Even worse, oxygen is supposed to have two bonds and carbon is supposed to have four bonds. We're clearly not doing something right.

So, how do we fix this? We can do one of two things:

- We can change the ordering of the atoms so they're connected in different ways.
- We can turn some of these single bonds into double or triple bonds.

It doesn't really matter which of these you try first, because you'll find that sometimes one will work and sometimes the other will. However, I've noticed that some students, when given a problem like this, will decide that they should just give up and/or freak out. After all, if you don't know what to do, why should you do it anyway?

Here's why: Because that's how the world works. If you think that this is the first and only time you'll ever run into something you don't intuitively know the answer to, you're a knucklehead.

Instead of feeling like you're lost, try something to see if it works. Try *anything* to see if it works. The worst thing that can happen is that you'll find out you tried the wrong thing and you'll try something else. If this upsets you, buy an eraser and quit whining!

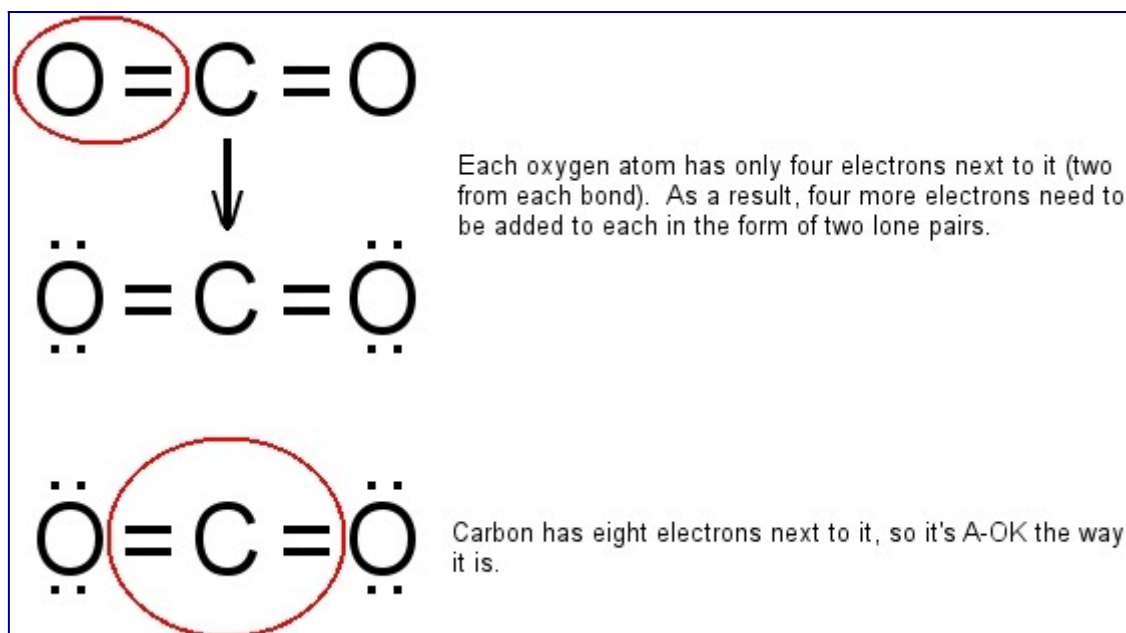
That said, I'm going to suggest that you put a couple more bonds in the structure you have. This suggestion is based on the fact that I know the answer to the problem, but it's also based on the fact that if you rearrange everything, carbon won't be in the middle anymore and since there's only one of them, that doesn't seem likely. Anyway, let's double up the bonds and see what happens:



Well, it appears that everything worked out OK. We used up the four bonds we figured out we'd need, each oxygen is bonded twice, and carbon has bonded four times. So far, so good.

6. Draw lone pairs

Let's see who (if anybody) needs lone pairs:



And there's our answer: Double bonds, lone pairs, and all.

Don't worry: I'll give you some more examples after I finish this last section:

Some words of caution:

During the process of learning how to write Lewis structures, some of you will run into odd problems that you can't figure out. Let's take a look at how to handle them:

- **You forget the steps I taught you:** No worries – if you keep practicing, you'll eventually learn it whether you want to or not.
- **The number of bonds you calculate in step 4 is a fraction:** No worries – do the math over again.
- **The number of bonds is simply too much or too little for the molecules:** No worries – do the math over again.

- **You did the math over again and it still doesn't work:** Either you're still doing the math wrong (in which case you should come back to it in a few minutes), it's a polyatomic ion (in which case you should read the next tutorial), or it's a compound with an extended octet (in which case you should check out the corresponding tutorial). No matter which it is, don't panic.
- **You get the right answers without doing the method in this tutorial:** That's not a problem. While this method works, it's not the only way to figure out the Lewis structure of a compound. As long as you get the right answer, it doesn't really matter how you do it!

No matter, what, don't panic. If you practice, practice practice, you can't help but get it right!

More practice problems:

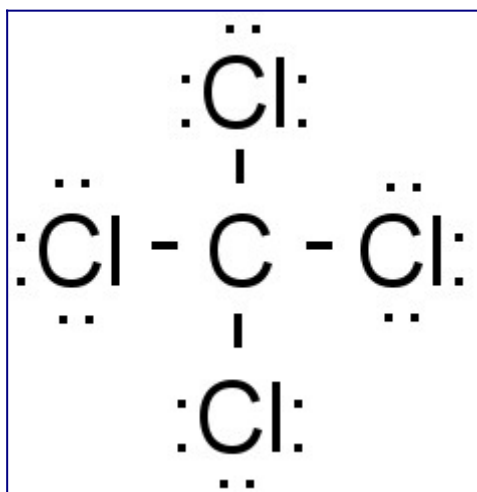
Draw the Lewis structures of the following covalent compounds. Solutions will be given below:

1. carbon tetrachloride
2. nitrogen (N_2)
3. boron trifluoride
4. formaldehyde (H_2CO)

Answers to the practice problems:

Carbon tetrachloride:

If you did this right, you should have calculated that there are a total of four bonds. Drawing this out and adding three sets of lone pairs for each chlorine atom, you should get a structure of:



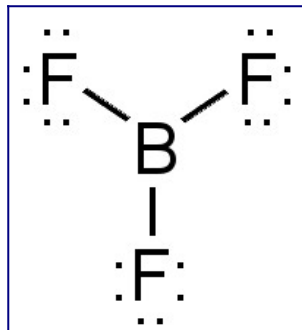
Nitrogen:

With three bonds and only two atoms, there has to be a nitrogen-nitrogen triple bond. This gives each nitrogen six electrons, so the extra two come from a lone pair on each:



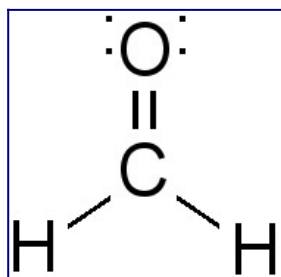
Boron trifluoride:

With three bonds, we find that boron bonds once to each fluorine atom – no surprises there. However, boron only needs six octet electrons and not eight, so it doesn't need any lone pair. (Note: If you drew this molecule with the fluorines at right angles, that's just fine – the reason I did it this way is because of the magic of VSEPR, which is in a different tutorial).



Formaldehyde:

There are four bonds and only three atoms stuck to carbon. How do we handle this? Well, the only way we can handle it is to bond oxygen twice to carbon, as there's no other possible way of handling this. Wrapping up, oxygen needs two lone pairs of electrons to finish the molecule:



Footnotes:

1. Unfortunately for you, empirical formulas have absolutely no value in the modern chemical world. You will undoubtedly have to learn that if you have an empirical formula, you can find the molecular formula by dividing the molar mass by the mass of the empirical formula and multiplying the subscripts accordingly. People haven't done this since the days of combustion analysis, which isn't done anymore. There is one case where you do see empirical formulas, and that's in ionic compounds, where the formulas are usually empirical. However, since ionic compounds don't form molecules, the empirical formula and the molecular formula (such as it is) are identical. Which again makes empirical formulas worthless.

Figure credits:

- All of the examples of structural formulas in the first drawing are public domain images, via Wikimedia Commons.
- I made the rest of them. Because I'm awesome.