

Polarity

You've undoubtedly noticed that your teacher has been talking a lot about polarity and intermolecular forces lately. He or she has probably even made it sound as if they are important in understanding the properties of chemical compounds and how they may react. Your teacher happens to be right: They *are* really important.

You thought I was going to make some smartass comment, didn't you? Nope. Polarity and intermolecular forces are actually really important. No smartass comment here.



You'll just have to settle for a dumbass picture instead.

What is polarity?

I'm glad you asked! Polarity occurs when there's an unequal sharing of electrons. *Bond polarity* occurs when one atom in a bond is more electronegative than another, while *molecular polarity* occurs in asymmetric molecules where the atoms on one part of the molecule are more electronegative than the others.

Polarity only occurs in covalent compounds. After all, if you look at ionic compounds, there are no molecules and no real bonds. What you *do* see in ionic compounds is a bunch of cations and anions crammed together, but that's not really what we're talking about when we're discussing polarity.

How can you tell if a bond is polar?

To tell if a bond is polar, look at the two atoms in the bond. If they are exactly the same atom, then the bond is nonpolar and there's an equal sharing of electrons. If they're different from each other, the bond is polar to some degree.

The problem with this definition is that polarity is meant to describe a significantly unequal sharing of electrons, which leads to unusual molecular behavior. Though carbon and hydrogen are different elements, their electronegativities are only very slightly different (2.5 and 2.1, respectively) so they pull on their shared electrons with almost exactly the same strength.¹ The molecule is technically polar, but for the purposes of classification, we'd say it's a nonpolar bond.

To figure out if a bond is classified as polar or nonpolar, you need to know the electronegativities of both atoms. The difference between the two electronegativities will allow you to classify the bond according to the following table:²

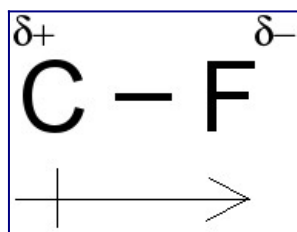
Difference in polarity between atoms	Bond classification	Examples
0 - 0.5	nonpolar	N-N, C-H, P-S
0.5-1.7	polar covalent	H-O, S-O, P-F
> 1.7	ionic	K-F, Li-O, Ba-Cl

Notes:

- Remember, these differences are subjective and may be cited differently.
- Ionic "bonds" are actually electrostatic interactions and not real bonds.
- All polar covalent and ionic bonds share electrons to some greater or lesser degree.

Though the electronegativity difference is used to find out whether a bond is strictly nonpolar, polar-covalent, or an ionic interaction, in practice we usually just assume that if a metal bonds to a nonmetal, it's an ionic compound, and most nonmetal-nonmetal bonds are polar covalent (unless they're the same element). Metalloids are usually classified as nonmetals when doing this.

To show whether something is polar in a diagram, we use the following notation:



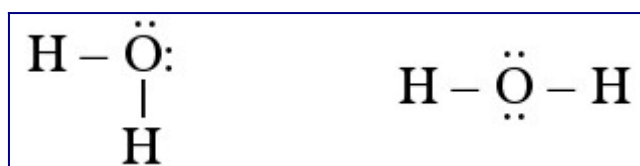
In this diagram, you'll notice two odd things:

- The $\delta+$ and $\delta-$ terms refer to where the electrons are pulled more strongly in the bond. $\delta+$ means "partial positive charge" because the fluorine (which has the $\delta-$ term) is pulling electrons away from it. This causes oxygen to be a little more positive than usual and the fluorine to be a little more negative.
- The arrow at the bottom is called a *dipole arrow* and points toward the more electronegative atom, with the line on the back indicating that this atom has partial positive charge. Generally speaking, charts will show either the δ terms or the dipole arrow in a diagram, but not both.

How can you tell if a molecule is polar?

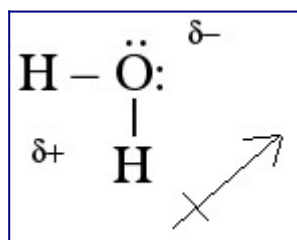
This is an easy one: Draw the molecule and check to see if all of the things connected to the central atom are the same. If they are, then the molecule is nonpolar (i.e. it has an equal sharing of electrons). If not, then it's polar (unequal sharing).

Let's draw a couple of diagrams:



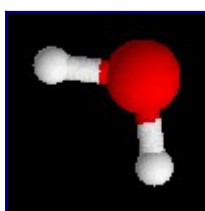
Water, water everywhere.

In the left diagram, we can see that the central atom, oxygen, has two different types of things stuck to it: hydrogen atoms and lone pairs. Because a hydrogen atom is different than a lone pair, this molecule is polar. Because the oxygen atom is more electronegative than the hydrogen atoms, the dipole arrow points toward the oxygen atom and the little δ signs are located on the side of the molecule with partial positive or negative charge.



If you look at the diagram of water to the right, you'll see something similar, with oxygen having two lone pairs and two hydrogen atoms – this indicates a polar molecule. Unfortunately, it's impossible to draw the dipole arrow for this molecule because the atoms are drawn symmetrically. What do we do?

Fortunately, we're saved by one fact: Water isn't a 2-D object. Though each of the diagrams shown portray it as being 2-dimensional, it's actually a 3-D structure as shown in the cool picture below:



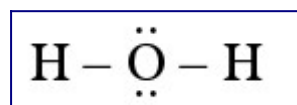
Because it's a 3-D structure, there are different valid ways of drawing it depending on the observer's viewpoint. Let me show you with a cool picture I took:



Pictured: My snack

The cherries on the left show the orientation of water that's shown on the left above. The cherries on the right are arranged in an identical way (getting them to balance like that was a pain in the butt), showing that from another angle they appear linear. They're the same molecule but seem to have different shapes depending on how you look at it.

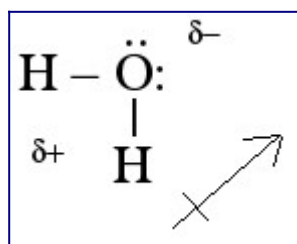
Going back to the diagram of water that looks like this:



we now understand that we drew the lone pairs and hydrogen atom in completely arbitrary positions. In order to properly show the polarity of the molecule, follow this rule:

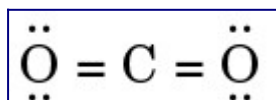
If you have a polar molecule, you can switch around any combination of atoms and lone pairs if you want to. The correct arrangement for showing polarity is the one where the molecule *looks* uneven.

So, for this water molecule, if we switch a lone pair with a hydrogen, we end up with exactly the same thing as we got for the other:



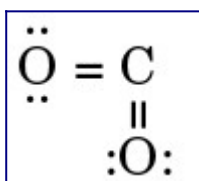
Another example:

Let's look at carbon dioxide:



In this structure, we can see that carbon has two singly-bonded oxygen atoms stuck to it. Because these atoms are the same, the molecule is nonpolar and there's no need to draw any of those δ things or an arrow.

However, what if we actually drew carbon dioxide like this:

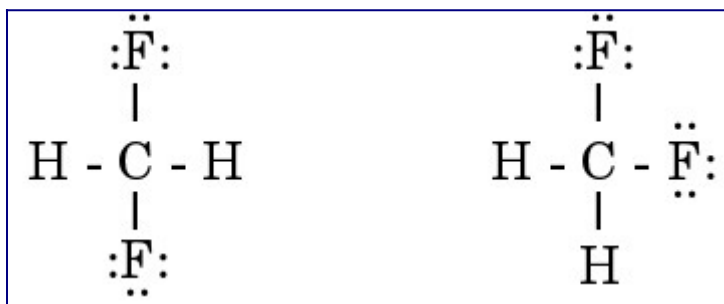


This Lewis structure is completely valid and follows all of the rules. It also makes it look like the oxygen side of the molecule should have a partially negative charge and the carbon side should have a partial positive charge. Which, according to our method for finding polarity, we know it doesn't.

Rest easy! This structure may be technically right, but it doesn't follow VSEPR, which says that all of the things stuck to an atom want to spread themselves out as far as possible. In this case, we can clearly see that the O=C=O bond isn't shown as spread out as it is in the first diagram. This makes the second portrayal technically correct, but misleading. And still nonpolar.

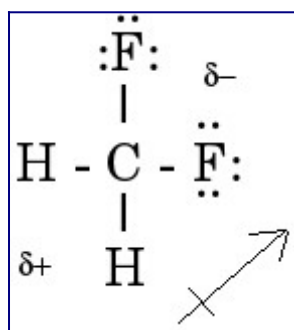
Another example: Difluoromethane (CH₂F₂):

Difluoromethane looks like this:



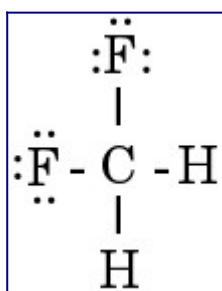
As with water, both of these diagrams are identical – just shown from different angles. If we want to figure out whether the molecule is polar or not, just look at the central carbon – you’ll see that it has two fluorine atoms and two hydrogen atoms stuck to it. Given that fluorine and hydrogen are different from each other, this indicates that it’s a polar molecule.

Now, where to draw the dipole? If we look at the first drawing, we can see that everything looks pretty symmetrical, making it impossible to show the electrons more on one side of the molecule than the other. However, the drawing on the right switches one of the hydrogen atoms with a fluorine, giving us a structure where we *can* show the dipole:

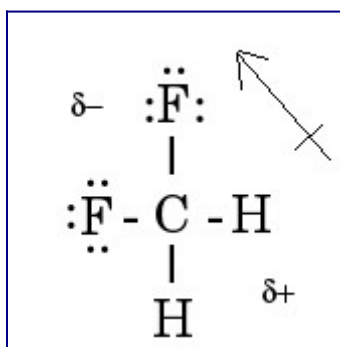


And this is your final answer.

Except... What would happen if you had instead switched two different atoms with each other than the one I showed? For example, let’s say we got this diagram instead:



It's a valid Lewis structure and it's shown asymmetrically. As a result, we can go ahead and draw the dipole arrow:



Even though it's a little different than the other answer I said was correct, you can see that they're basically the same thing, with the dipole pointing toward the side of the molecule with the fluorine atoms and the hydrogen atoms having partial positive charge. It's the same thing, only viewed from a different angle.

And with that, you're a polarity pro!

Footnotes:

1. It's a little more complicated than this in practice due to something called the [inductive effect](#). The idea behind the inductive effect is that the "pull" of atoms in a bond may be somewhat different depending on the other atoms stuck to it. For example, a C-C bond will be nonpolar if both carbons have the same atoms attached, but if one of the carbon atoms has three fluorine atoms, the highly electronegative fluorine atoms will cause that carbon to pull strongly on the C-C bonding electrons. This is fairly important in organic chemistry, but you're not likely to see it in high school.
2. These numbers are more-or-less arbitrary, and you may see different values cited by different sources. This takes place because polarity is a continuum where some things can be a little polar while others can be a lot polar. The cutoff between a little and a lot is arbitrary, so when people cite different numbers they're both basically right.