Fascinating Education Script

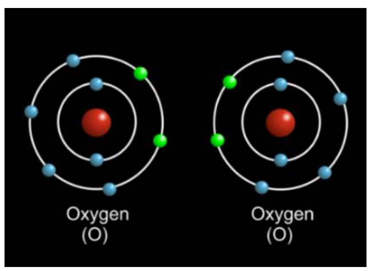
Fascinating Chemistry Lessons

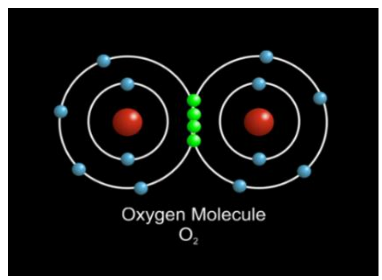
Lesson 4: The Polar Covalent Bond

Slide 1: Introduction

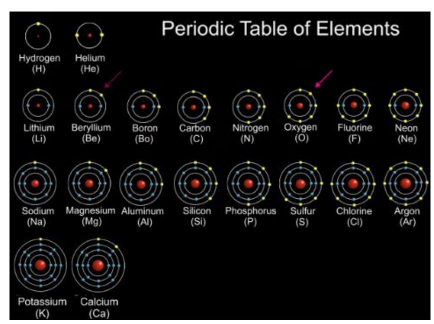
Slide 2: Giving away electrons

Oxygen atoms can bond to other atoms in a number of ways. One way is for an oxygen atom to pair up with another oxygen atom and share two electrons covalently, just as two chlorine atoms bonded covalently by sharing one electron. And like molecules of chlorine, molecules of oxygen would be nonpolar. Small nonpolar molecules like chlorine and oxygen form gases at room temperature.





Oxygen could also fill up its Ring 2 by taking two electrons from an atom looking to give away its two outer electrons, an atom like beryllium. An ionic intramolecular bond like this would make beryllium oxide polar and thus form solid crystals.



A third way oxygen could fill up its outer Ring 2 is to take an electron from two atoms, each willing to give away one electron.

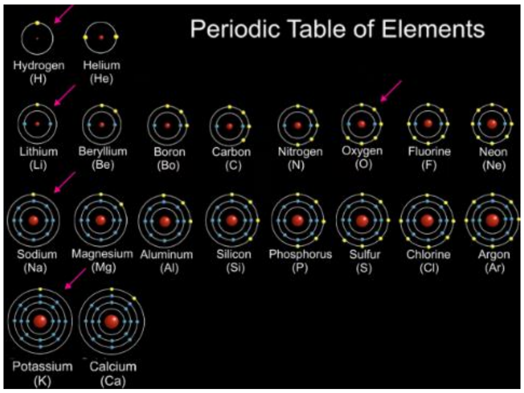
Which atoms would be willing to give away a single electron? Hydrogen, lithium, sodium, and potassium.

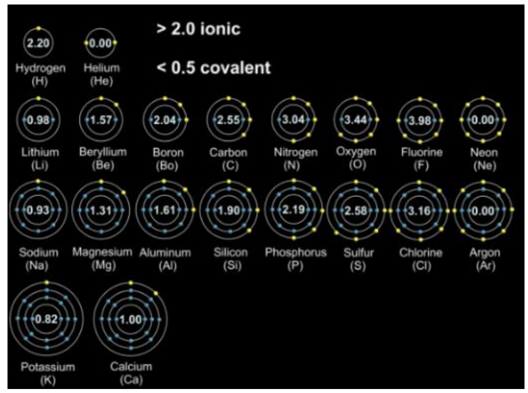
Slide 3: Pauling's electronegativity chart

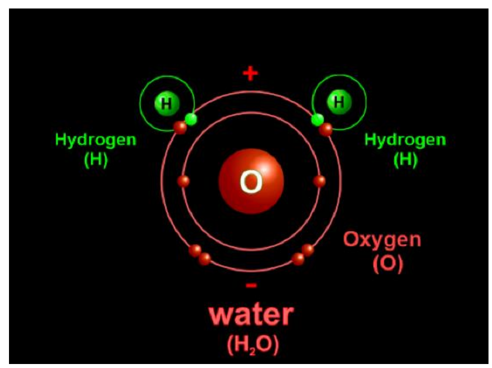
According to Pauling's chart, an electronegativity difference of greater than two predicts an ionic bond, and less than one-half predicts a covalent bond. Hydrogen with an electronegativity value of 2.20 and oxygen with an electronegativity value of 3.44, a difference of 1.24, would therefore form neither a covalent bond nor an ionic bond.

Oxygen is very electron hungry; only fluorine has a higher electronegativity value than oxygen. But hydrogen, you recall, also has a high electronegativity value because its electron is so close to its nucleus with no other electrons around to shield it from the nucleus.

As a result, oxygen can only share hydrogen's single electron, but because oxygen's nucleus is so much more positive than hydrogen's, and because oxygen's nucleus is not that far away from its Ring 2, oxygen is able to dominate the sharing and hog hydrogen's electron.





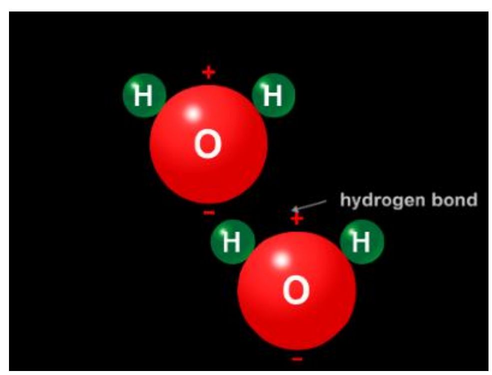


Meaning, that while oxygen and hydrogen both share an electron with each other, oxygen makes the shared electrons spend most of their time around the oxygen nucleus.

What this does is make the oxygen side of the molecule somewhat negative and the hydrogen side somewhat positive.

The scientific name for the unequal sharing bond between oxygen and each hydrogen atom is "polar covalent bond."

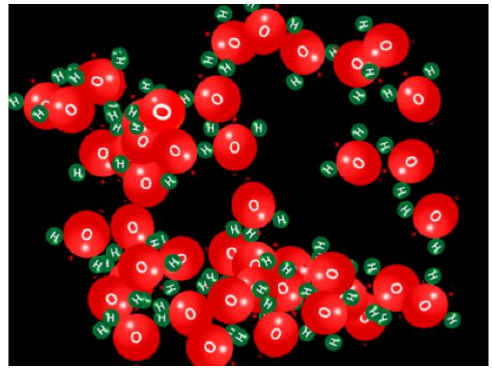
Oxygen bonded to two hydrogen molecules, H_2O, is of course water.



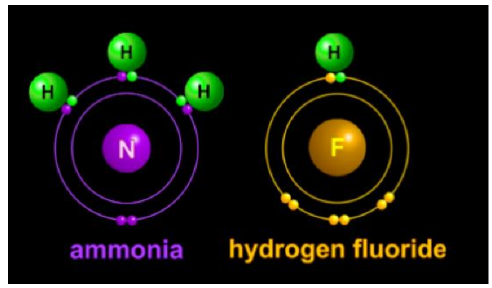
Why, though, would a molecule that's somewhat positive on one side and somewhat negative on the other side, be liquid at room temperature?

Slide 4: Hydrogen bond

The reason water molecules are liquid at room temperature is that by being somewhat positive on one side and somewhat negative on the other side, the positive side of one water molecule is now attracted to the negative side of another water molecule. So, when water molecules bump into each other, their opposite sides stick to each other for a brief moment with an intermolecular bond called a “hydrogen bond.”

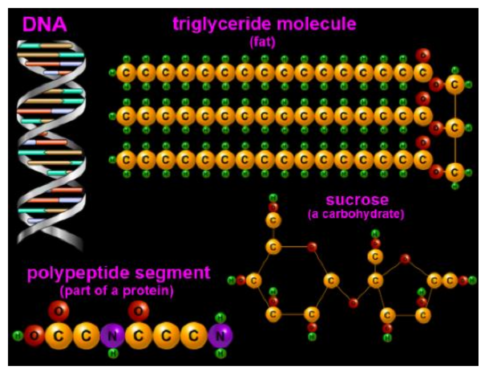


We can’t yet see molecules moving about, but even if we could, with water molecules bumping into each other a billion times a second, we still wouldn’t be able to see them sticking to each other before bouncing away. All we could ever see would be water molecules slipping and sliding over each other as a liquid.



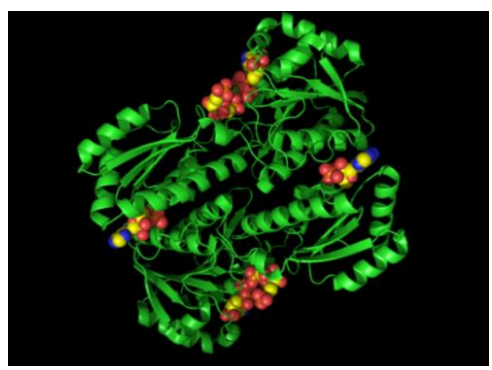
Oxygen and hydrogen are the most atoms involved in hydrogen bonding, but other atoms, like nitrogen and fluorine, are also strongly electronegative and as capable as oxygen at hogging hydrogen’s shared electron. So, molecules like ammonia, NH_3, and hydrogen fluoride, HF, also hydrogen bond to other polar molecules.

Conversely, hydrogen is not the only atom that oxygen can hog electrons from. Oxygen is so electronegative -- so electron-hungry -- that it can even form a polar covalent bond with carbon and hog carbon’s electrons. We’ll see some examples of this in the next slide.



Slide 5: Hydrogen bonds in nature

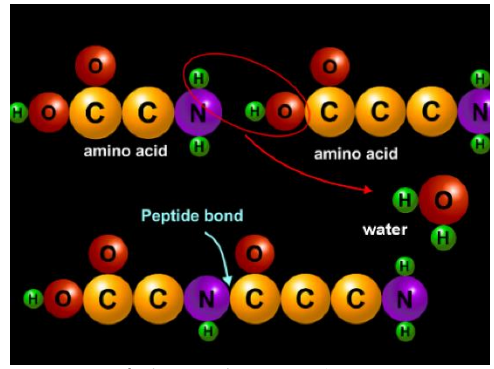
Hydrogen bonding is widespread in biological molecules – molecules like carbohydrates, DNA, proteins, and fats. The reason hydrogen bonding is so common is that carbohydrates, DNA, proteins, and fats, are made with oxygen and nitrogen atoms, which, being strongly electronegative, hog electrons shared with hydrogen and carbon atoms.

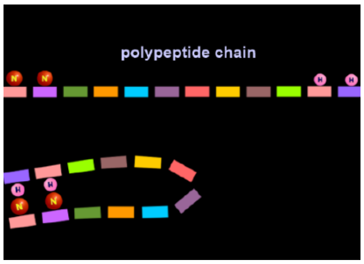


Proteins in particular depend on hydrogen bonding to fold themselves up into a specific shape that allows each protein to perform a specific task inside or outside a cell.

Proteins are made by attaching amino acids together into long chains called “polypeptides,” so-called because the bond holding one amino acid to the next amino acid is a “peptide bond.”

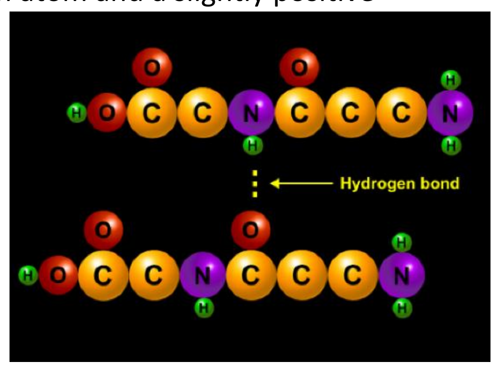
Each peptide bond is formed when a hydrogen atom at one end of an amino acid detaches from a nitrogen atom and joins the hydrogen and oxygen atom on another amino acid to form a molecule of water.





As more and more amino acids link up, a long chain of amino acids gradually assembles into a long polypeptide chain, which then folds up with the help of hydrogen bonding between different parts of the polypeptide chain. The protein now has a very specific shape and is ready to perform its single task.

Here is a typical hydrogen bond between a slightly negative oxygen atom and a slightly positive hydrogen atom whose own electron is being hogged by a nitrogen atom.



Hydrogen bonds are only moderately strong. It takes only a little heat energy to break hydrogen bonds and cause the protein molecule to unfold. A common example is the heating of the protein albumin in egg white. The process of unfolding a protein, called "denaturing," turns the protein in egg white opaque.

Cooking meat does the same thing to meat proteins. Even proteins in our brain and internal organs run the risk of denaturing when our body temperature rises above 105° Fahrenheit, 41° Celsius, a condition called "heat stroke."

Hydrogen bonding also occurs in DNA. DNA is a long strand of special molecules carrying a code of instructions to make the thousands of different proteins necessary for life. To protect the code from damage until the DNA code is ready to be read, the long strand of DNA code is covered up with a mirror image of the DNA code. The two strands of DNA stick to each other with hydrogen bonding. Because hydrogen bonding is only moderately strong, the two DNA strands can be split apart when it's time for the DNA code to be read and begin protein synthesis.

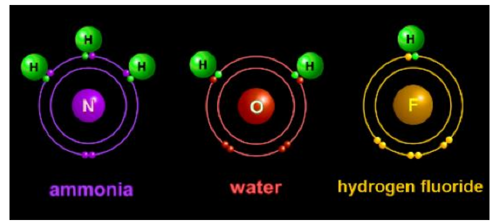
Much of clothing is made of cotton fibers consisting of long chains of carbon, oxygen, and hydrogen atoms. Hydrogen bonding between fibers pulls the fibers together, creating wrinkles in the fabric. Hot steam from an iron is able to break those hydrogen bonds and smooth out the wrinkles.

The reason foods like spaghetti stick to pots and pans is in large part due to hydrogen bonding between carbohydrates in the food and polar molecules in the pots and pans.

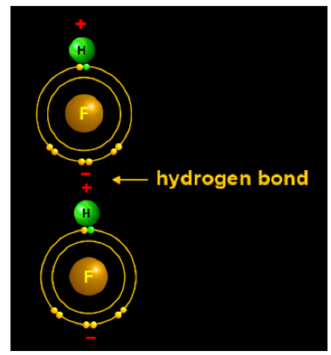
Slide 6: Catching our breath

Let's catch our breath.

So, what do we know about the polar covalent bond?

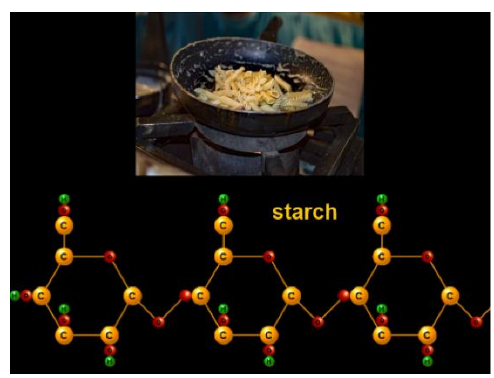


We know that the polar covalent bond is an intramolecular bond wherein one atom of the molecule, typically an atom of oxygen, nitrogen, or fluorine, has a stronger pull on the shared electrons than the other atom, which is usually a hydrogen atom. Because the shared electrons are forced to spend more time around the oxygen, nitrogen, or fluorine atom, the molecule becomes somewhat polar, with the hydrogen side becoming somewhat positive and the oxygen, nitrogen, or fluorine side becoming somewhat negative.



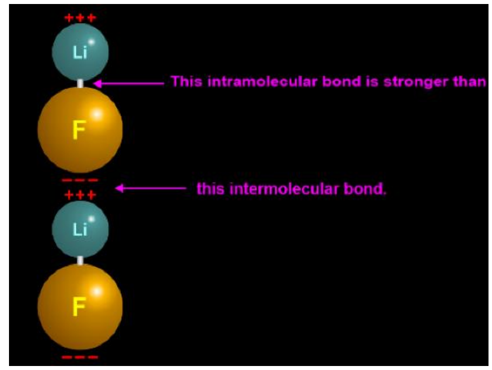
As you'd expect, the positive side of one polar covalent molecule is attracted to the negative side of another polar covalent molecule. The intermolecular bond between two polar covalent molecules is called a "hydrogen bond," because hydrogen is so often the positive atom, but even when the positive side is not a hydrogen atom, it's still called a hydrogen bond.

The strength of hydrogen bonding between polar covalent molecules is greater than the attraction between nonpolar molecules, but not as strong as the attraction between ionic molecules. Polar covalent molecules tend to stick weakly to each other, and then quickly detach to form another brief hydrogen bond with another polar covalent molecule. When seen from a distance, polar covalent molecules look like they're slipping and sliding over each other. To the naked eye, polar covalent molecules look fluid and liquid. Water is a perfect example.



Besides hydrogen and oxygen, carbon and oxygen also share electrons unequally. Carbohydrates, like the starch in spaghetti and other pastas, are loaded with carbon-hydrogen and carbon-oxygen bonds that, if you let the food dry out, hydrogen bond to metal atoms in the utensils and pots. Soaking everything in water before the food dries out allows water molecules to hydrogen bond to the carbohydrates before they can hydrogen bond to any metal surfaces.

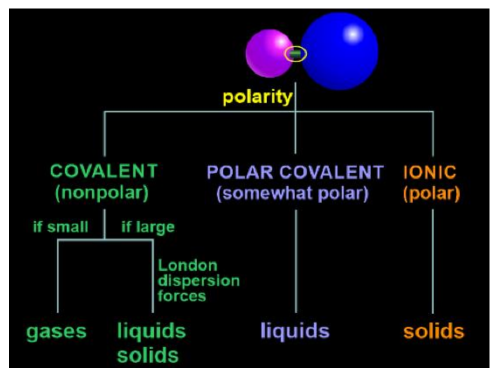
Slide 7: Intramolecular and intermolecular bonds



Recall that intramolecular bonds hold atoms together while intermolecular bonds hold molecules together.

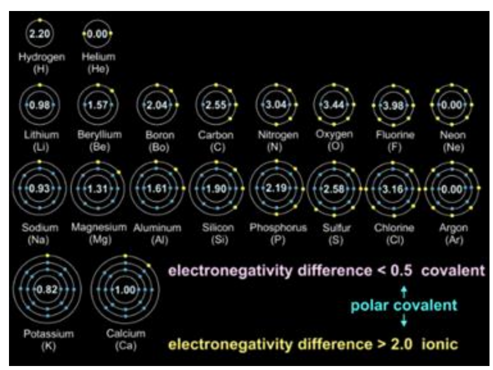
The intramolecular bond is stronger than the intermolecular bond, which is why, when molecules bump into each other, they break their intermolecular bond and simply bounce away instead of breaking their intramolecular bonds and busting apart.

The intramolecular bond determines the polarity of the molecule and thus the strength of the intermolecular bond. The strength of the intermolecular bond, in turn, is the major predictor of a molecule's behavior. So, in order to predict how a molecule behaves, we need to know which intramolecular bond formed the molecule: the covalent bond, the polar covalent bond, or the ionic bond.



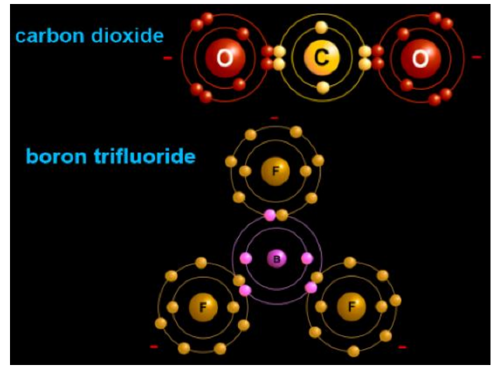
Covalent bonds produce the least polar molecules, which, if they're small, form gases. Larger covalent molecules are also nonpolar but instead of remaining as gases, large nonpolar molecules can form liquids and solids by using London dispersion forces to pull them all together and form liquids and solids.

Ionic bonds form the most polar molecules and become solids. In between are the polar covalent bonds that produce somewhat polar molecules that stick only briefly to each other and form liquids, water being the best example.

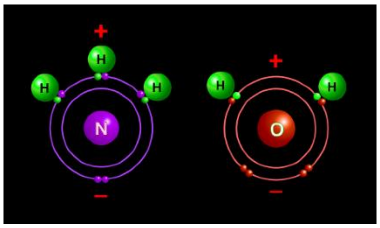


When a molecule is made up of only two atoms, Pauling's electronegativity chart predicts covalent bonding when the difference in electronegativity is less than 0.5, ionic bonding when the difference in electronegativity is greater than 2.0, and polar covalent bonding when the difference in electronegativity is between 0.5 and 2.0.

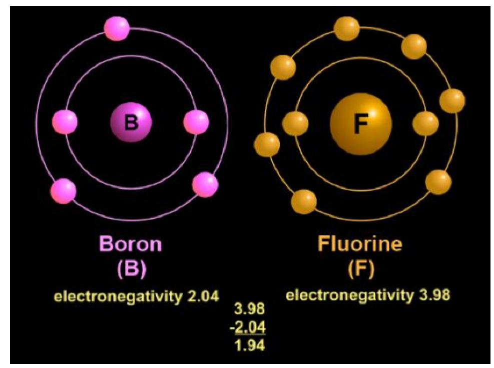
What I want to show you now is that while the electronegativity difference between two atoms may very well predict the type of intramolecular bond and thus the molecule's polarity, the electronegativity difference, alone, doesn't predict the polarity of molecules made up of more than two atoms – molecules like carbon dioxide or boron trifluoride. For molecules made up of three or more atoms, polarity is also determined by the arrangement of the atoms in the molecule.



We're about to see two very important things. One is that in molecules made up of three or more atoms, polar bonds between the atoms can still form a nonpolar molecule if the bonds are placed symmetrically in the molecule. The other thing we're about to see is that a polar molecule made up of three or more atoms can be made even more polar if one of the atoms has valence electrons not being shared with another atom.



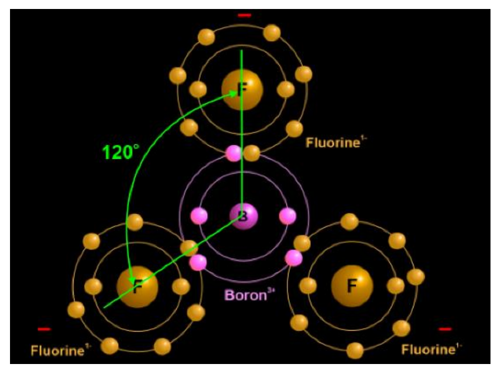
We'll begin by showing how the arrangement of polar intramolecular bonds in a molecule can make the molecule nonpolar.



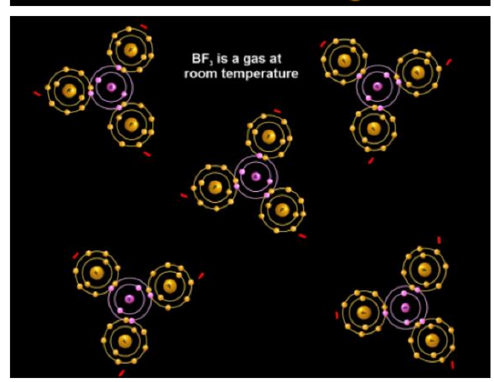
Slide 8: Nonpolar molecules with polar bonds

Boron with three electrons in its Ring 2 could give one electron to three different fluorine atoms looking for one electron to fill up their Ring 2.

The electronegativity difference for each boron-fluorine bond would be 3.98 for fluorine minus 2.04 for boron, or 1.94, right on the edge of being ionic.



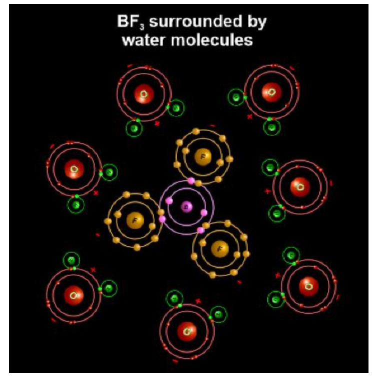
Yet, boron trifluoride, an industrial chemical, is not attracted to other boron trifluoride molecules even though all three of its intramolecular bonds are very polar. Each fluoride of a boron trifluoride molecule has a negative charge, so why isn't boron trifluoride polar?



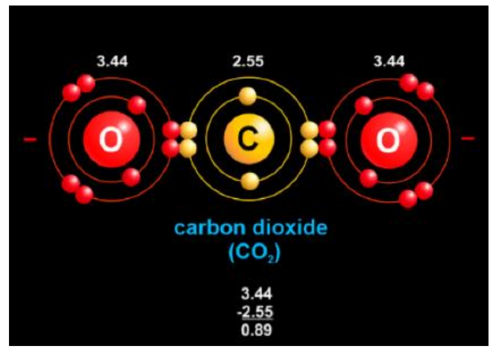
Because the three fluorine atoms are located as far away from each other as possible, meaning that no matter what direction you view boron trifluoride from, its electrical charge is the same.

Being nonpolar, then, means one thing when the molecule is among other identical molecules, and another thing when the molecule among truly polar molecules. Boron trifluoride, for example, is only nonpolar to other boron trifluoride molecules.

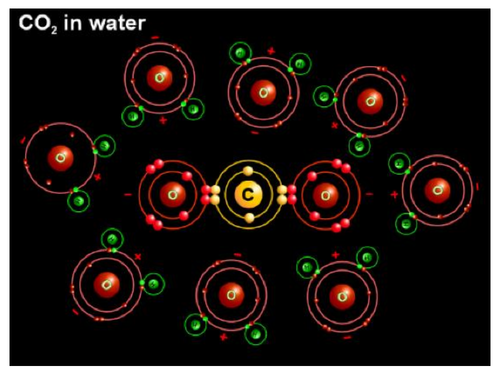
If boron trifluoride molecules were to come into contact with truly polar molecules, positive on one side and negative on the other, the negative sides of boron trifluoride would be attracted to the positive side of those polar molecule.



The same goes for carbon dioxide. Carbon dioxide's electronegativity difference is 3.44 for oxygen minus 2.55 for carbon, or 0.89. Each oxygen atom hogs the electrons it shares with carbon, making each end of the carbon dioxide molecule somewhat negative.

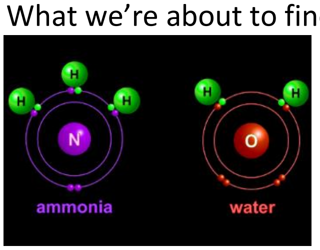


Carbon dioxide is a gas at room temperature because carbon dioxide molecules are nonpolar to each other. However, should carbon dioxide molecules come into contact with polar molecules, the negative oxygen ends of the carbon dioxide molecule would be attracted to the positive end of the polar molecule. That's why carbon dioxide is soluble in water. Each oxygen end of a carbon dioxide molecule is attracted to the positive side of a water molecule.

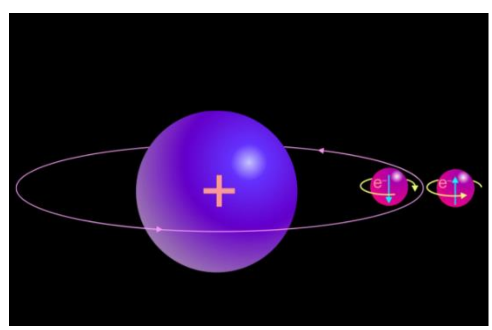


The polarity of a molecule, then, does depend on the electronegativity difference between its atoms, and the placement of atoms around a central atom, but even molecules that are nonpolar because of their symmetrical electrical charges can become polar in the presence of other molecules that are polar, the best example being molecules of water.

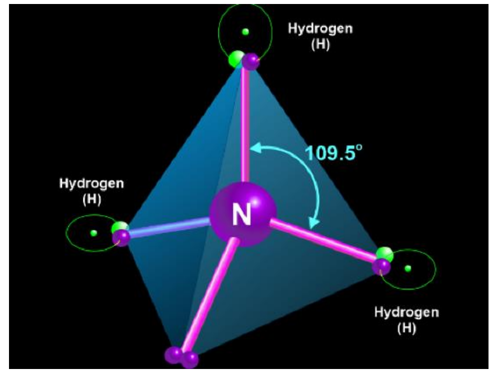
What we're about to find out now is that molecules like water and ammonia, consisting of a central oxygen or nitrogen atom bonded to two or three hydrogen atoms, are made even more polar by pushing the hydrogen atoms closer together. This is done with unshared electrons around the central atom. Let's see.



Slide 9: Electrons travel in pairs

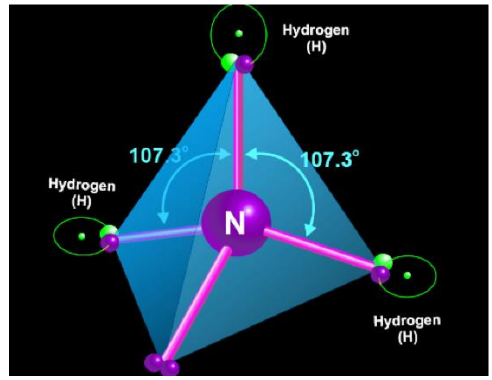


This is the first time that I've introduced the fact that electrons travel in pairs. The reason two electrons pair up is that electrons spin. When electrons spin, they create a magnetic field (represented by the up and down arrows). Electrons pair up because one electron spins clockwise and the other counterclockwise generating opposite magnetic fields that attract each other. Being negatively charged, though, each pair of electrons tries to remain as far away from every other pair of electrons as possible.



In three dimensions, eight electrons around a nucleus pair up into four pairs. Each pair situates itself as far away from every other pair as possible, which places each pair at the four points of a pyramid. The angle between any two pairs of electrons is 109.5 degrees.

When atoms share electrons, however, as nitrogen is doing with three hydrogen atoms in this molecule of ammonia, the electrons being shared lose some of their repulsive strength because they now have to spend some time around the hydrogen atom.

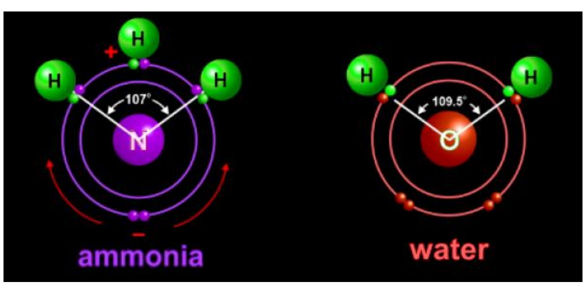


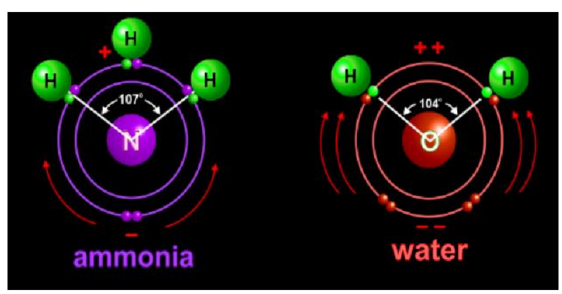
The pair of unshared electrons is able to push the three shared electrons – and their attached hydrogen atoms – slightly to one side of the ammonia molecule. The angle between the three hydrogen atoms narrows to 107.3 degrees.

Slide 10: Unshared and shared electrons

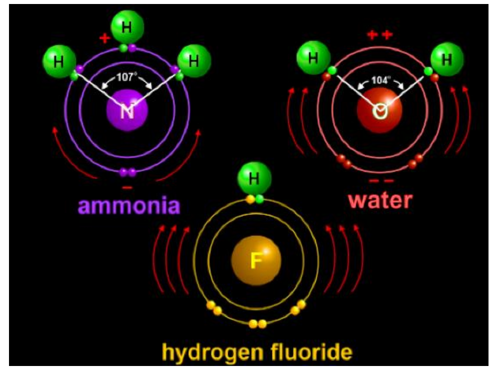
We just saw that in an ammonia molecule, NH_3, the three hydrogens are squeezed together by nitrogen's two unshared electrons. The hydrogen side of the ammonia molecule becomes slightly positive, leaving the opposite side of the nitrogen atom with the unshared pair of electrons slightly negative. This shift in bond angle makes ammonia somewhat polar.

Oxygen, though, is able to push its shared electrons much closer together than nitrogen and make water molecules even more polar than ammonia. How come?

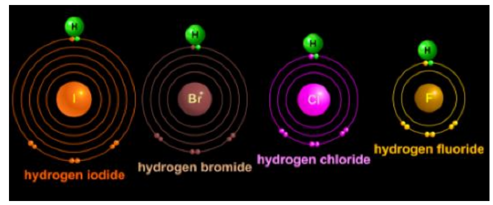




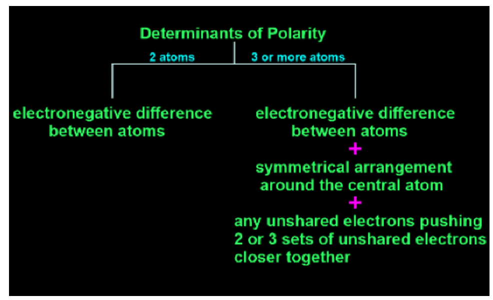
Because oxygen has two sets of unshared electrons and only two sets of shared electrons to be pushed together. When the two pairs of unshared electrons push the two pairs of shared electrons closer together, the angle between the two hydrogen atoms narrows from 109.5 to 104 degrees. The hydrogen side of the water molecule becomes quite positive, and the opposite side quite negative.



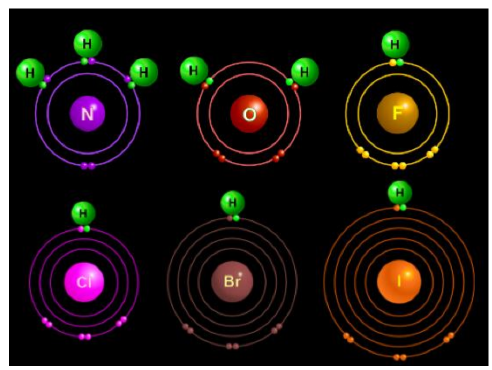
How about hydrogen fluoride? It has three sets of unpaired electrons and should therefore be able to push with a great deal of force. They sure can, but what are they going to push?



While hydrogen fluoride, hydrogen chloride, hydrogen bromide, and hydrogen iodide all have three sets of unshared electrons, they only have one set of shared electrons to push. The three sets of unshared electrons can push as hard as they want, but they can't make the hydrogen side of the molecule any more positive than it already is. That's why when there are only two atoms in a molecule, the polarity of the molecule does depend solely on the electronegativity difference.



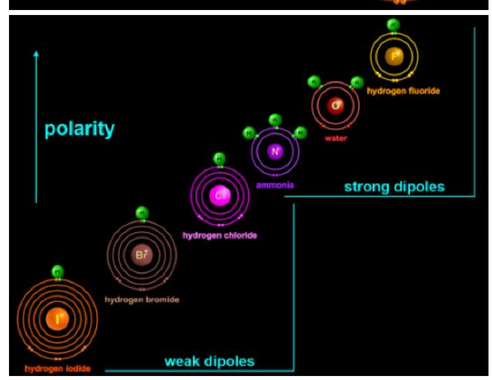
For a molecule made up of three or more atoms, though, polarity depends not only on the electronegativity difference between all the atoms in the molecule, but also on how symmetrically atoms are situated around a central atom, and on the ability of unshared electrons around the central atom to push two or three hydrogen atoms closer together. This narrows the bond angle between the atoms surrounding the central atom and makes the molecule even more polar.



How, then, would you rank the polarity of ammonia, water, hydrogen fluoride, hydrogen chloride, hydrogen bromide, and hydrogen iodide?

Slide 11: Polarity strength

These 6 molecules are arranged according to their polarity. Polarity is either strong or weak, but there is a way to measure polarity with numerical units, in which case, the degree of polarity is referred to as a "dipole" or "dipole moment."



By taking into consideration both the electronegativity difference and the ability of unshared electrons to push any shared electrons closer together, it turns out that the least polar molecule is hydrogen iodide, because even though iodine has many more protons in its nucleus than hydrogen and is able to hog the electrons being shared with hydrogen, the hogging is very weak because iodine's nucleus is shielded from those shared electrons by a very large number of interior electrons.

Bromine has fewer interior electrons than iodine, so hydrogen bromide is slightly more polar than hydrogen iodide, and for the same reason hydrogen chloride is slightly more polar than hydrogen bromide. Hydrogen iodide, hydrogen bromide, and hydrogen chloride are all weakly polar.

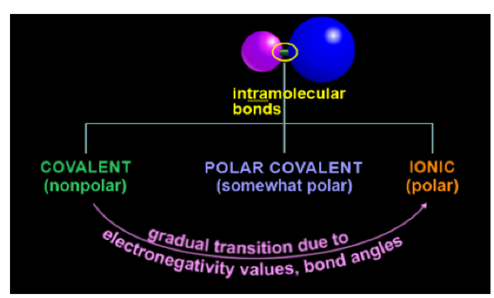
The strongly polar molecules include ammonia, NH_3, with only one pair of unshared electrons pushing together three pairs of shared electrons, then the somewhat more polar water molecule with two pairs of unshared electrons pushing together two pairs of shared electrons, and finally the strongest, hydrogen fluoride with a highly electronegative atom of fluorine.

Something doesn't make sense, though. If hydrogen fluoride is more polar than water, hydrogen fluoride should have a higher boiling point than water, because it would take more heat energy to pull hydrogen fluoride molecules apart and allow them to escape into the air. But hydrogen fluoride's boiling point is less than 20 degrees Celsius while water's boiling point is much higher, 100 degrees Celsius. Why is water a liquid at room temperature but hydrogen fluoride is a gas? Why is water stickier than hydrogen fluoride when hydrogen fluoride is more polar than water?

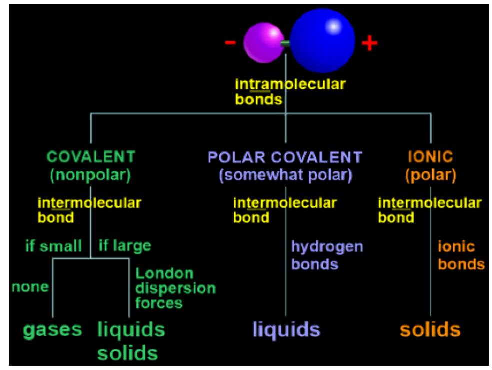
Because water, with two hydrogen atoms, can hydrogen bond to two other water molecules while hydrogen fluoride can only hydrogen bond to one hydrogen fluoride molecule. In order for water to break both of its intermolecular bonds and escape into the air, heat energy has to break both of water's hydrogen bonds, but only one hydrogen bond in order for hydrogen fluoride to boil, so hydrogen fluoride boils at a lower, in fact, much lower, temperature than water.

Slide 12: Polarity of molecules

What you should take away from all this are two things. The first is that the intramolecular bond determines the polarity of a molecule, which is not just nonpolar, somewhat polar, and polar, but rather a gradual transition from nonpolar to polar with all sorts of gradations in between caused by electronegativity values and changes in bond angles brought about by unshared electrons.



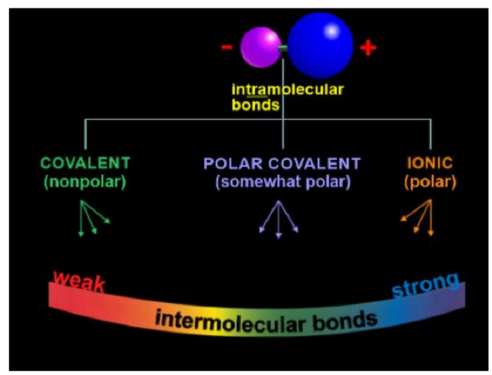
The second thing to take away is that while the intramolecular bond determines the polarity of the molecule, the polarity of the molecule determines the strength of the intermolecular bond holding molecules together. If the intermolecular bond is very weak, the molecules form gases, liquids if the intermolecular bond is somewhat strong, and solids if the intermolecular bond is very strong.



The only exception is large nonpolar molecules, because their large size and weight allows them to remain close by so that their London dispersion forces are able to pull them together into liquids and solids.

In short, once you know the intramolecular bond that two atoms use to bond together into a molecule, you know the polarity of the molecule, and from that, the strength of the intermolecular bond that the molecule forms with other molecules as they form gases, liquids, and solids.

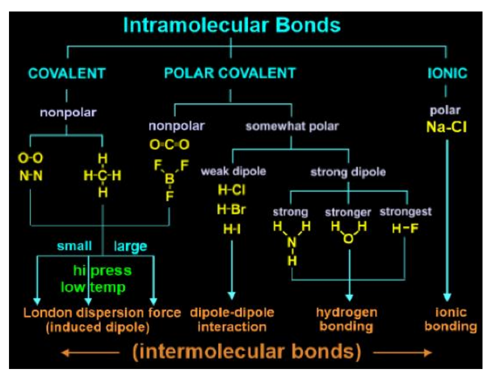
What this chart doesn't show, though, is that within each of the three types of intermolecular bonds there are gradations of strength. So while there may only be 3 states of matter – gases, liquids, and solids, depending on the strength of the intermolecular bonds between molecules and depending on how the atoms are arranged in the molecule, every gas, every liquid, and every solid will have different properties. By understanding those different molecular properties, we can then select which molecules are most suitable for materials that need to be, say, strong, flexible, waterproof, lightweight, durable, and so on.



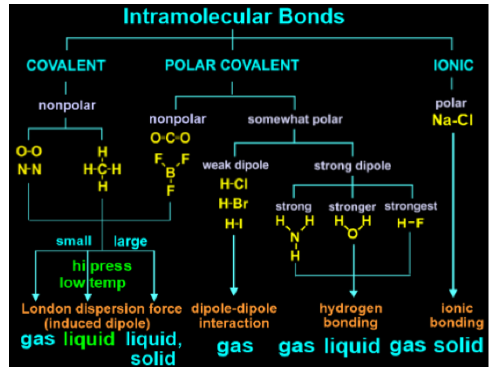
What we need to do now is look at how covalent, polar covalent, and ionic bonds are able to produce intermolecular bonds with such a wide spectrum of strengths.

Slide 13: Spectrum of intermolecular bonds

Shown up top in blue are the 3 intramolecular bonds: covalent, polar covalent, and ionic. At the bottom, in orange, are the intermolecular bonds resulting from each intramolecular bond, ranging from weak London dispersion forces to slightly stronger dipole-dipole interactions, even stronger hydrogen bonding, and finally reaching the strongest intermolecular bond, the ionic intermolecular bond.



This flow chart shows how the covalent, polar covalent and ionic intramolecular bonds end up producing a full spectrum of intermolecular bonds.



Beginning with the intramolecular bonds up top, covalent bonds produce nonpolar molecules either by bonding two identical atoms or by placing polar bonds symmetrically around a central atom as with methane.

Polar covalent bonds can also produce nonpolar molecules by placing polar covalent bonds symmetrically around a central atom, as in carbon dioxide and boron trifluoride.

Small, lightweight nonpolar molecules are gases at room temperature, but with high pressure and low temperature, London dispersion forces can make small, lightweight nonpolar molecules stick to each other and form liquids such as liquid nitrogen and liquid argon.

Unlike small, lightweight nonpolar molecules, large and heavy nonpolar molecules, like long-chain triglycerides, are so sluggish that they don't ricochet very far after bumping into other long-chain triglycerides. By remaining close, and having so many atoms, long-chain triglycerides are easily pulled together by London dispersion forces to form liquids and even solids at room temperature and pressure. The longer the triglyceride, the more London dispersion forces between two triglycerides. And the more saturated the triglyceride, the easier it is for them to lie flat against each other and allow London dispersion forces to pull them together into liquids and solids.

Do you see, then, that all nonpolar molecules, whether they were made from covalent intramolecular bonds or polar covalent intramolecular bonds, require momentary London dispersion forces to pull them together into liquids and solids. Recall, though, that carbon dioxide and boron trifluoride molecules, being negative all the way around, are attracted to polar molecules like water, and thus dissolve in water using hydrogen bonding.

The polar covalent bonds form weakly polar and strongly polar molecules. The weakly polar molecules form when chlorine, bromine, and iodine become the negative side of the molecule. As these "weak dipoles," hydrogen chloride, hydrogen bromide, and hydrogen iodide, approach each other, the weakly positive side of one molecule is drawn to the weakly negative side of another molecule, and they pull together in what's called a "dipole-dipole interaction." A dipole-dipole interaction is stronger than a London dispersion force, which is only a momentary force. A dipole-dipole interaction is strong enough that these molecules don't require the high pressures and low temperatures that nonpolar molecules do in order for them to stick together as liquids.

The strongly polar molecules are those involving the more electron-hungry atoms: fluorine, oxygen, and nitrogen. These molecules, known as the "strong dipoles," range from ammonia, a strong dipole, to water, a little stronger dipole, to hydrogen fluoride, the strongest dipole. The strong dipoles bond to each other with strong hydrogen bonding, stronger than dipole-dipole interaction.

The most polar intramolecular bond, the ionic bond, produces the strongest intermolecular bond, the ionic intermolecular bond. Sodium chloride is an excellent example. The ionic intermolecular bond almost always produces solids at room temperature and pressure.

What you should realize from this chart is that the only thing molecules understand is positive and negative. The polarity of a molecule is essential because it's the only thing that attracts one molecule to another. Understand a molecule's polarity and you understand how it behaves around other polar and nonpolar molecules. To understand a molecule's polarity, though, you have to understand how the molecule came to be polar or nonpolar, or somewhere in between. That's what this chart will help you with.

Slide 14: What you know so far

1. Oxygen could fill up its outer ring by sharing an electron with two separate hydrogen atoms, but the sharing would be unequal, because the electronegativity difference between oxygen and hydrogen is 1.24, which lies between give-and-take and equal sharing.

2. Oxygen hogging two electrons from two separate hydrogen atoms is water.

3. A water molecule is quite polar. Its dipole (two poles) makes water molecules somewhat sticky, enough to form a liquid at room temperature.

4. Polar intramolecular bonds can produce nonpolar molecules if there is more than one polar bond and they are placed symmetrically around a central atom.

Slide 15: What you know so far

5. Electrons travel in pairs whenever possible.

6. Pairs of electrons try to stay as far away as possible from other pairs of electrons.

7. An unshared pair of electrons exerts a stronger repulsive force on other electrons than a shared pair of electrons. Unshared electron pairs can thus push intramolecular bonds away from them and narrow their bond angle.

8. In the water molecule, oxygen's unpaired electrons narrow the bond angle of both hydrogen atoms (from 109^o to 104.5^o), making the water molecule more polar. Similar narrowing of the bond angle occurs in ammonia, between the nitrogen atom and its three hydrogen atoms, making ammonia quite polar, too.

Slide 16: What you know so far

9. Besides oxygen and nitrogen, polar molecules commonly form when hydrogen atoms bond to fluorine atoms. Nitrogen, oxygen, and fluorine are the three most electron hungry atoms.

10. Only one factor determines the polarity of a molecule with two atoms: their electronegativity difference. In molecules with three or more atoms, polarity also depends on the arrangement of atoms around a central atom, and the bond angle between those surrounding atoms.

11. The bond angle between atoms surrounding a central atom can be narrowed by any valence electrons around the central atom.

12. Hydrogen bonding is an intermolecular bond that occurs between the hydrogen atom of one molecule and the nitrogen, oxygen, and fluorine atoms on another molecule. Hydrogen bonding helps proteins fold up into their proper shape.

Slide 17: What you know so far

13. A molecule's intramolecular bond is always stronger than its intermolecular bond.

14. Water is less polar than hydrogen fluoride, but water's boiling point is much higher than hydrogen fluoride because each water molecule hydrogen bonds with two other molecules while hydrogen fluoride only hydrogen bonds with one other molecule.

15. The spectrum of intermolecular bonds ranges from no intermolecular bond at all to weak London dispersion forces to dipole-dipole interactions to hydrogen bonding to strong ionic bonding.