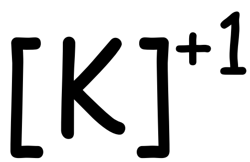
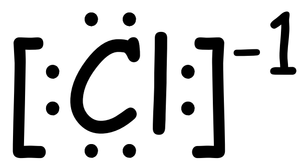
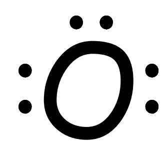
Bonding Basics

Chemistry is the study of the stuff of the universe, and importantly, how it forms into new substances by bonding in certain ways (or un-bonding to become simpler). There are several ways atoms can bond together in high school chemistry. We will look over each type, learning the particular ways they work, nd understand their differences.

Most of the bonds we will see are inside compounds, bonding hydrogen to oxygen when water forms, or sodium ions to chloride ions when sodium chloride forms. There are some types of bonds between particles as well. Finally, there are bonds that hold metals together as solids, and help us to understand how metals exist with their special properties of electrical conduction, and their ability to bend and not shatter.

LEWIS DOT DIAGRAMS

In order to help "see" how bonding works, a chemist named Dr. Lewis developed a diagram method for atoms, ions, and compounds. We will draw many in class. The diagrams look like these.

			
Helium atom	Potassium cation	Chloride anion	Oxygen atom

Atoms like helium, and oxygen show all of their valence electrons. These electrons tend to PAIR UP, which is part of the suborbital system of chemistry that we don't spend any time on, just remember that.

The potassium cation has lost it's outermost electron, and the whole valence orbital at the same time. It becomes 2-8-8-0 It ends up with $19p^+$ and only 18 e^-, making it have an overall charge of +1.

The chloride anion started out as a chlorine atom. It started with a 2-8-7 electron configuration, but gained one electron into it's third, or valence orbital. It becomes a -1 anion with 2-8-8 electron configuration. We draw all 8 dots here.

Lewis dot diagrams for atoms show all valence electrons. Cations show the new, "empty" valence orbital in brackets with a charge to show you KNOW what's going on. Anions end up with FULL VALENCE orbitals, which show ALL dots, and have brackets and charges as well.

IONIC BONDING

This type of bonding is the simplest to understand for the new students of chemistry. Whenever metals and nonmetals bond together it's this way: Metals will lose electrons and form into positive CATIONS. The metal ions TRANSFER their valence electrons to the nonmetals, which form into negative ANIONS. This transfer of electrons is always "perfect", the number of electrons lost by the metals are the same number "gained" up by the anions. No extra electrons, or left over electrons are allowed. The most common ionic compound is sodium chloride, table salt, the formula is NaCl.

To quickly review what we learned earlier in the year, metals will lose enough electrons to become ISOELECTRIC to a noble gas. They lose enough electrons to get a noble gas electron configuration.

A sodium atom has a 2-8-1 electron configuration. It will become a Na^{+1} cation with a 2-8 configuration, it loses one electron (it transfers this electron to a nonmetal, possibly chlorine, it doesn't actually LOSE it!)

Aluminum has a 2-8-3 electron configuration. It must "lose" 3 electrons to become isoelectric to neon.

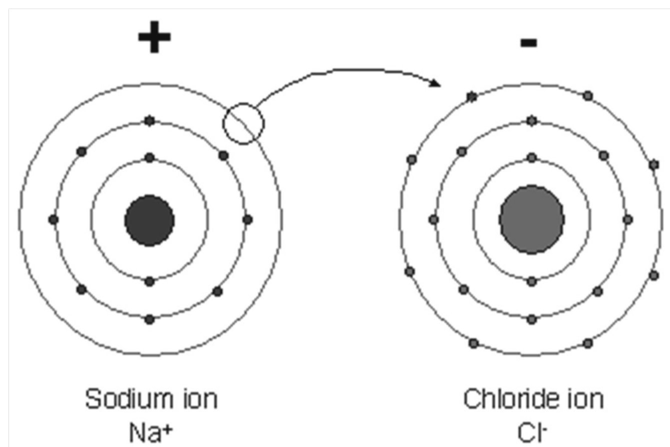
$Al \rightarrow Al^{+3}$ with a 2-8-0 cation electron configuration.

Nonmetals, like chlorine, gain electrons to form into negative anions. Chlorine atoms have 17 electrons in a 2-8-7 configuration. When Cl^{-1} forms, it's configuration changes to 2-8-8, which is isoelectric to argon.

Cations can only form simultaneously with anions. The transfer of electrons is always perfectly balanced, and that keeps all ionic compounds that form electrically neutral (the positive charges = the negative charges). You can't have a jar of cations, nor can you have a test tube of anions, they only exist at the same time.

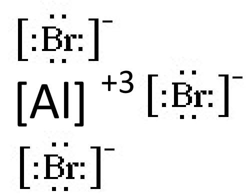
To draw atoms and ions (and compounds) in ways to help us understand this electron transfer we use LEWIS DOT DIAGRAMS. These show the VALENCE ELECTRONS only, which are the electrons in the outermost orbital of the atom, NOT all of the electrons, just the outside electrons.

At right is a model of the sodium cation that has already "lost" an electron (and it's WHOLE VALENCE ORBITAL), and the chloride anion which gained that electron. The ions opposite charge makes them very attracted together, which is called an ionic bond. The electron has been TRANSFERRED from $Na \rightarrow Cl$

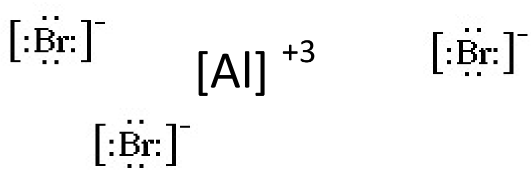


When ionic compounds form, we can draw Lewis dot diagrams for them as well. They are not exactly "pretty" but they are obvious. Just push the cation diagram close to the anion diagram together to indicate that they are making a compound. This is KCl...

If the compound has more than two ions (say $CaCl_2$ or $AlBr_3$) just push the ion diagrams close. There is no "correct" way to do this, literally, just make them close.



The differences here should be noted. The chloride ions are both labeled with the (-) sign but not (-1). That is fine. The Calcium cation has a 2+ instead of +2, another difference not worth worrying about. The bromides at right all show with a (-) sign, not (-1) too. These are minor style differences.



This arrangement at left DOES NOT show ionic bonding, it's more like 4 Lewis dot diagrams that are near each other.

Ionic bonding diagrams are CLOSE together, like $CaCl_2$ and $AlBr_3$ above.

COVALENT BONDING

When metals and nonmetals bond, they form ions first, then are attracted together by opposite charge. When 2 or more nonmetals bond together (no metals allowed, ever), they DO NOT FORM IONS. The atoms still try to become ISOELCTRIC to the noble gases, but they do not transfer electrons to do this. Instead, when two atoms make a covalent bond they SHARE valence electrons. By sharing, both atoms can share full orbitals.

This SHARING of ELECTRONS can be a perfectly even sharing (like best friends) or be uneven sharing (like me and you and one piece of cherry pie ala mode!).

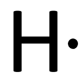
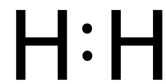
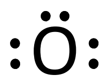
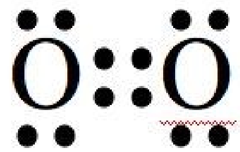
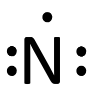
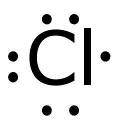
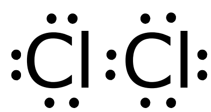
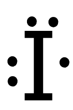
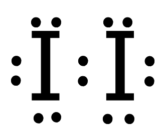
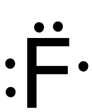
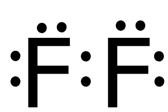
Nonmetals share enough electrons to get FULL ORBITALS, usually that means 8 electrons, but it's only 2 electrons for the smallest atoms. Rarely there are exceptions.

Water, carbon dioxide, and nitrogen gas all make types of covalent bonds. We will examine them now.

We can make Lewis dot diagrams that show ALL valence electrons, being shared, or simplify them by only showing the bonding, each pair of electrons is shows as a single dash.

Simple molecular compounds share either one pair of electrons (like H_2, Cl_2, Br_2, I_2 and F_2) or share two pairs of electrons, like O_2, or share three pairs of electrons, like N_2.

These bonding pairs of electrons can be changed to dashes once you can readily count electrons without making silly mistakes!

ATOM	MOLECULE	Dots replaced by dashes
		H–H
		O=O
		N≡N
		Cl–Cl
		Br–Br
		I–I
		F–F

The HONClBrIF twin molecules exhibit a variety of covalent bonds. H_2, Cl_2, Br_2, I_2 and F_2 all share one pair of electrons. This is called a single covalent bond. Since these atoms have the same exact electronegativity value, these bonds are also nonpolar. We call them: SINGLE NONPOLAR COVALENT bonds.

Oxygen molecules must share 2 pairs of electrons. The electronegativity difference is also zero, so these are DOUBLE NONPOLAR COVALENT bonds.

Nitrogen must share 3 pairs of electrons. The electronegativity difference is again zero, so these are called TRIPLE NONPOLAR COVALENT bonds.

Bonds are IONIC when formed from ions.

Bonds are COVALENT when two or more nonmetals share electrons.

Covalent Bonds can be POLAR or NONPOLAR bonds.

Covalent Bonds can be SINGLE, DOUBLE, or TRIPLE bonds.

These are important things to notice, and not get mixed up about

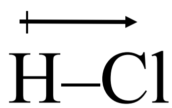
molecules	Share this many pairs of electrons	Share this many total electrons	Name of the bond
F_2	1	2	single nonpolar covalent
H_2	1	2	single nonpolar covalent
O_2	2	4	double nonpolar covalent
N_2	3	6	triple nonpolar covalent
HCl	1	2	single polar covalent
NaCl	0	0	ionic
MgO	0	0	ionic
CH_4	1	2 in each of the 4 bonds	4 single polar covalent

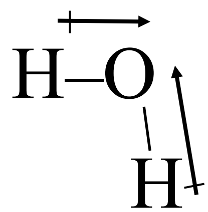
Electronegativity means tendency to gain electrons in a bond. The higher electronegativity "gets" electrons more of the time than the lower value. So, atoms with higher electronegativity will tend to get the electron and "be" negative more of the time. The atom that gets the electron LESS of the time tends to be the more positive side of the bond. The POLAR BOND has a positive and a negative pole (most of the time).

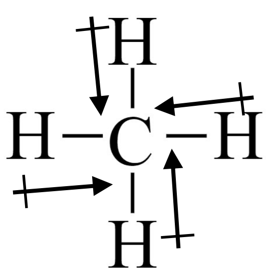
We can SHOW this polarity with a DIPOLE ARROW. This arrow shows the DIRECTION that the electrons go - to what side of the bond. The arrow ALSO shows what side is left "MORE POSITIVE" because the electrons moved to the other side of the bond. A dipole arrow looks like this:

This side is more positive	This side is more negative
The electron <u>left</u> this side	The electron <u>ends up over here</u> most of the time

A couple of examples of molecules with dipole arrows are:







With methane, CH_4, each bond is polar because the carbon atom has the higher electronegativity value, making the electrons spend MORE TIME with carbon, leaving the hydrogen atoms "more positive" more of the time.

The dipole arrow "shows" this quickly.

In these examples, hydrogen has a LOWER electronegativity value, so the arrow heads point to the atom with the higher electronegativity value. The Cl in HCl becomes negative most of the time because chlorine "gains" that electron most of the time. The O in water also becomes more negative more of the time. The carbon atom becomes more negative more of the time too.

The two atoms DO NOT SHARE ELECTRONS EVENLY. In water, the oxygen atoms make two different bonds with the two hydrogen atoms, each one is sharing unevenly. In methane there are FOUR single polar covalent bonds.

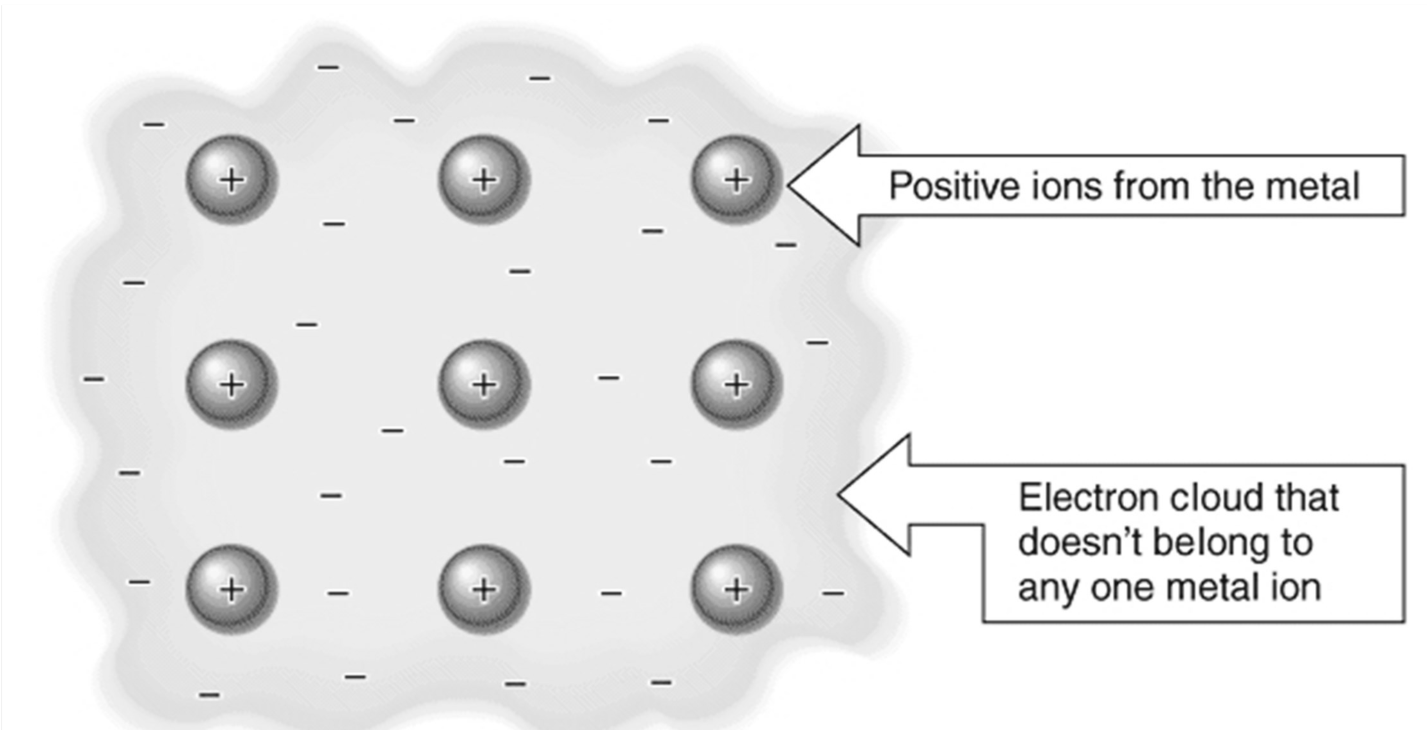
All of these three bonds are SINGLE POLAR COVALENT, because there are NO METALS bonding it must be covalent; they share one pair of electrons, and they share them unevenly.

METALLIC BONDING

When a frying pan is made, or if you take a fork out of your kitchen drawer, that hunk of metal displays many interesting properties. First, it has been stuck in that position for a long time, and it will likely hold this shape indefinitely.

It conducts heat well. It will also conduct electricity (as do all metals). If you smash it with a hammer, it will bend but it won't crack. The reason scientists believe metals do these three things, instead of not conducting electricity, or cracking like glass, is because of how they describe how metals bond together.

Metals are described not as packed atoms, awash is a sea of valence electrons. Poetic really



If the metal cations are crushed together, say by banging your house key with a big rock, the positive cations would get closer, and would want to repel; they would crack apart. This DOES NOT HAPPEN, the metal is able to change form, because the loose electrons moving at near the speed of light, move towards these squished together cations, and they offset that overly positive charge with some negative charge, keeping the metal electrically neutral, and the meal bends instead of cracking.

Electricity is described as moving electrons. If you run electricity into one side of a metal wire, electrons move into the metal, immediately disrupting the neutrality of the metal. Out the other side flows an equal number of electrons to complete the current. The electrons that flow into the metal are NOT NECESSARILY the ones flowing out the
other side. The electrons are almost like water flowing through a pipe, although this pipe is a wire and it's a solid.

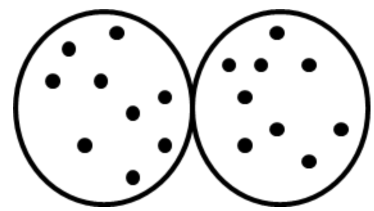
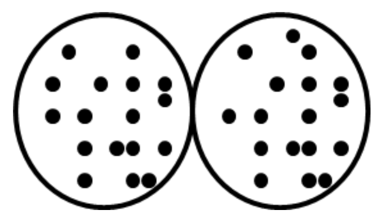
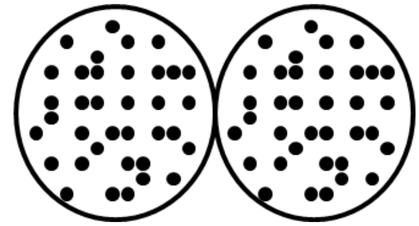
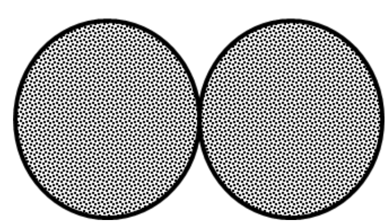
Loose valence electrons, in a packed cation solid, explain most of the metallic properties.

INTERMOLECULAR BONDING

Bonding INSIDE compounds or inside metals are clearly covered with ionic bonds, covalent bonds, and with metallic bonding. There is also bonding between particles of a gas, particles of a liquid, or particles of a solid. These are called inter-molecular bonds. Of all bonds, these are the weakest types, but they are still important and help determine phases of substances.

The three we will cover, from weakest to strongest are called: ELECTRON DISPERSION, DIPOLE ATTRACTION, and HYDROGEN BONDING.

Electron Dispersion Attraction or electron dispersion forces are due to the electrons of any atom or compound. Let's look at the atoms of group 17, the halogens to describe this intermolecular attraction.

Atom	Formula	Number of e^- in an atom	Number of e^- in a molecule	diagram
F	F_2	9	18 You can count them at right	
Cl	Cl_2	17	34 You can count these too.	
Br	Br_2	35	70 Hard to count, but there they all are.	
I	I_2	53	106 Too many electrons to count.	

In every atom and every compound, the electrons are moving very fast, and not in any exact orderly way (like the planets going around the Sun). The electrons are in orbitals, or ZONES, where they are most likely to be found, but they are not limited to any special exact spots.

At any instant of time, the electrons are somewhere. They might be completely spaced out evenly, and the whole molecule would be balanced or neutral. If the electrons were all slightly off to one side, for that instant that one side would be MORE NEGATIVE and the other side would be MORE POSITIVE, then it would change.

Over time, these instant points of negative or positive in the electron clouds have some attraction and some repulsions for other atoms or molecules. If you have few electrons, it's hard for this to amount to much positive or negative, and the force is terribly weak (real, but nearly insignificant).

At STP, both F_2 and Cl_2 are gases. The REASON they are gases is that the only attractive force pulling them together is ELECTRON DISPERSION forces. As the electrons are instantaneously dispersed, creating small and instantaneous positive and negative spots on their orbital clouds, this force even with 34 electrons in Cl_2 cannot overcome the kinetic energy of STP. These two elements are gases at normal temperature and pressure.

Bromine, Br_2, has 70 total electrons, and that many electrons dispersing at any point in time make MORE points of temporary positive and negative than fluorine or chlorine. Br_2 at STP is a liquid, because the electron dispersion attraction with this many electrons is enough attraction to hold the molecules together as a liquid (but not solid).

Iodine, I_2, has 106 total electrons, and a greater amount of electrons dispersing. When that many electrons are moving about, they create more moments of positive and negative, enough of them to pull this halogen into a solid at STP.

The three phases, gas, liquid, and solid are present in one group on the table; and these phases are caused only by the motion of the electrons in time, which create the weak but real electron dispersion force of attraction.

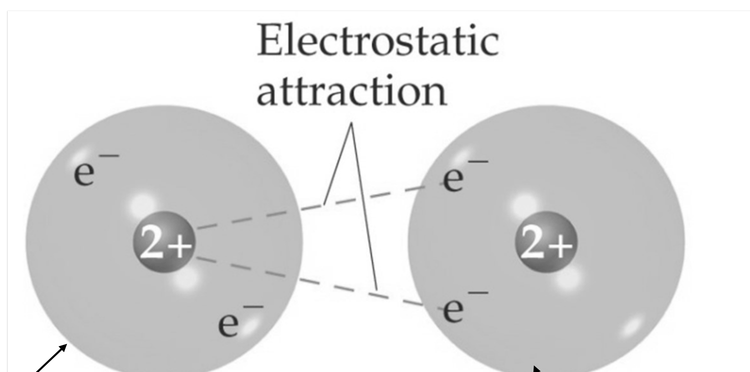
Electron Dispersion force used be called the

London Dispersion Forces

or the

van der Waal Forces,

but that's old fashioned.



On the right, the electrons in helium atom #2 are off to the left, creating a TEMPORARY, but real negative spot on the orbital cloud. That negative point is attracted to the positive nucleus of the left side helium atom. This lasts for A MOMENT, but new ones constantly appear as the electrons keep dispersing, or moving.

In this atom of helium at left: the electrons are dispersed evenly in this moment.

DIPOLE ATTRACTION

When atoms bond in covalent bonds they can share electrons perfectly together, if their electronegativity values are equal. Electronegativity means tendency to gain an electron in a bonding situation. If two atoms of fluorine bond, each has an EN Value of 4.0 which means that the difference in their EN Values is zero.

Neither atom of F gets the electrons they share more of the time. The bond is single NONPOLAR covalent.

Same with H_2, or Br_2. Both have single NONPOLAR covalent bonds.

And it's the same with O_2 although that is a double NONPOLAR covalent bond.

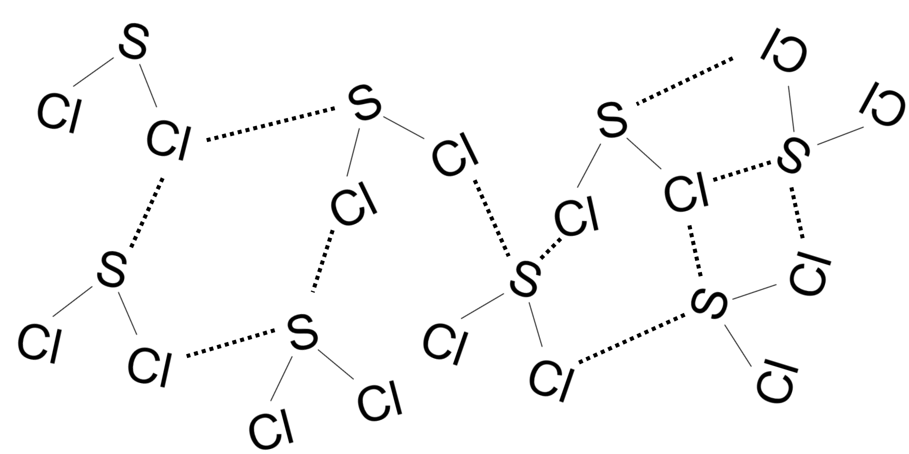
N_2 has a triple NONPOLAR covalent, but again NO DIFFERENCE in EN VALUES.

When bonds form and the bond is POLAR, because there IS A DIFFERENCE in EN Value, the atom with the higher electronegativity "gets" the electrons of the bond more often. That side of the bond is said to be more negative, the other side would be more positive.

These polar covalent bonds are sharing electrons, just not sharing equally. This unequal sharing will make the bonds almost always uneven, or POLAR.

When a whole bunch of sulfur dichloride molecules (SCl_2) are together, since the bonds between sulfur and chlorine have an EN Value difference (3.2 - 2.6 = 0.6 which is a polar bond), these bonds are almost always skewed so that the chlorine side is negative and the sulfur side is positive. Not a lot of positive or negative, actually just a little, but most of the time this polarity exists.

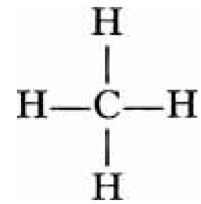
When these molecules are close together, they attract to each other's opposing poles. Chlorines (-) of one molecule are attracted to the sulfur (+) of other molecules, this near constant attraction is called DIPOLE ATTRACTION.



Above are many molecules of SCl_2, randomly dispersed. Remember, all the sulfur atoms are usually positively charge because of the EN Value differential with chlorine. The chlorine atoms are all usually negatively charged. This ALMOST CONSTANT POLARITY creates dipole attraction. This is also somewhat weak, but will make these molecules stick together better than molecules without it (see next page).

This DIPOLE ATTRACTION is indicated with the dotted lines.

With molecules of methane you need to see that the C-H bonds are polar, but…



The shape of this molecule is very important. The molecule has a balanced shape. Even though the hydrogen atoms are all mostly positive because the carbon in the center attracts their electrons most of the time, the positive charges sort of cancel each other out, because of the shape.

This molecule has RADIAL SYMMETRY. That is the same sort of balance as a pizza pie.

No matter how you cut a pizza pie (or cherry pie, or circle). If you go through the middle point, you get two equal halves. NO MATTER WHAT. If a molecule exhibits radial symmetry, it's balanced in shape, and the polar bond charges cancel each other out. This whole methane molecule has a positive outside, a negative in-side. When you put a bunch of methane together, all of the outside to all of the molecules are positively charged (most of the time), there is NO DIPOLE ATTRACTION.

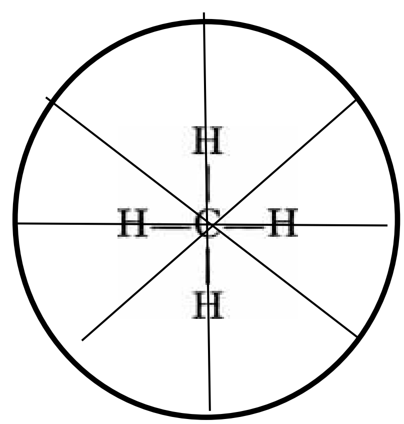
To get dipole attraction, you need polar bonds IN polar molecules. The shape of a molecule that determines if if is polar or nonpolar.

Radial symmetry is the only symmetry we care about in chemistry.

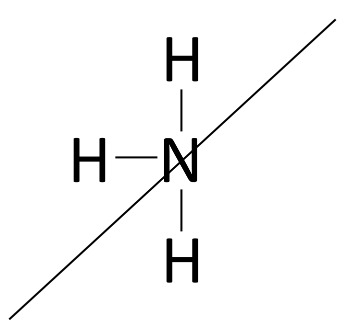
CH_4 has 4 polar bonds, but it has radial symmetry.

The molecule is nonpolar, or balanced.

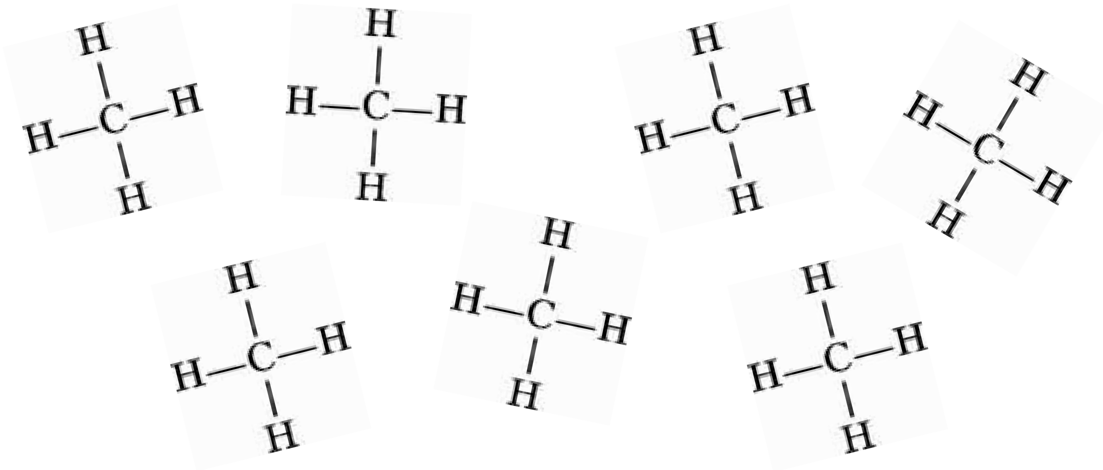
The distribution of the positive charges (in H) are symmetrical.



Humans (and gingerbread men) have bilateral symmetry. So does a water molecule. They have only one plane of symmetry. Only radial symmetry matters in chemistry. Water molecules are POLAR, the polarity of the bonds just can't cancel out.



In ammonia, each of the 3 bonds between N-H are polar. Since there are 3 hydrogen atoms bonded to the ammonia, when you cut it this way, you get one side with 2 H atoms, the other side has one H and two nonbonding electrons. This molecule does NOT HAVE radial symmetry, this is a polar molecule.



All of these molecules of methane have radial symmetry. They have polar bonds, but the molecules are balanced and NONPOLAR. They have almost no intermolecular attraction, except for electron dispersion forces. Unlike HCl or NH_3 CH_4 is a gas at STP.

BONDS can be polar, MOLECULES can be polar.
Bonds are polar when there is a difference in EN Values.
Molecules are polar when no radial symmetry exists.

HYDROGEN BONDING

When molecules that are polar and the bonds contain hydrogen, molecules like water and ammonia, not only are the bonds polar, but they are EXTRA POLAR because hydrogen has such a low EN VALUE.

For example:

SCl_2 has an EN Value difference of 0.6 between S and Cl. In water, the difference between oxygen and hydrogen EN Value is 3.4 - 2.2 = 1.2 which is a much greater polarity in the bond.

With Ammonia, nitrogen and hydrogen make an EN Value differential of 3.0 - 2.2 = 0.8, again much greater than the 0.6 differential in SCl2, these dipole attractions are so much greater, they have a different name.

This is sort of silly. Dipole attraction + super-duper dipole attraction would be fine with me, but not with the State Education Department!

So, there are dipole attractions when polar bonds exist in polar molecules. But if these polar bonds contain hydrogen, they are called HYDROGEN BONDS.

These intermolecular attractions are not really bonds either, they are super-duper dipole attractions.

In order of weakest to strongest, the intermolecular bonds are:

- electron dispersion
- dipole attraction
- hydrogen bonding

With intermolecular bonding, the electron dispersion attractions, the dipole attractions and the hydrogen bonding, these are real but much weaker than covalent bonding or ionic bonding.

This chart will hopefully help you keep track of what's going on in bonding.

All atoms and all compounds have electron dispersion attractions. They are caused by the temporary "moments" of positive or negative charge created by the motion of the electrons swirling around every atom and every bonded atom and every bonded compound. These are real but weak. The more electrons present, the stronger they are, but they never get as strong as ionic, covalent or metallic bonds.

Dipole attraction is found in SCl_2 (and many other polar molecules <u>without</u> hydrogen atoms). This is created by the near permanent dipoles (positive side/negative side) of polar bonds in polar molecules. All of these compounds have electron dispersion attractions AS WELL, but the electron dispersion attractions are minor compared to the dipole attractions.

Hydrogen bonding is found in water, ammonia, and many other polar molecules with polar bonds <u>that contain</u> hydrogen atoms. Water, ammonia (etc.) with hydrogen bonds have electron dispersion attractions AS WELL, but the electron dispersion attractions are minor compared to the hydrogen bonding.

Ionic Bonds are the strongest bonds. They form from cations and anions TRANSFERRING electrons. Additionally, the ionic compounds ALSO have electron dispersion attractions but these are very minor compared to the ionic bonds.

Covalent Bonds are very strong, but slightly weaker than ionic bonds. They form from nonmetals SHARING electrons. Additionally, the covalent compounds ALSO have electron dispersion attractions but they are very minor compared to the covalent bonds.

Metallic Bonds are strong enough to keep metals in the same shape forever, and we understand them to be packed cations, SHARING loose valence electrons. They ALSO have electron dispersion attractions but they are very minor compared to the metallic bonds.

Electron dispersion is found in all atoms and all compounds. It's biggest "impact" is seen in group 17, the halogens. Each is diatomic, with a nonpolar bond and these all have radial symmetry so they are nonpolar. These molecules have very little attraction to each other (nonpolar bonds, nonpolar molecules), but they show increasing numbers of electrons, all of which are moving about, creating "moments" of positive or negativeness in the electron orbitals due only to the random motion of the electrons. The more electrons, the more movement, and the more "positiveness" or "negativeness" that temporarily is created. These moments of polarity cause attractions that are weak, but increase with the number of electrons.

This group has two gases, one liquid and one solid. These phase differences are ONLY due to the numbers of electrons they have, and the ever increasing attraction that they generate.

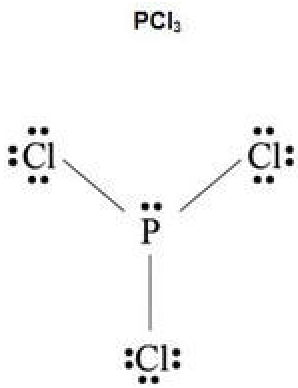
UNUSUAL BONDING, exceptions, weirdo bonds that are important to us as people, etc.

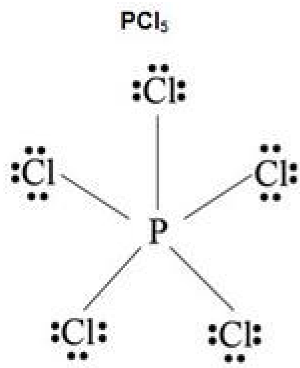
There are several bonds that we need to learn about that do not follow the "rules" of bonding, but somehow they exist, and they are worth looking at.

Rules include the octet rule, meaning only 8 electrons fit into the valence orbital unless it's too small. This rule is broken with the compound PCl_5

In this, phosphorous has 5 valence electrons, and they break apart, allowing 5 chlorine atoms to bond in. That gives the P atom 10 electrons. This is not normal, but possible.

PCl_3 and PCl_5 are shown here. They have 3 or 5 polar covalent bonds. In the first, there are one pair of UN-SHARED electrons shown at top. This molecule does not have radial symmetry, it's polar. So is PCl_5





A

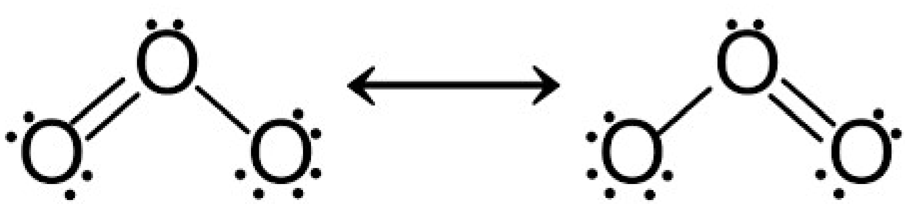
This diagram shows a combo design of structural with the nonbonding unshared pairs of electrons. At left, P has 2 electrons that are not involved in bonding, and paired for their stability.

The chlorine atoms all have 3 pair each of unshared electrons.

The P atom at right seems to have 10 electrons, IT DOES!

It breaks the octet rule and makes 5 single bonds.

As you might remember, I am from a place in Queens called Ozone Park. Ozone is a type of pure oxygen, but it's got a different formula, and different bonding. You can't stay alive if you breathe ozone, and pure O_2 won't protect you from harmful rays of the Sun. They're both pure oxygen but with different structures, different bonding, and different properties. These are ALLOTROPES. Ozone is O_3. But it won't bond in a stable way, the only way to keep it bonded is to make a single bond, and a double bond that switches back and forth, one side stable, the other nonstable. These bonds were described as resonating between the two structures shown below.



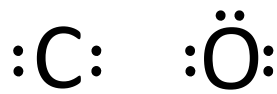
Ozone exists, and the bonding makes no sense, unless you "agree" that this can flip back and forth. In reality what really forms is a sort of 1½ bond on both sides rather than a double/single as shown. That also is hard to show, and the old resonance name stuck. The electrons bond in a stable way, but it's hard to draw.

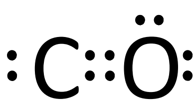
COORDINATE COVALENT BONDS

CO_2 makes two double polar covalent bonds. Since the molecule has radial symmetry (balance) the molecule is nonpolar. When CO, carbon monoxide forms, there is NO WAY that you can get the electrons to balance unless you know a tricky bond called coordinate covalent.

This is a common substance in your life, so you need to learn this, but it's not common bonding.

Carbon has 4 valence electrons, or 2 pairs. Oxygen has 3 pairs of valence electrons. There is possible combination where carbon shares 2 electrons with oxygen so oxygen gets an octet. This leaves carbon with just six, not an octet. SO, now oxygen will "loan" two of it's nonbonding electrons (the top 2 electrons) so carbon can have an octet as well. Since they are sharing just 2 pairs (double polar covalent bond) and oxygen "loans" one pair (a coordinate covalent bond) this is the only way both get an octet. It's also the only way to explain how these 2 atoms could bond. Start here:

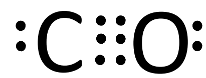




So what happens, strangely,
is that CO makes a double polar covalent bond like this at first,

This "satisfies the oxygen with an octet, but not carbon (and carbon is not happy without an octet). It's got to do more!

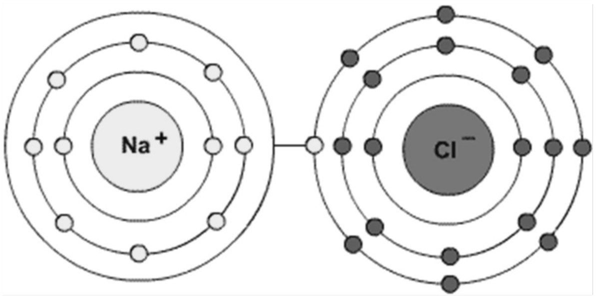
A coordinate covalent bond can be described as oxygen coordinating an octet for carbon by "lending" 2 of it's (top) unshared electrons to the double bond, making what LOOKS LIKE a triple bond, but is really a double polar covalent bond, PLUS a coordinate covalent bond. It looks like this:

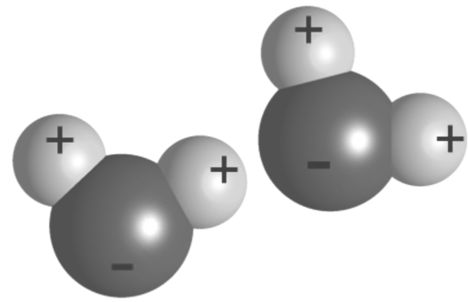


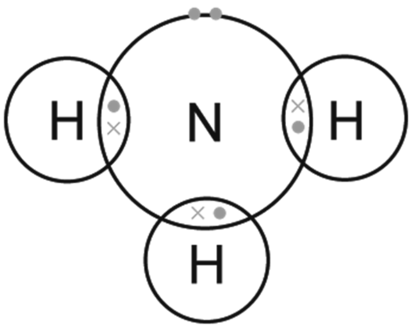
This "appears" to be a triple bond, but notice that carbon has 2 unshared electrons at left, and only has 2 more of it's electrons in the center. The extra electron carbon seems to "have" are from the oxygen atom.

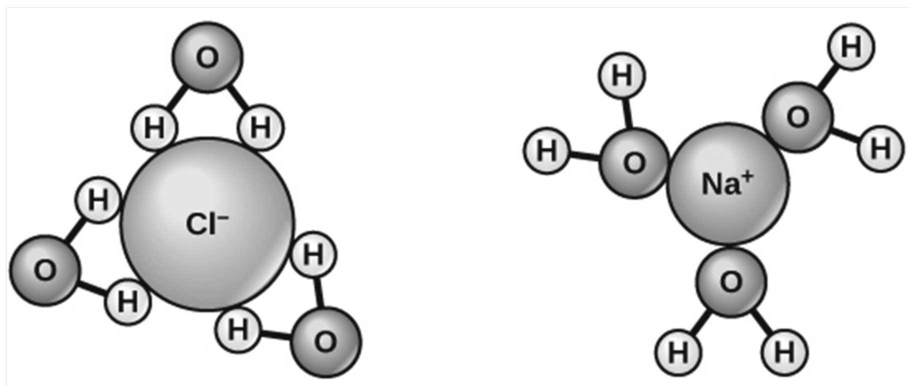
This is a double polar covalent bond PLUS two electrons just being put into the middle so carbon "feels" like it gets an octet too. Not normal, but CO exists in our life, we need to keep track of this bond.

There are many ways to show bonding, many are here. These are not used often, but you should be able to figure them out if you care to. NaCl forms when Na transfers an electron to Cl. Two water molecules with polar bonds (hydrogen are +). One H is attracted to the neighboring oxygen, via hydrogen bonding.

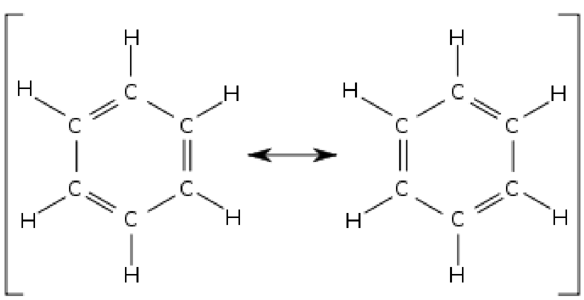






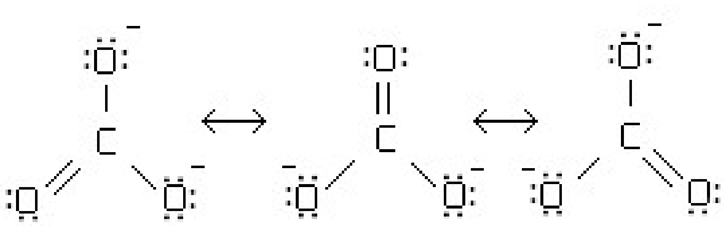


Ammonia forms when three H atoms bond to one N atom. The nitrogen has one unshared pair of electrons at top. Next is water molecules that surround sodium and chlorine ions ($NaCl_{(AQ)}$)
Note the orientation of the water molecules, their polarity "point" them at the ions in a particular way.



C_6H_6 resonates with 3 triple bonds and
3 single bonds between the carbon atoms.
Neither is more stable than the other,
so they will resonate back and forth.

At bottom left is the carbonate anion CO_3^{-1}
It too has no stable form, and the extra electron
making it charged moves about, shifting the
bonding to the oxygen atoms around.



Alloys

Alloys are mixtures of elements that contain at least one metal, often 2 or more metals. These elements are usually melted together, stirred up, and then let to chill to a solid. The new solid that forms is a mixture of the metals and together they have "better" properties than the original metals alone.

Common Alloys include…

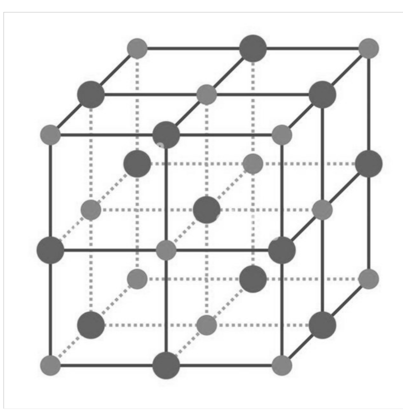
1	Carbon	and Iron	Form Cast iron for pipes that don't rust
2	Copper	and Zinc	Form Brass for trumpets and tubas
3	Silver	and Copper	Form Sterling Silver for table silverware and jewelry
4	Chromium	and Iron	Form Stainless steel for kitchen wear

Coordination numbers

When ionic solids form, the ions "fit" together in a pattern that is repeated in three dimensions. Although it's totally invisible at the ION level, once enough ions pack together, a shape emerges that is based upon the ionicn structure you can't see with your eyes.

The number of cations that surround each anion is called the coordination number. Oppositely, the number of anions that surround each cation is also a coordination number. These two coordination numbers are revealed to us when the number of ions is so many we can hold the crystals of ionic solid in our hands and see them with our eyes.

For us, it's a vocabulary word, but the physical shapes of different ionic crystals is based upon the coordination numbers that we can't see.



Sodium chloride (above) is boxy, the coordination numbers for both chlorine and sodium are SIX.

At right is CaF_2, and it's "cool" shape" is because the coordination numbers are not both the same.