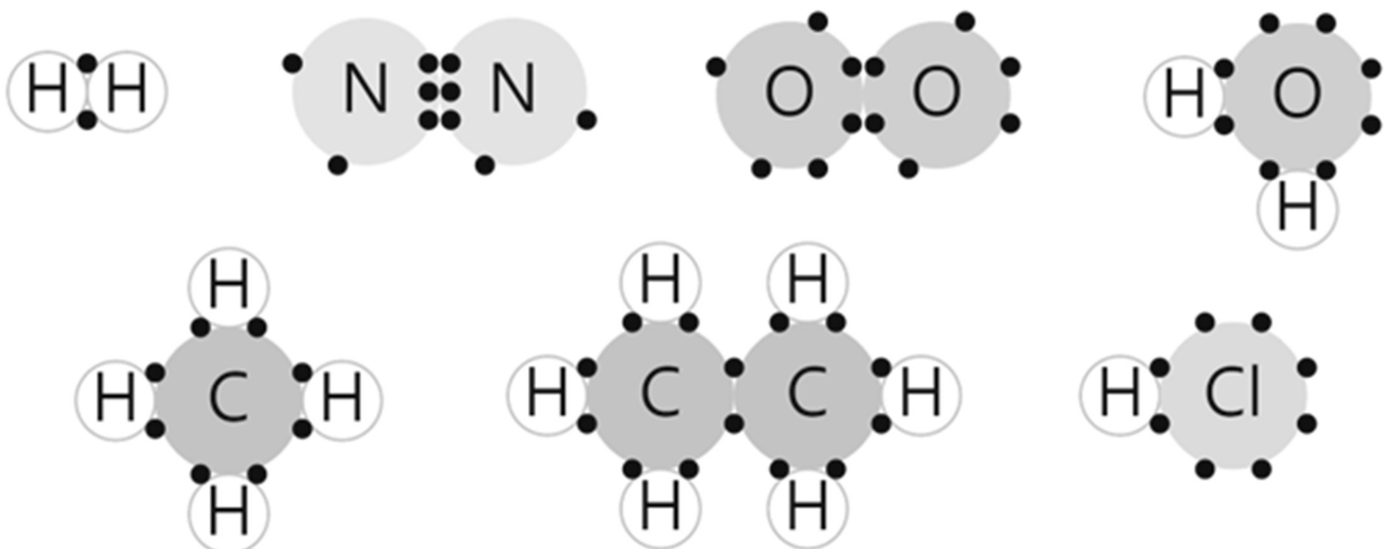
Name:

CHEMICAL BONDING PACK

BASICS & CLASS NOTES



ionic NaCl	covalent CO_2
metallic Ag	intermolecular $H_2O_{(L)}$ $Br_{2(L)}$

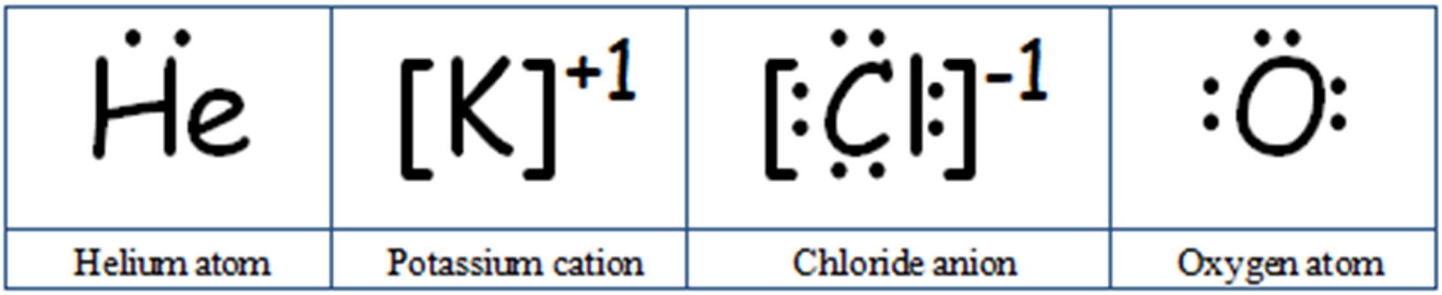
Bonding Basics

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, and 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.

<u>LEWIS DOT DIAGRAMS</u>

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.



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

The potassium cation has lost its outermost electron, and the whole valence shell at the same time. It becomes 2-8-8-0 It ends up with $19p^+$ and only $18\ e^-$, giving it 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 its third, or valence shell. It becomes a -1 anion with a 2-8-8 configuration. Anions get 8 dots.

Lewis dot diagrams for atoms show all valence electrons. Cations show the new, “empty” valence shell in brackets with a charge to show you KNOW what’s going on. Anions end up with FULL VALENCE SHELLS, which show 8 dots, and have brackets and charges too.

<u>IONIC BONDING</u>

This type of bonding is the simplest to understand for the new students. Whenever metals and nonmetals bond together, they first make ions. We say that metals lose electrons and form into positive CATIONS. The metal atoms do not really “lose” electrons, rather they TRANSFER electrons to the nonmetal atoms, which transform 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. There are never extra electrons, left over electrons, or any sort of “IOU” some electrons. The most common ionic compound is sodium chloride, table salt, NaCl.

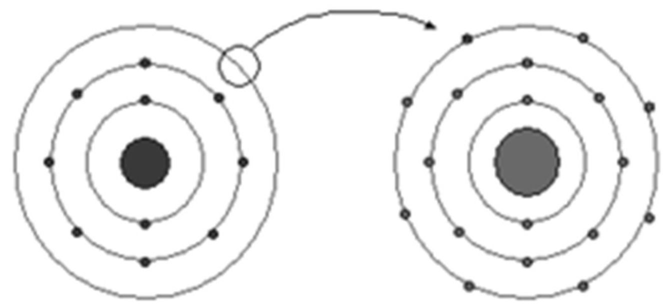
To draw atoms and ions (and compounds) in ways to help us understand this electron transfer we use LEWIS DOT DIAGRAMS. These show students the VALENCE ELECTRONS, which are the electrons in the outermost shell of the atom, *NOT* all of the electrons, just the outside electrons.

This is a model of the sodium cation that has "lost" an electron, and it's entire VALENCE SHELL.

The chloride anion has gained that electron, into its VALENCE SHELL.

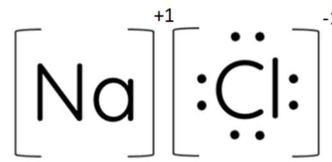
The opposite charges of the ions make the ions bond very tightly together, in what is known as an IONIC BOND.

The electron has been TRANSFERRED from Na → Cl

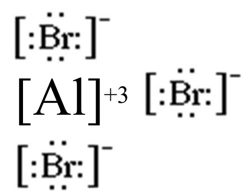


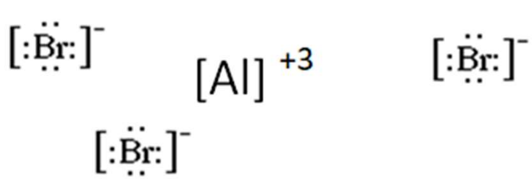
Na^{+1} cation Cl^{-1} anion

When ionic compounds form, we can draw Lewis dot diagrams for them. 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. Include the charges that sum up to zero. This is NaCl.



If the compound has more than two ions (say $CaCl_2$ or $AlBr_3$) just push the ion diagrams close. There's no "correct" way to draw them, just make them close together.





This arrangement at left DOES NOT show ionic bonding.

It is more like 4 dot diagrams that are near each other.

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

COVALENT BONDING

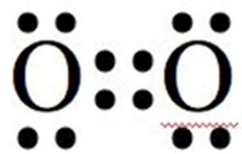
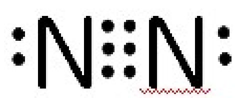
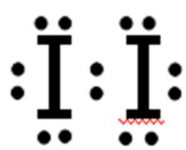
When metals and nonmetals bond, they form ions first, then are attracted together by opposite charge. When 2 or more nonmetals bond together they DO NOT FORM IONS. The atoms still try to become ISOELCTRIC to the noble gases, but they do not transfer electrons to do this.

Instead, when two atoms make a covalent bond they SHARE their valence electrons. By sharing, both atoms can share full shells.

This SHARING of ELECTRONS can be a perfect 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 SHELLS, usually that means 8 electrons, but only 2 electrons for the smallest atoms. Sometimes there are exceptions, which we will see too.

Water, carbon dioxide, and nitrogen gas all make types of covalent bonds. 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 shown 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 count electrons without making silly mistakes!		
		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. They only make one single covalent bond. These pairs of atoms have the same electronegativity values; these bonds are all nonpolar. Their "whole" name is: SINGLE NONPOLAR COVALENT bonds.

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

Nitrogen shares 3 pairs of electrons. The electronegativity difference is also zero. We call these types of bonds TRIPLE NONPOLAR COVALENT bonds.

IONIC bonds form from ions. COVALENT bonds occur between two or more nonmetals sharing electrons. Covalent Bonds can be POLAR or NONPOLAR bonds, AND be SINGLE, DOUBLE, or TRIPLE bonds.

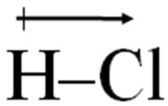
compounds	share this many pairs of electrons	share this many total electrons	name of the bond
F_2	1	2	single nonpolar covalent
H_2	1	2	single nonpolar covalent
O_2	2	4	double nonpolar covalent
N_2	3	6	triple nonpolar covalent
HCl	1	2	single polar covalent
NaCl	0	0	ionic
MgO	0	0	ionic
CH_4	1 pair, 4 times	8 (2 x 4 bonds)	4 single polar covalent

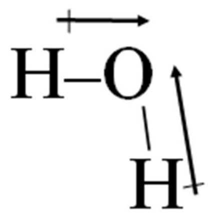
Electronegativity means tendency to gain electrons in a bond. The atom with a higher electronegativity “gets” electrons more of the time than the atom with the lower value. Atoms with higher electronegativity tend to be negative most of the time. The atom that gets the electron LESS of the time tends to be the more positive side of the bond. The POLAR BOND has a positive and a negative pole.

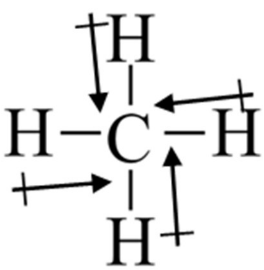
We can SHOW this polarity with a DIPOLE ARROW. The arrowhead shows which side of the bond gets the electrons “most of the time”. The tail of the arrow shows what side of the bond is “MORE POSITIVE” because some electrons moved away from this atom. A dipole arrow looks like this:

This side is “more positive”
The electron has <u>left</u> this side

This side is “more negative”
The electron <u>ends up here</u> most of the time







H (2.2) and Cl (3.2) make a very polar bond.

In water, the H (2.2) and oxygen (3.4) make the most polar bonds of these three molecules.

Methane, CH_4, has four single polar covalent bonds. Each bond is polar because of electronegativity value differences (C = 2.6 and H = 2.2)

Carbon gets the electrons MORE, leaving the H atoms “more positive” most of the time.

The dipole arrow “shows” this concept quickly.

In these examples, hydrogen has the LOWER electronegativity value, so the arrowheads point to the atoms with the higher electronegativity value. The hydrogen atoms in these three molecules always become the more positive side of the bond. The atoms do not share these electrons “fairly”.

Chlorine, Oxygen, and Carbon all pull the electrons harder than Hydrogen, they become “more negative”.

When atoms DO NOT SHARE ELECTRONS EVENLY, they make POLAR BONDS.

In water, the oxygen atoms make two separate bonds with each of the two hydrogen atoms, each bond is polar. Methane makes FOUR single polar covalent bonds.

All these bonds are SINGLE POLAR COVALENT, because the sharing of electrons is uneven.

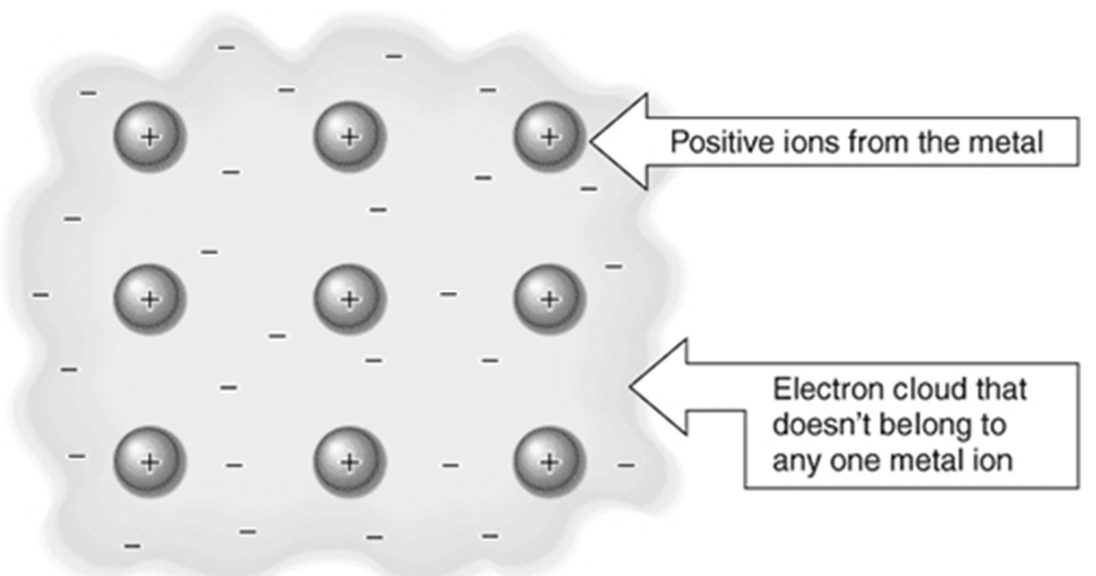
METALLIC BONDING

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 second, it will hold this shape indefinitely.

The metal conducts heat well. It also will conduct electricity. If you smash it with a hammer, the metal will bend but it won't crack. The reason scientists believe metals do these things, instead of not conducting electricity, or cracking like glass, is because of how scientists describe metallic bonding.

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 they would repel apart; they would crack. This DOES NOT HAPPEN, the metal is able to change shape, 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 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 really a solid wire.

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

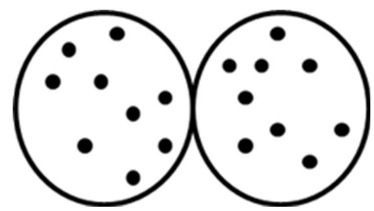
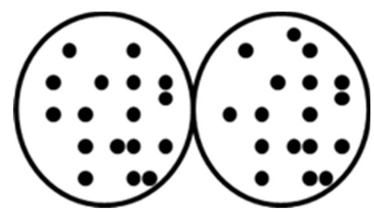
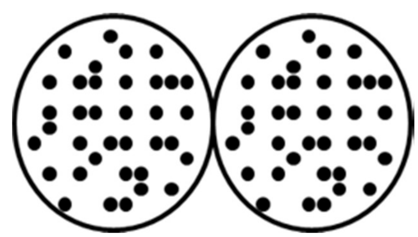
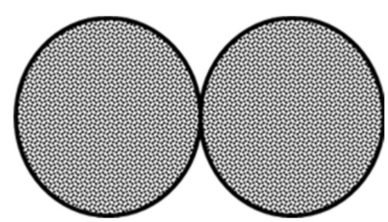
Malleable means a metal can be flattened when pounded.
Ductice means a metal can be squished into the shape of a wire (drawn out or pushed through a hole).
Conducting electricity means electrons easily flow through the solid metal.

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 one atom	Number of e^- in a molecule	diagram
F	F_2	9	18 Count them	
Cl	Cl_2	17	34 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 move very fast, and not in any exact orderly way like planets going around the Sun. The electrons are in shells, 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.

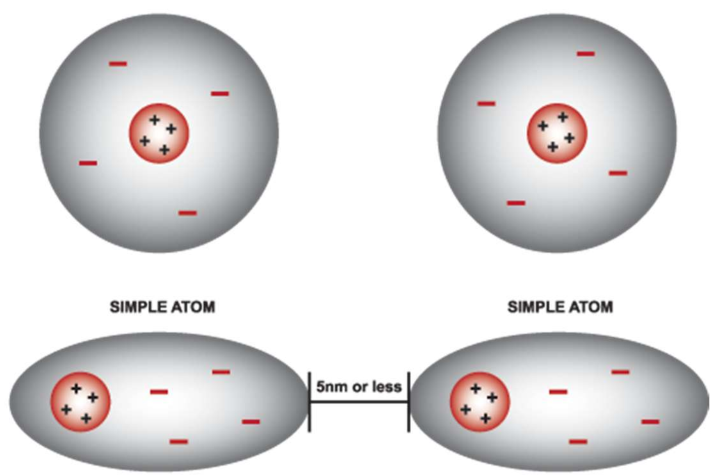
Over time, these instant points of negative or positive in the electron clouds have some attraction and some repulsions for other atoms or molecules. If you have few electrons, it's hard for this to amount to much positive or negative, and the force is terribly weak (real, but nearly insignificant).

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 electron 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 these 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 number of electrons dispersing. When that many electrons are moving about, they create more moments of positive and negative charge, 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.

<table>
<tr>
<td>The weakest of the
intermolecular molecular
attractions of force (IMF)
is Electron Dispersion force.

Sometimes it's called
London Dispersion Force,
or
van der Waal Force.</td>
<td>
</td>
</tr>
<tr>
<td colspan="2">The top two atoms have electrons flying around the nucleus, seemingly spread out evenly.
If the electrons get uneven, this makes moments of positive or negative in the electron clouds.
That creates temporary attractions, atoms or molecules do have attractions to each other.
These are short lived and weak, but they make a difference that adds up to the particles.</td>
</tr>
</table>

DIPOLE ATTRACTION

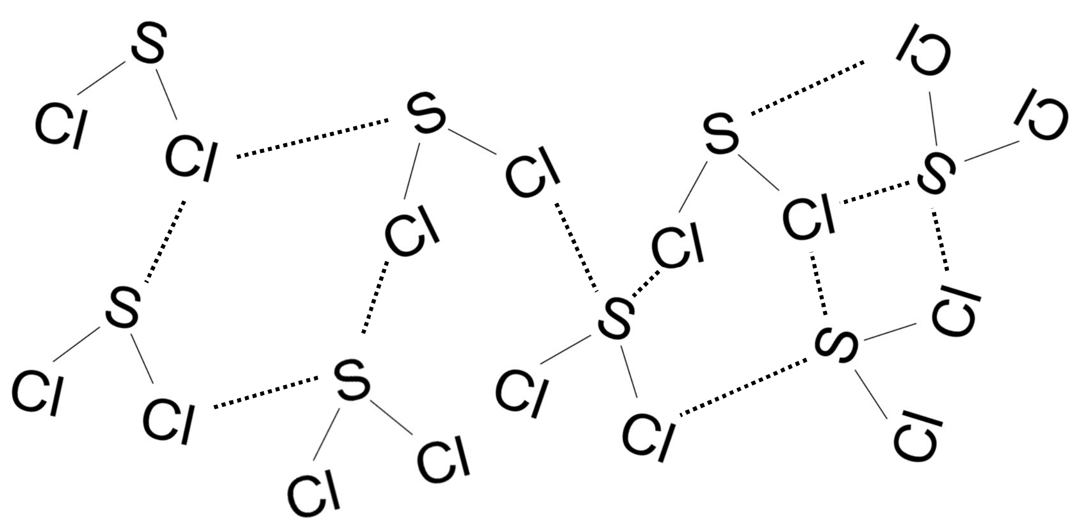
When atoms bond in covalent bonds they can share electrons perfectly together, if their electronegativity values are equal. Electronegativity means tendency to gain an electron in a bonding situation. When F_2 bonds, their electronegativity difference is zero. The bond is single NONPOLAR covalent.
The same happens with H_2, or Br_2. Both also 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, again NO DIFFERENCE in electronegativity.

When bonds form and the bond is POLAR, that's because there IS A DIFFERENCE in electronegativity.

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, for most of the time, near permanent polarity.

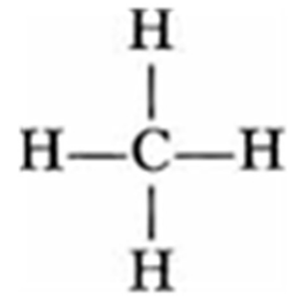
These polar covalent bonds are sharing electrons; they are just not sharing them equally. This unequal sharing makes the bonds uneven, or POLAR.

Inside a sulfur dichloride molecule, there are two single polar covalent bonds. The sulfur side is more positive, and the chlorine side is more negative. This polarity being almost permanent makes nearby molecules of SCl_2 attract each other. It is not a super strong positive or negative, but is nearly permanent, so the opposite poles of neighboring molecules will attract to each other's opposing poles. Chlorines (-) of one molecule are attracted to the sulfurs (+) of other molecules, this attraction is called DIPOLE ATTRACTION.



Above are many molecules of SCl_2, randomly dispersed. The sulfur atoms are more positively charged because they have a lower electronegativity value than chlorine. Chlorine atoms are more negatively charged. This ALMOST CONSTANT POLARITY creates the dipole attraction. Dotted lines = DIPOLE ATTRACTION

With molecules of methane, notice 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.

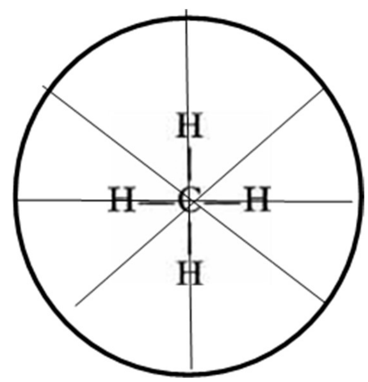


This molecule has RADIAL SYMMETRY. That is the same sort of balance as a pizza pie. No matter how you cut a pizza pie (or cherry pie, or circle), if you go through the middle point, you get two equal halves. NO MATTER WHAT.

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 inside. When you put a bunch of methane together, all the outside atoms of hydrogen are positively charged most of the time. There is NO DIPOLE ATTRACTION.

There are polar bonds and nonpolar bonds. There are ALSO polar molecule and nonpolar molecules. Dipole attraction requires polar bonds in polar molecules. Molecular shape determines molecular polarity.

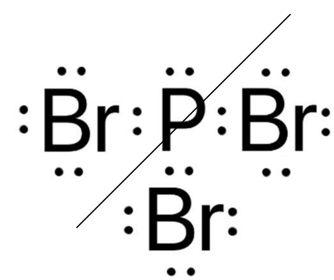
Radial symmetry is the only symmetry we care about in chemistry.
CH_4 has 4 polar bonds, but it has radial symmetry.
The molecule is nonpolar, it's balanced.
The distribution of the positive charges is symmetrical.



Humans (and gingerbread men) have bilateral symmetry.
So does a water molecule. These have just one plane of symmetry.

Radial symmetry is the only symmetry that matters in chemistry.
Water molecules are POLAR, the polarity of the bonds just don't cancel out.

In phosphorous tribromide, each of the 3 single bonds between P-Br are polar.
There are three bromine atoms bonded to phosphorous,
When you imagine cutting it this way, the left side has 2 Br atoms,
the right side has one Br atom plus 2 nonbonding electrons in phosphorous.
This molecule does NOT HAVE radial symmetry. It is a polar molecule.



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 electronegativity value.

For example:
SCl_2 has an electronegativity difference of 0.6 between S and Cl.
In water, H_2O, the difference between oxygen and hydrogen electronegativity is 1.2 (3.4 - 2.2) which creates a much greater bond polarity.
In ammonia (NH_3) the nitrogen - hydrogen electronegativity difference is 0.8 (3.0 - 2.2) which is greater than the 0.6 differential in SCl_2. These dipoles with hydrogen involved are so much stronger that we (unfortunately) give them a different name. This is sort of silly. Dipole attraction or super-duper dipole attraction would be fine with me, but not with the State Education Department!

Dipole attraction with hydrogen atoms involved is called HYDROGEN BONDING.

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

From weakest to strongest, the intermolecular bonding strengths are: electron dispersion force, dipole attraction, and hydrogen bonding

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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.

All atoms and all compounds have electron dispersion attractions. They are caused by the temporary "moments" of positive and negative charge created by the motion of the electrons swirling around every atom and in every 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 all polar molecules without hydrogen atoms). This is created by the near permanent dipoles (positive side/negative side) of polar bonds in polar molecules. All these compounds have electron dispersion attractions AS WELL, but the electron dispersion attractions are very weak compared to the dipole attractions.

Hydrogen bonding is found in water, ammonia, and other polar molecules with polar bonds that contain hydrogen atoms. Water and ammonia (etc.) has hydrogen bonding AND electron dispersion attractions too. The electron dispersion attractions are very minor compared to the hydrogen bonding.

Ionic Bonds are the strongest bonds. They form when cations TRANSFER electrons to anions.
These ionic compounds ALSO have electron dispersion attractions, but these are very minor compared to the strength of the ionic bonds.

Covalent Bonds are very strong, but slightly weaker than ionic bonds. They form when nonmetals SHARE electrons. 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; we understand them to be packed cations, SHARING loose valence electrons.

Electron dispersion is found in all atoms and all compounds. Its biggest “impact” is seen in group 17, the halogens. All are diatomic (F_2, Cl_2, Br_2, and I_2) and all have single nonpolar covalent bonds and with radial symmetry, the molecules are all nonpolar as well.

These molecules have very little attraction to each other (nonpolar bonds, nonpolar molecules), but they have increasing numbers of electrons, which create “moments” of positive or negativeness in the electron clouds. The more electrons, the more moments of charge and attractions between the molecules.

F_2 and Cl_2 are both gases, they don’t have that many electrons, so their electron dispersion attraction is relatively weak. Br_2 has enough electrons that make the electron dispersion attraction stronger, strong enough to make bromine a liquid at room temperature! You have seen I_2 in class; iodine has so many electrons and a much larger electron dispersion attraction, making iodine is a solid at room temperature.

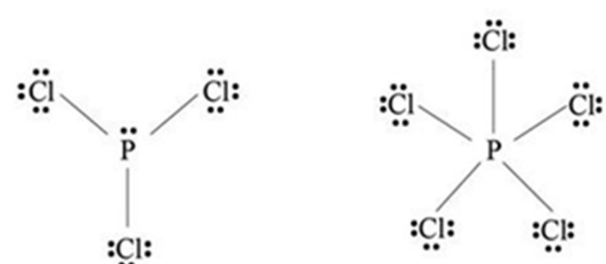
UNUSUAL BONDING exceptional, weirdo bonds that are important to know

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 shell unless it’s too small.
This rule is broken with the compound PCl_5 - phosphorous pentachloride.

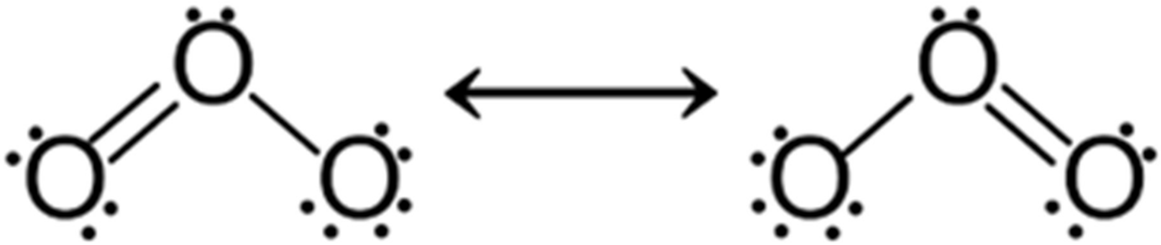
In PCl_5 phosphorous has 5 valence electrons, and they separate to allow 5 chlorine atoms to bond to it.
That gives the P atom a total of 10 electrons. This is not normal, but possible.

PCl_3 and PCl_5 are shown here. They have 3 single polar covalent bonds, or 5 single polar covalent bonds.
In PCl_3, there are one pair of UNSHARED electrons shown at top. This molecule does not have radial symmetry, it’s polar. Although it is too complicated to discuss here, PCl_5 is a nonpolar molecule.

This diagram shows modified Lewis Dot diagrams; the shared pairs of electrons (bonds) are indicated by dashes. The nonbonding unshared pairs of electrons are dots here. At left, P has 2 electrons at top called an unshared pair of electrons that are not involved in bonding. All chlorine atoms all have 3 pairs of unshared electrons. The P atom on the right seems to have 10 electrons. It does! PCl_5 breaks the octet rule, it makes 5 single bonds.	

Resonating 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 just ozone; pure O_2 won't protect you from harmful rays of the Sun. They're both only oxygen atoms, but with different structures, different bonding, and different properties. These are ALLOTROPES. Ozone is O_3. But it doesn't follow our simple bonding rules. It was originally described as making one single bond, and one double bond that was unstable and would "switch" back and forth, one side stable, the other nonstable. These bonds were described as resonating between the two bonding structures shown here.

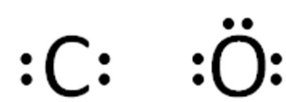


Ozone exists, and the bonding is confusing, unless you accept that these bonds flip back and forth – or resonate with each structure. What really forms is a sort of 1½ sized bond on both sides rather than a double/single as shown. That is hard to draw, and the old resonance name stuck. The ozone electrons do bond in a stable way, but it's hard to draw. Sulfur dioxide has resonating bonds as well.

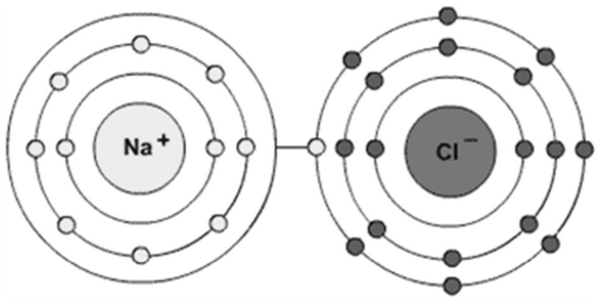
COORDINATE COVALENT BONDS

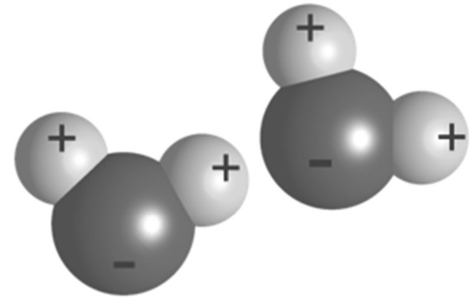
CO_2 makes two double polar covalent bonds. The molecule has radial symmetry (balance) - the molecule is nonpolar. When CO, carbon monoxide forms, there is NO WAY that you can get the electrons to bond unless you know a tricky bond called coordinate covalent bond.

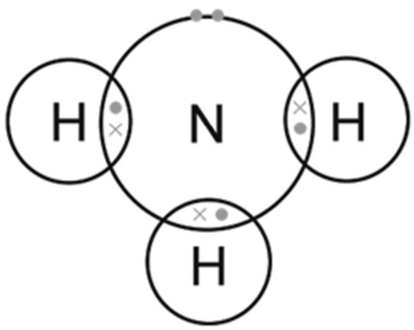
This is a common substance in your life, so you need to learn this, but it's not common bonding.
Carbon has 4 valence electrons (2 pairs). Oxygen has 6 valence electrons (3 pairs). Bonding chaos!

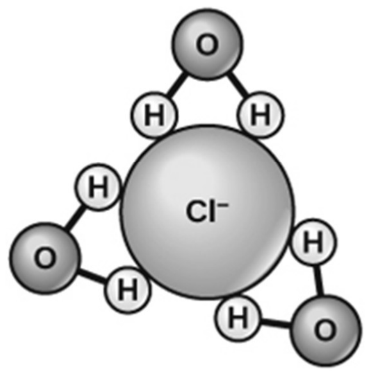
Start here with one carbon and one oxygen atom	
What seems to happen is that the CO molecule begins by making a double polar covalent bond like this. This "satisfies oxygen with an octet, but not carbon. (carbon is not happy without an octet).	
A coordinate covalent bond can be described as oxygen coordinating an octet for carbon by "lending" 2 of its (top) unshared electrons into the existing double bond. That makes this look like a triple bond but isn't! It is a double polar covalent bond, PLUS a coordinate covalent bond. It looks like this → The "top" pair of electrons are BOTH from oxygen, who has "lent" these to electrons into the middle to create two octets.	

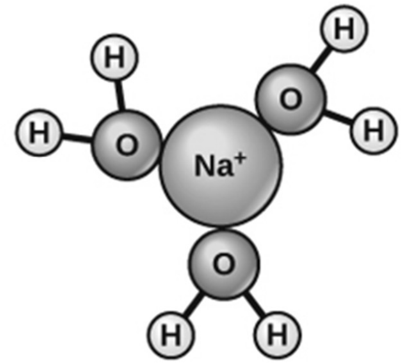
There are many ways to show bonding, many are here. These are not used often, but you should be able to figure them out if you care to.   NaCl forms when Na transfers an electron to Cl. Two water molecules with polar bonds (hydrogen are +). One H is attracted to the neighboring oxygen, via hydrogen bonding.



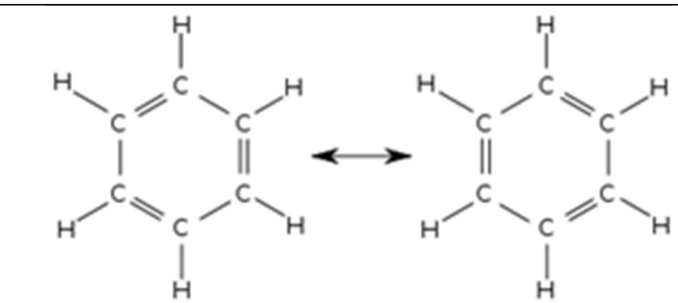


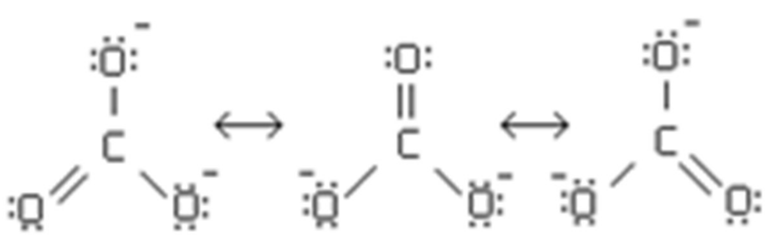






Ammonia forms when three H atoms bond to one N atom. The nitrogen has one unshared pair of electrons at top. Next is water molecules that surround sodium and chlorine ions ($NaCl_{(AQ)}$)
Note the orientation of the water molecules, their polarity "point" them at the ions in a particular way.

<table>
<tr><td></td><td rowspan="2">C_6H_6 resonates with 3 triple bonds and 3 single bonds between the six carbon atoms.
Neither structure is more stable than the other, the bonds resonate back and forth.

Here is a CO_3^{-2} carbonate anion. It has no stable form; the two extra electrons that make this ion have a -2 charge move about, shifting from one structure to another. This resonates as well.</td></tr>
<tr><td></td></tr>
</table>

Alloys

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

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

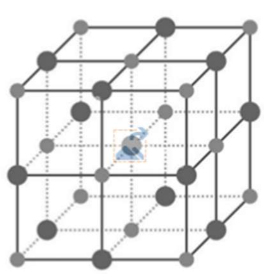
When ionic solids form, the ions "fit" together in a pattern that is repeated in three dimensions. Although it's totally invisible at the ION level, once enough ions pack together, a shape emerges that is based upon the ionic structure you can't see with your eyes.

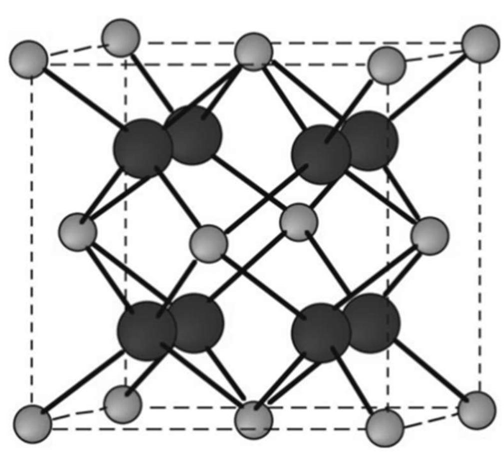
The number of cations that surround each anion is called the coordination number. Oppositely, the number of anions that surround each cation is also a coordination number. These two coordination numbers are revealed to us when the number of ions is so many we can hold the crystals of ionic solid in our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

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

Sodium chloride (left) is boxy, the coordination numbers for both chlorine and sodium are SIX. Each sodium cation is surrounded by six chloride anions. Each chloride anion is surrounded by six sodium cations. This makes for a nice regular cubic shape crystal.

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





Bonding Notes

	Type of bond	Important things to know about each type of bond
1		form between metal cations and nonmetal anions. Electrons are TRANSFERRED All ions end up with FULL electron shells (isoelectric to noble gases)
2		form between 2 or more NONMETALS - no metals ever. Electrons are SHARED. There are NO IONS. Bonding atoms usually share FULL outer shells.
3		form between atoms of metals. These hold solid metals together, no new compounds form. Can be pure, or mixtures of more than one metal (alloys)
4		fhese are relatively weak attractions <u>between</u> molecules (not like bonds inside of them). There are 3 different kinds of IMF.

5	The outermost electrons are the ________________________ electrons	
6	The outermost electron shell is the ____	
7	To bond, atoms or ions always* end up ____ to a noble gas; they get a full valence shell full of electrons * there are some exceptions.	
8	In a Lewis Dot Diagram	The DOTS represent the outermost electrons or _________________________________
9	In a Lewis Dot Diagram	The dots represent only valence electrons. We do not draw the __________________________________ electrons that don't bond

Electrons do not just fly around the nucleus; they are in specific zones that are of a certain size and shape. The shells can "hold" only a certain number of electrons each. If there are "too many electrons", another electron shell opens to accommodate these extra electrons. Let's look at the Group 18 electron configurations.		
10	Shell 1	The first shell is tiny; it only holds up to _____electrons.
11	Shell 2	The 2nd shell is bigger; it holds up to ________electrons.
12	Shell 3	The 3rd shell is a bit trickier. It can be "full" when it has ____ electrons (like Ar), but it can also stretch to hold up to _____ electrons (like Kr).
The larger shells (4th, 5th, 6th and 7th) are fancier.		
14	The ________________________________ shells can also be full with 8 or 18 electrons, or stretch to hold more electrons.	
15	Rule: ____ If you need to know how many electrons fit into any shell, look at the group 18 atoms, they have only full, stable, perfectly full shells.	

16 Atomic Number	Atomic Symbol	Lewis dot (atom)	Ion Symbol	Lew dot (ion)
1	H		H^{+1}	
2	He		X	X
3	Li			
4	Be			

Atomic Number	Atomic Symbol	Lewis dot (atom)	Ion Symbol	Lew dot (ion)
5				
6				
7				
8				
9				
10			X	X
11				
12				
13				

Atomic Number	Atomic Symbol	Lewis dot (atom)	Ion Symbol	Lew dot (ion)
14				
15				
16				
17				
18			X	X
19				
20				

17 Remember that…	Metals lose electrons to become isoelectric to a noble gas, Metal Cations have
18 Remember that	Nonmetals gain electrons to become isoelectric to a noble gas, Nonmetal Anions have

Bonding class #2 Objective: Metallic Bonds, More Lewis Dots, and the Octet Rule. Get your periodic tables open on your desk now please (not later)	
19	To bond ionically, a metal atom forms into a cation a nonmetal atom form into an anion by transferring electrons (1,2,3, or more electrons) The Na° atom → _________________ with a +1 charge The Cl° atom → ______________________, with a -1 charge

20. Draw the Lewis dot diagrams for the Na and Cl atoms, the Na^{+1} and Cl^{-1}ions, and NaCl compound

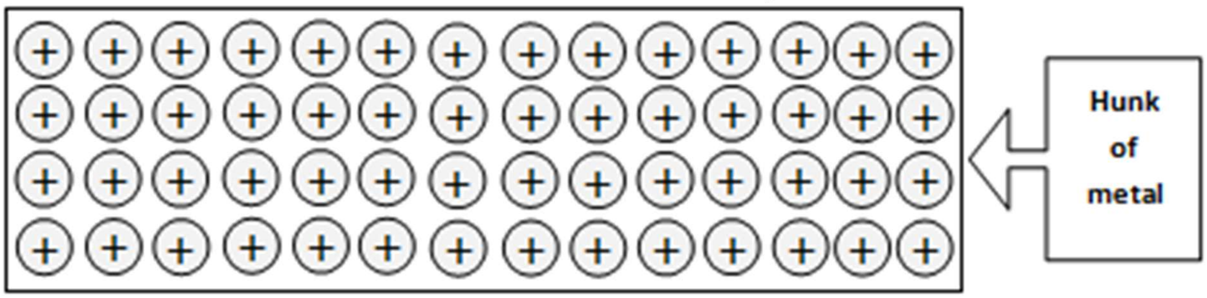
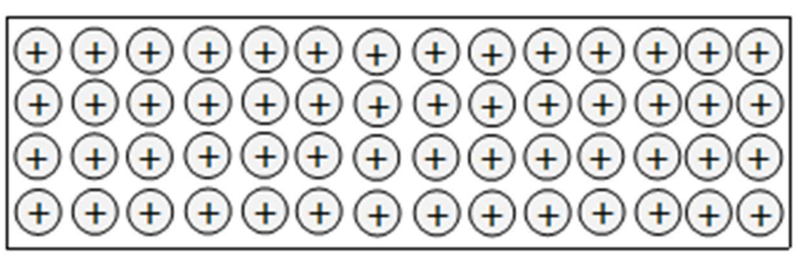
Atoms	Ions	Compound

21	Na atom loses 1 electron, changing the configuration from 2-8-1 to "2-8-0", which is really just 2-8 Na loses, or transfers, one electron, it becomes isoelectric to Ne ______________________________________
22	Cl atom gains one electron, changing the electron configuration from 2-8-7 to 2-8-8 Cl gains 1 electron, it becomes isoelectric to ________________________________

Noble gases have the most stable electron shells of all atoms.
They have only FULL, and PERFECT electron configurations.

23	Ions tend to follow the ____. The octet rule means most ions end up with 8 valence electrons (full outer shells) There are a few exceptions to the octet rule.
24	The exceptions include smaller ions with only ONE SHELL. They are too small for more than 2 electrons, like _______________________________________

25. Compound name	Compound Formula	Cation	Anion	Lewis Dot Diagram
Magnesium oxide	MgO	Mg^{+2}	O^{-2}	
	LiF			
	$CaCl_2$			
Sodium			S^{-2}	
Cesium oxide				

26	Metallic Properties that you should remember include…
27	Metals are understood to be…
27	Draw in the “loose valence electrons” as dots in between these “packed metal cations 
28	Metals are ___________________________, they have the same number of protons + electrons, but they seem to exist at packed cations, with shared valence electrons, not as packed atoms.
29	Imagine smashing the metal with a hammer to make the metal exhibit its malleable nature. The cations will be crushed closer together – they should repel from each other, causing a crack. The loose valence electrons ____. This also occurs when you squish metals into wires.
30	Ductile means can be____. Pulling creates a “stretch” not a snap.
31	Show the flow of electricity through the wire, from left to right. 

Objective: introduction to covalent bonding	
32	Covalent bonding occurs… when ____ share their valence electrons to make a bond.
33	Atoms ____ transfer electrons like ions do.
Review	
34	with ionic bonding from the METAL ___________________ → NONMETAL ______________________________ They follow the octet rule. The Metal comes first in the formula. Ionic compounds come in formula units (FU's), not molecules.
35	In Covalent Bonding, there is a SHARING OF ____
36	There are _____________________________________ in covalent bonds
37	Atoms bond covalently to form ________________________________
38	Molecules form using covalent bonds *(sharing electrons) and* follow the _____________________________________ almost every time.

Let's draw some molecules that share electrons covalently.			
	Formula	Dot Diagrams	Notes
39	F_2		
	Cl_2		
40	H_2		
	Br_2		

41	In covalent bonds, all atoms get to share enough electrons so that they get full valence shells ____.
42	All these bonds, F_2 - Cl_2 – H_2 - Br_2 are ____ bonds because they share _______________________________________- 1 electron from each atom AND because there is no difference between their ____________________________________ values

43	Electronegativity	Electronegativity	Electronegativity Difference	That means the bond is
F_2	F 4.0	F 4.0		
Cl_2	Cl 3.2	Cl 3.2		
H_2	H 2.2	H 2.2		
Br_2	Br 3.0	Br 3.0		

43	Draw HCl dot diagram	hydrogen monochloride and dihydrogen monoxide (water)
44	Draw H_2O dot diagram	
45	Redraw as structural diagrams	

46. Draw the Lewis Dot and the Structural Diagram for ammonia NH_3. Name each bond. First, AMMONIA is not AMMONIUM. Ammonia, NH_3 is a molecule. Ammonium, NH_4^{+1} is a cation. They are NOT THE SAME THING!		
Dots	Structural	Notes

47. Draw the Lewis Dot and the Structural Diagram for methane CH_4. Name each bond.		
Dots	Structural	Notes

The greater the difference in electronegativity values between two atoms, the greater the polarity of the bond. Greater electronegativity differences = MORE POLAR bonds. Less difference = LESS POLAR bonds.

48. Which bonds are more polar, the C-H bonds in methane or the N-H bonds in ammonia?

Bond	Electronegativity of 1st atom	Electronegativity of H	Electronegativity difference	Both bonds exhibit polarity
C-H	2.6	2.2		
N-H	3.0	2.2		

49. Draw the Lewis Dot and the Structural Diagram for oxygen O_2. Name each bond		
Dots	Structural	Notes

Bonding Class #4 Your objective today is to become masterful with the Double covalent bonds and Triple Covalent Bonds.

50. The HONClBrIF Twins bonding			
Twin	Dot diagram	Structural diagrams	Name the bond
H_2	H:H	H−H	Single polar covalent
O_2			
Cl_2			
Br_2			
I_2			
F_2			

51. What about Nitrogen, N_2?			
N atom	2-5 electron configuration	Needs to make 3 bonds, to get an octet	Each N atom needs to borrow 3 electrons from the other N atom
To make this work, you must become Master of the Universe, move the electrons, to where they belong.			
N_2	Dot diagram	Structural diagram	Name the bond

52. C_2H_6 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
Carbon makes an octet by borrowing 4 electrons, 3 from the H atoms, one from the other carbon atom. H get "full" baby sized shells	C–C bond is single nonpolar covalent C–H bonds are single polar covalent 7 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

53. C_2H_4 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
Carbon makes an octet by borrowing 4 electrons, 2 from the H atoms, 2 from the "other" carbon atom. H get a "full" baby sized shells	C=C bond is double nonpolar covalent C–H bonds are single polar covalent 5 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

54. C_2H_2 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
Carbon makes an octet by borrowing 4 electrons, 1 from the H atom, 3 from the "other" carbon atom. H get a "full" baby sized shells	C≡C bond is triple nonpolar covalent C–H bonds are single polar covalent 3 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

55. C_3H_8 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
Carbons make octets by borrowing electrons, from the H atoms, and from the "other" carbon atoms. H get a "full" baby sized shell.	C–C bonds are single nonpolar covalent C–H bonds are single polar covalent 10 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

56. CO_2 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
Carbon make an octet by borrowing 2 electrons from <u>each</u> oxygen atom. Both "O" atoms borrow 2 electrons from the carbon atom.	O=C bonds are double polar covalent THIS IS A STRAIGHT LINE 2 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

57. $AsCl_3$ Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
Arsenic make an octet by borrowing 1 electrons from each chlorine atom. The Cl atoms borrow 1 electron each from the arsenic atom.	As–Cl bond is single polar covalent 3 total bonds in molecule	Just one cut shows that this molecule does not have radial symmetry. One side is a "double Cl", the other is a single Cl plus a pair of electrons. This is a POLAR molecule.

58. C_4H_{10} Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
At some point you either can count electrons or you can't. If you want, you can draw as many dots as you like, or you can just "know" the dots (the electrons) are there or not there. You know and can know because you're smart. I am not drawing dots anymore. I think you are smart enough to know this.		
☺	C–C bonds are single nonpolar covalent C–H bonds are single polar covalent 13 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

59. OBr_2 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
All the atoms make octets. "O" borrows 2 e^-, one from each Br Each bromine borrows an electron from the oxygen.	O–Br bonds are single polar covalent 2 total bonds in molecule	The red line cuts this molecule into 2 equal halves, but the black line shows a double Br side vs. no Br side This molecule does not have radial symmetry, it's POLAR.

60. CCl_4 Draw the Lewis dot diagram, the structural diagram, then REDRAW the structural diagram to determine if this molecule has radial symmetry and is nonpolar, or if it's "unbalanced" and polar.		
All four atoms make octets. Carbon borrows 4 electrons, one from each chlorine, while each chlorine borrows 1 electron from the carbon.	C–Cl bonds are single polar covalent 4 total bonds in molecule	Every cut through the center point results in 2 equal halves, this molecule has radial symmetry, it is NONPOLAR

61. Draw Calcium oxide (last one today). Be careful, it's tricky! Calcium is atom #20, oxygen is #8.

Bonding Class #5: Reviewing some random bonding vocabulary, look at more Lewis Dot diagrams and Structural diagrams, and meet the weird hybrid bonds of O_3, CO, and PCl_5	
62	Alloys are ____ They are not ____ The metals get mixed together by melting. When cooled, the mixed metals are not bonded, but mixed, they have different physical properties. They can be melted apart (reverse physical change).
63	Examples of ALLOYS

Sterling Silver		for strength
		for strength & non-corrosiveness
		for strength & non-corrosiveness
		for durability, beauty, & French horns J

64	Coordination number is Each Na^{+1} cation is surrounded by 6 Cl^{-1} ions. The reverse is also true: Each Cl^{-1} ion is surrounded by 6 Na^{+1} cations
65	Since both coordination numbers are 6, NaCl has a ________________________________
66	Different Coordination Numbers form into different ________________________________.

67. Carbon Monoxide gas, CO First draw both ATOM Lewis Dot diagrams, then we discuss.

Carbon atom +	Oxygen atom →	Start with double polar covalent bond ↓
CO has is a double polar covalent bond, and a		Finish with the proper CO bonding here

67. A shorthand for this bond, a double polar covalent plus a coordinate covalent bond looks like this →

68. Phosphorous pentachloride   P has ____ valence electrons,   Cl has ____ valence electrons

This breaks the octet rule — phosphorous ends up with 10 valence electrons (two too many!)
Draw the structural diagram

69. Draw the Boron Trifluoride dot diagram and structural diagram

70. Oxygen & Ozone are both pure elemental oxygen but are bonded differently. O_2 and O_3 are

71. Draw Ozone with "resonating bonds"	72. Draw Ozone with "hybrid bonds"

Bonding Class #6 Intro to the Simple Intermolecular Attractions (IMF) (not LMAO, not LOL, not DULL)

73. Intermolecular Bonds are the:

They are much weaker than ionic or covalent bonds, or metallic bonds.
They do impact a substance's melting and boiling points (etc.).
At room temperature, substances with weak IMF are usually gas.
Stronger IMF means stronger intermolecular attractions, and usually results in a liquid or even solid.

74. There are 3 kinds of intermolecular bonds, from weakest → strongest they are	Weakest	Medium	Strongest

75. Electron Dispersion Attraction is clearly displayed by the halogens of group 17.

F_2, Cl_2 Br_2, and I_2, are all ______________________,

They have ____ bonds,

and they are all _____________________________________ molecules.

76. Fluorine F_2 Draw 3 molecules, show dipoles.	Each F_2 molecules has two 2-7 electron configurations, a total of 18 electrons. When these 18 electrons all "move" to one side, for a nanosecond, a temporary dipole is created, positive and negative, creating weak temporary attractions. These many electrons make very weak dipoles. F_2 is a gas at STP, The kinetic energy exceeds the attractive force of the electron dispersion attraction.

Chlorine Cl_2 Draw 3 chlorine molecules, show dipoles	Cl atoms have a 2-8-7 electron configuration Cl_2 molecules have 34 electrons. When these electrons all "move" to one side, for nanoseconds, they make mildly stronger temporary dipoles, compared to F_2. With so few electrons moving, the electron dispersion attractions are still very weak. Cl_2 is a gas at STP, because the kinetic energy exceeds the attractive forces.
Bromine Br_2 Draw 3 bromine molecules, show dipoles	Br atoms have a 2-8-18-7 electron configuration Br_2 molecules have 70 electrons! These electrons all shift around so much, when a temporary dipole is created, it's much stronger than in F_2 or Cl_2. The electron dispersion force is strong enough to make Br_2 a liquid! The weak <u>but constant</u> intermolecular attraction accumulates. This impacts the phase. Br_2 is a liquid at STP Because the intermolecular attraction strength exceeds the kinetic energy. The molecules can't shake apart into a gas.
Iodine I_2 Draw 3 iodine molecules, show dipoles	Each I_2 molecule has 106 electrons. The electrons shift around so much that so many temporary dipoles form that molecules of I_2 have the most moments of attraction of all the halogens. So much electron dispersion attraction exists in iodine molecules, I_2 is a solid at STP. Electron dispersion attraction is weak, but enough of it can affect the phase of these similar, nonpolar molecules like this.

77. Phases at STP	Halogens	The difference in phase is only due to the amount of …
Gases		
Liquid		
Solid		more electrons = stronger IMF

78. A dipole occurs when a molecule has polar bonds and doesn't have radial symmetry. These form a near permanent dipoles. Draw one SCl_2 molecule, with dipole arrows.	79. Draw 5 random SCl_2 molecules, SHOW dipole attractions IN COLOR

80. Draw the CH_4 molecule, with dipole arrows. CH_4 is a nonpolar molecule, it does not have dipole attractions,

These form a near permanent dipoles. Draw one SCl_2 molecule with dipole arrows.

81. Compare and contrast SCl_2 and CH_4

	Boiling Point (K)	Phase at room temp.
SCl_2		
CH_4		

Why would these 2 small molecules have different phase at room temp?

82 Hydrogen bonding is the same as dipole attraction, but ______________________________ must be present in the molecule. (It's like super-duper dipole attraction)

compound	atom 1 electronegativity	atom 2 electronegativity	Electronegativity difference	How polar are these molecules?
H_2O	H 2.2	O 3.4		
SCl_2	S 2.6	Cl 3.2		

83. Draw these 2 dipole molecules with dipole arrows	
SCl_2	H_2O
SCl_2 has…	H_2O has…

84. Draw 6 water molecules, the <u>use a color pencil to show the hydrogen bonding</u>

85. Bond types	example compound formulas (memorize this page)
Ionic	
Single nonpolar covalent	
Single polar covalent	
Double nonpolar covalent	
Double polar covalent	
Triple nonpolar covalent	
Triple polar covalent	
Coordinate covalent	
Resonant	
Ionic + Covalent at the same time	
Breaks the octet rule (more than $8e^-$)	
Breaks the octet rule (less than $8e^-$)	

Bonding Class # 7 Objective - master relative oxidation numbers, review bonding

86. ____ are those positive and negative integers in the corners of the NONMETALS on the periodic table. They help us determine what ratios of atoms to atoms that molecular compounds make in compounds.

87. Using a T-chart, and oxidation numbers, show what compounds can form from H and O atoms bonding.

88. What are the relative oxidation numbers for	
BF_3	
SiO_2	
HF	
PCl_5	

ex	Sulfur dioxide	SO_2	S^{+4} O^{-2} O^{-2} = sums to zero
ex	Chromate ion	CrO_4^{-2}	Cr^{+6} O^{-2} O^{-2} O^{-2} O^{-2} = sums to -2
89	Permanganate anion		
	Ammonia (not ammonium)		
	Sodium hydroxide		
	Potassium chlorate		
	Carbon monoxide		
	Carbon dioxide		
	Dihydrogen sulfate		
	Nitrate anion		
	Nitrogen dioxide		
	Phosphorus trichloride		

90. Explain the difference between bond polarity and molecular polarity.

91. Explain the resonating bonds of ozone.

92. Draw the CO, carbon monoxide molecule properly (dots and structurally).
Name the <u>bond or bonds, use color pencils or else!</u>

93	Question	True or False?
A	Ionic bonds can form single, double or triple bonds.	T F
B	Covalent bonds are always polar.	T F
C	Oxygen molecules have one double polar covalent bond.	T F
D	Nitrogen molecules have one double *nonpolar covalent bonds.*	T F
E	Hydrogen atoms can make single or double bonds.	T F
F	Nonpolar molecules can't have polar bonding.	T F
G	Water can *sometimes form into a straight–line shape.*	T F
H	Oxygen atoms must make double bonds.	T F
I	Molecules with only nonpolar bonds cannot form polar molecules	T F
J	The weakest intermolecular bond is called dipole attraction.	T F
There are two more slides, for thinking…		

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