Bonding Basics

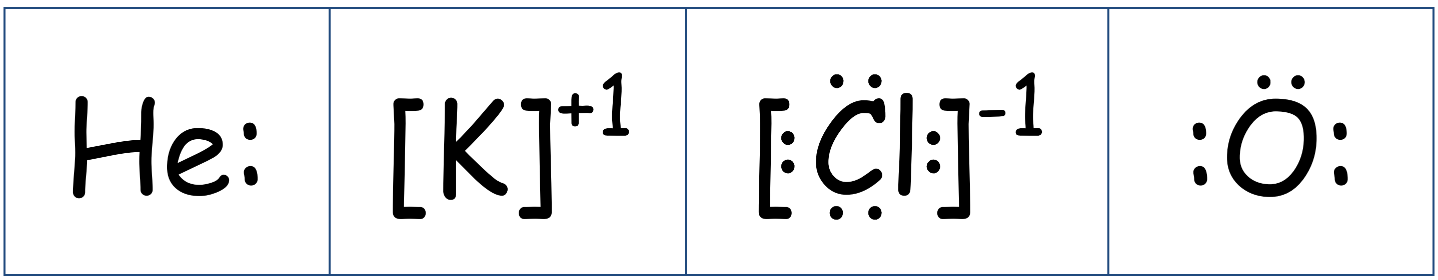
Chemistry is the study of the stuff of the universe, and importantly, how it forms into new substances by bonding in certain ways (or un-bonding to become simpler). There are several ways atoms can bond together in high school chemistry. We will look over each type, learning the particular ways they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

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 below. Which species (what are the actual specific names for each?) are each of these?

Atoms like helium, and oxygen show all valence electrons, and electrons will 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 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 8 valence electrons, all drawn in 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.

PRACTICE QUESTIONS Set 1 (answers on the last page)

1. What are the electron configurations for phosphorous and for magnesium?
   What are their ion electron configurations?
2. Which electrons are drawn in Lewis dot diagrams?
   A. all of them B. inner electrons C. outermost electrons D. no electrons
3. Draw Lewis Dot Diagrams for the lithium atom, boron atom, calcium cation, and the oxide anion.
4. Will an atom of aluminum ever normally lose 1 or 4 electrons? Why not?

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. They TRANSFER these valence electrons to nonmetals, which form into negative ANIONS. This transfer of electrons is always perfect", the amount of electrons lost by the metals are picked up by the nonmetals. No extra electrons, or left over electrons are allowed ever. 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!)

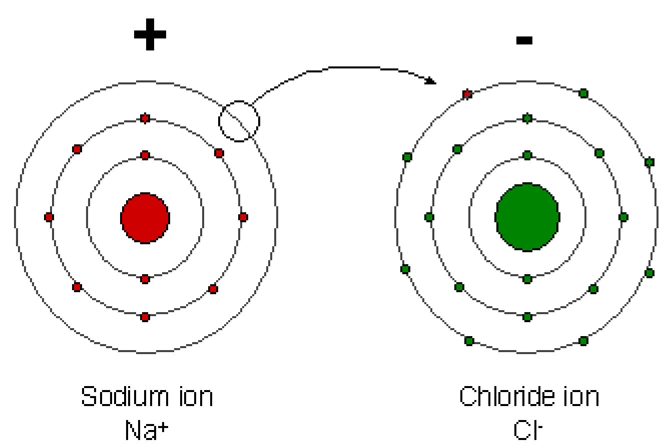
An atom of Aluminum has a 2-8-3 electron configuration. It must "lose" 3 electrons to become isoelectric to neon. $Al \rightarrow Al^{+3}$ with a 2-8 cation electron configuration.

Nonmetals, like chlorine, gain electrons to form into negative anions. Chlorine atoms start with 17 electrons in a 2-8-7 configuration. When $Cl \rightarrow Cl^{-1}$, it's electron 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, & 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.

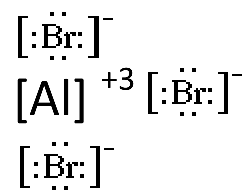
To draw atoms and ions (and compounds) in ways to help us understand this electron transfer we use LEWIS DOT DIAGRAMS. These show the VALENCE ELECTRONS only, which are the electrons in the outermost orbital of the atom, NOT all of the electrons, just the outside electrons.

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.

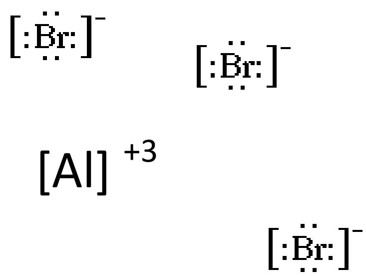


When ionic compounds form, we can draw Lewis dot diagrams for them as well. They are not exactly "pretty" but they are obvious. Just push the cation diagram close to the anion diagram together to indicate that they are making a compound. This is KCl...

If the compound has more than two ions (say $CaCl_2$ or even $AlBr_3$) 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. Different texts interchange some of these minor points.



This arrangement at left <u>DOES NOT</u> show ionic bonding, it's more like 4 Lewis dot diagrams that are near each other. Ionic bonding diagrams are CLOSE together, like the two above.

PRACTICE QUESTIONS Set 2

1. Draw the Lewis dot diagrams for Li^{+1}, P^{-3}, and for Neon.
2. Draw Lewis Dot diagrams for magnesium oxide and for potassium nitride.
3. Define ISOELECTRIC.

<u>**COVALENT BONDING**</u>

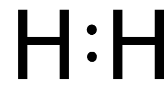
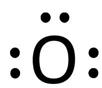
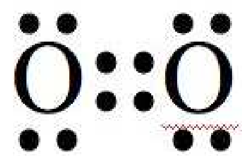
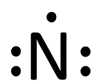
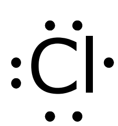
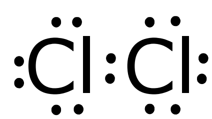
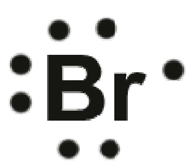
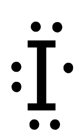
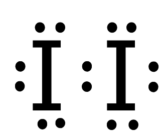
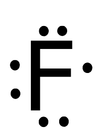
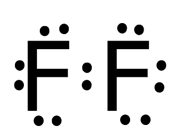
When metals and nonmetals bond, they form ions first, then are attracted together by opposite charge. When 2 or more nonmetals bond together (no metals allowed, ever), they DO NOT FORM IONS. The atoms still try to become ISOELCTRIC to the noble gases, but they do not transfer electrons to do this. Instead, when two atoms make a co-valent 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 will start with the HONClBrIF twins, they all show covalent bonding. The Lewis dot diagrams for them are here as atoms, and then as molecules. The simplest are H, Cl, Br, I and F, all making the smallest bonds.

		H–H
		O=O
		N≡N
		Cl–Cl
		Br–Br
		I–I
		F–F

The HONClBrIF twin molecules exhibit a variety of covalent bonds. H2, Cl2, Br2, I2 and F2 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 must share two pairs of electrons. The electronegativity difference is also zero, so these are DOUBLE NONPOLAR COVALENT bonds.

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

Bonds can be IONIC when formed from ions.

Bonds can be 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 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 for each of the 4 bonds	4 single polar covalent

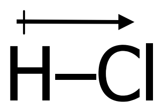
Electronegativity means tendency to gain electrons in a bond. The higher electronegativity "gets" electrons more of the time than the lower value. So, atoms with higher electronegativity will tend to get the electron and "be" negative more of the time. The atom that gets the electron LESS of the time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

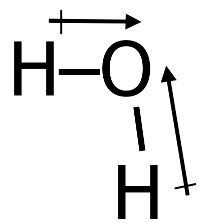
We can SHOW this polarity with a DIPOLE ARROW. This arrow shows the DIRECTION that the electrons go - to what side of the bond. The arrow ALSO shows what side is left "MORE POSITIVE" because the electrons moved to the other side of the bond. A dipole arrow looks like this:

This side is more positive — This side is more negative

The electron left this side — The electron ends up over here most of the time

A couple of examples of molecules with dipole arrows are:





In both 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 two atoms DO NOT SHARE EVENLY. In water, the oxygen makes two different bonds with the two hydrogen atoms, each one is sharing unevenly.

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.

TRICKY Questions

A common trick on the regents is to catch you on ionic bonds and sharing electrons. IONIC BONDS are about transferring electrons - NOT SHARING. Trick questions like:

A. In sodium chloride, NaCl, how many pairs of electrons are being shared? **(none!)**

B. In magnesium oxide, do these ions share two pairs of electrons? **(no, they don't share!)**

C. In calcium chloride, $CaCl_2$, are there two single polar bonds, or two single nonpolar bonds?) **(neither! The bonds are IONIC. Ionic bonds CAN'T BE single, double or triple either!)**

E. In sodium hydroxide, NaOH, are the bonds ionic or covalent? **BOTH! The Na^{+1} bonds ionically to the OH^{-1}, but INSIDE THE hydroxide, the O-H bond is polar covalent!**

F. How many electron pairs are shared between aluminum phosphide in AlP? **(NONE! It's ionic and there are no electrons shared, they are transferred!)**

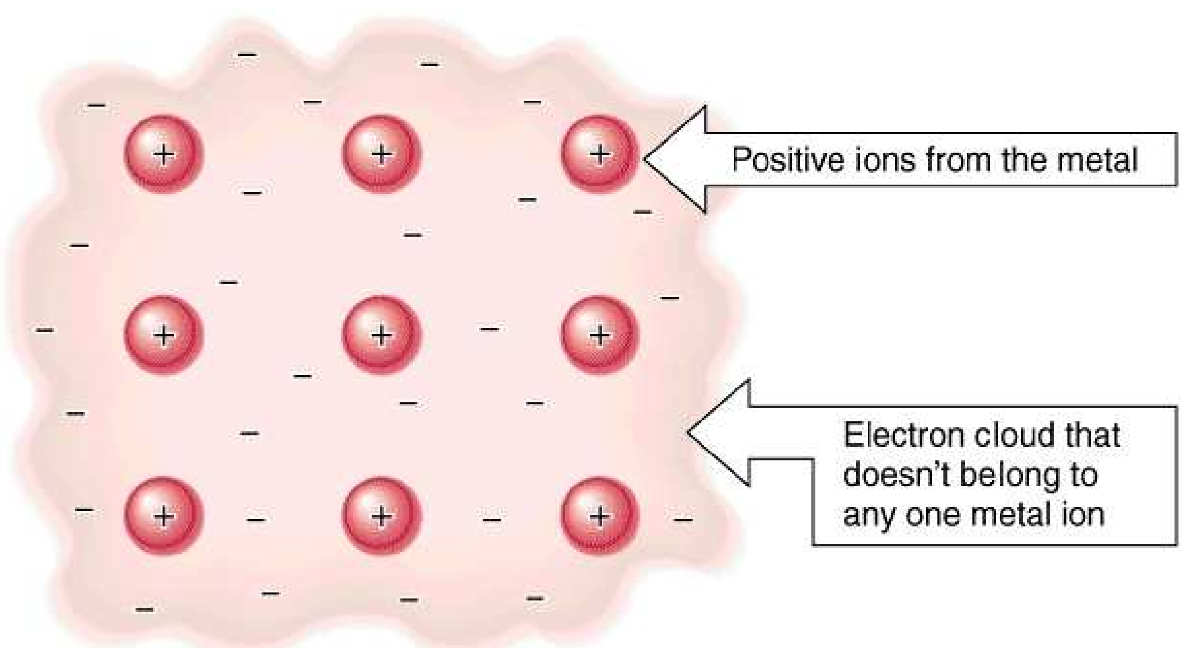
PRACTICE Questions set 3

1. Draw larger sized structural diagrams for KCl, for boron trifluoride, carbon tetra bromide.
2. Carbon dioxide makes two double bonds this way: O=C=O in a straight line. Name the type of bond between one atom of oxygen and one atom of carbon.
3. Name the bond types in oxygen difluoride.
4. Name the bond types in ammonium chloride.

METALLIC BONDING

When a frying pan is made, or if you take one out of your kitchen drawer, that hunk of metal displays many interesting properties. It's been stuck in that position for a long time, and it will likely hold it's shape indefinitely. It conducts heat well. It will also conduct electricity (as do all metals). It's been smashed into a shape and it didn't crack. If you bash it on a big rock, it might bend but it won't shatter like glass. The reason scientists believe metals do these things, instead of not conducting electricity, or not being malleable, is because of how they describe how metals bond together.

Metals are described not as packed atoms, rather as PACKED CATIONS with their valence electrons LOOSE and (literally) mobile in the metal. Like this picture.



If the metal cations are crushed together, say by banging the pan on a rock, the positive cations would get closer, and would want to repel, even to 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 closer 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.

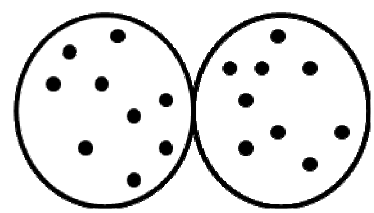
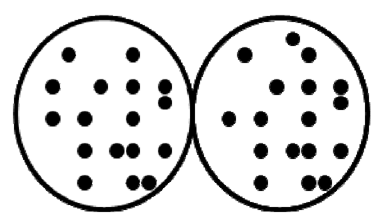
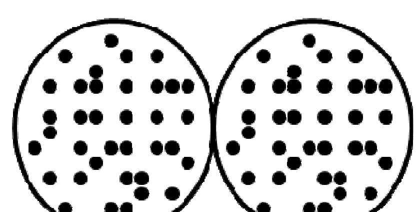
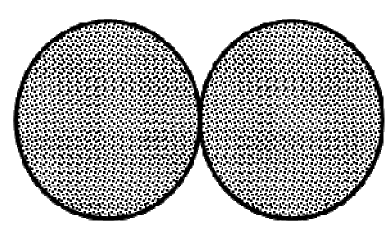
Loose valence electrons, in a packed cation solid, explain most of the metallic properties.

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	
Cl	Cl_2	17	34	
Br	Br_2	35	70	
I	I_2	53	106	

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 instantaneous positive and negative spots on their orbital clouds, this force even with 34 electrons in Cl2 cannot overcome the kinetic energy of STP. These two elements are gases at normal temperature and pressure.

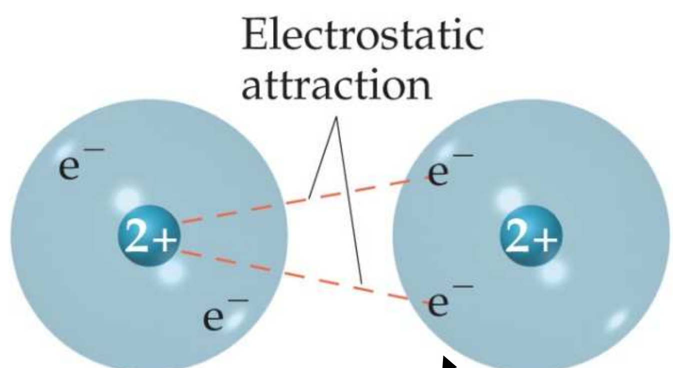
Bromine, Br_2, has 70 total electrons, and that many electrons dispersing at any point in time make MORE points of temporary positive and negative than fluorine or chlorine. Br_2 at STP is a liquid, because the electron dispersion attraction with this many electrons is enough attraction to hold the molecules 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.

This is sometimes called London Dispersion Force, but that's a bit old fashioned.

All atoms, and all compounds have electron dispersion attraction, but usually it doesn't impact the particles as much as other forces. Mostly it cannot compete with the temperature (kinetic energy) of substances.
It's the weakest attraction of them all.



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

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.

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.

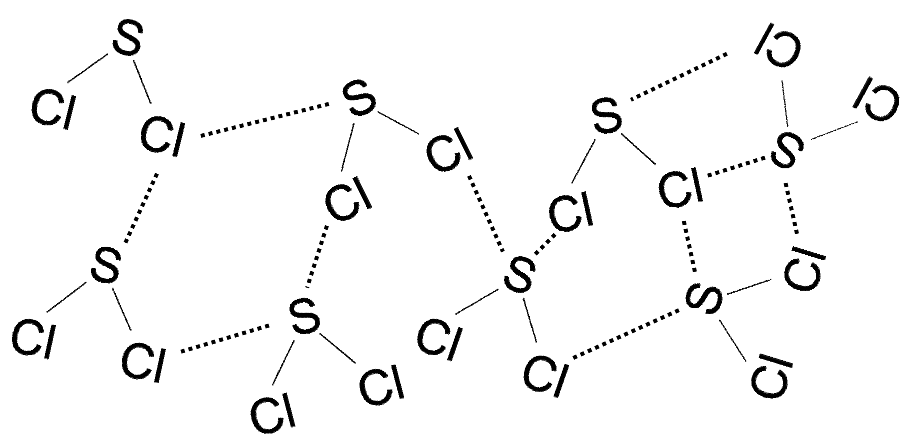
N_2 has a triple NONPOLAR covalent, but again NO DIFFERENCE in EN VALUES.

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.

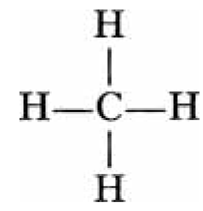
When these molecules are close together, they act like weak magnets. Chlorines (-) of one molecule are attracted to the sulfur (+) of other molecules, this near constant attraction is called DIPOLE ATTRACTION.



Above are eight 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.

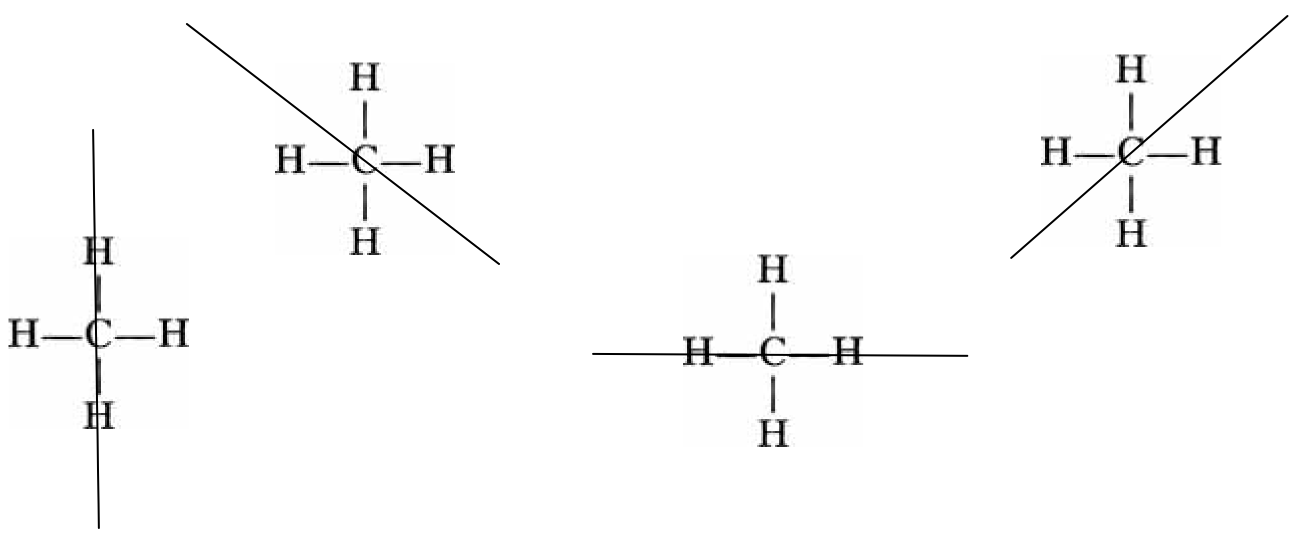
This molecule has RADIAL SYMMETRY. That is the same sort of balance as a pizza pie.

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 inside. 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.

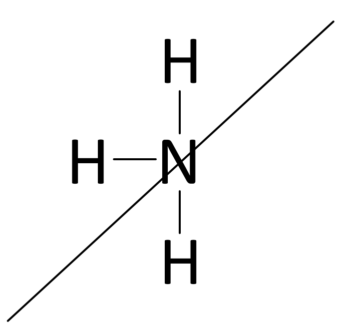
To get dipole attraction, you need polar bonds IN polar molecules. The shape of a molecule that determines if if is polar or nonpolar.

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

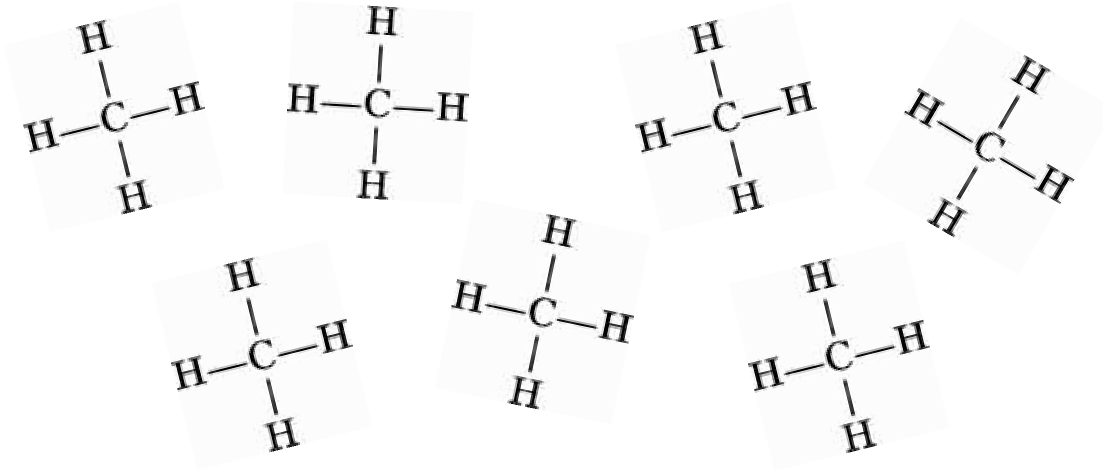
Humans (and gingerbread men) have bilateral symmetry. So does a water molecule. They have only one plane of symmetry. Water molecules are POLAR, the polarity of the bonds just can't cancel out.



Any cut, though the center of the molecule gives you two EQUAL sides. This is radial symmetry, the bond polarity can cancel out.



In ammonia, the bonds between N-H are polar. Since there are three hydrogen atoms bonded to the ammonia, when you cut it this way, you get one side with 2 H and one side with one H and two extra electrons that nitrogen is not bonding with. This is NOT 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 or other polar molecules, they are most likely going to be gases 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 NYS!

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

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

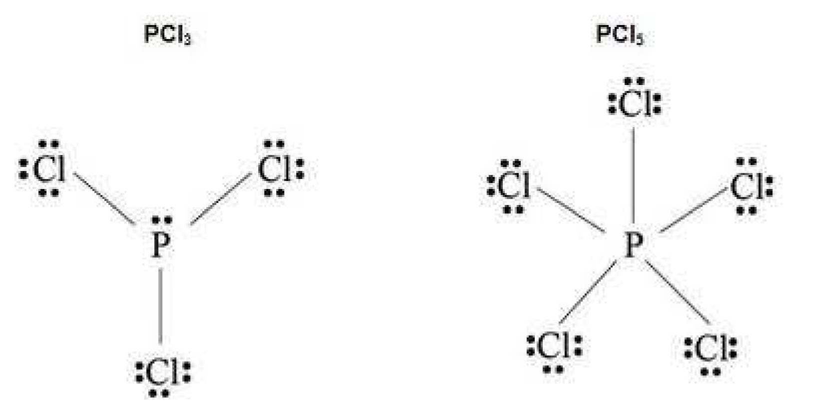
There are several bonds that we need to learn about that do not follow the "rules" of bonding, but somehow they exist, and they are worth looking at.

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 UNSHARED electrons shown at top. This molecule does not have radial symmetry, it's polar. So is PCl_5

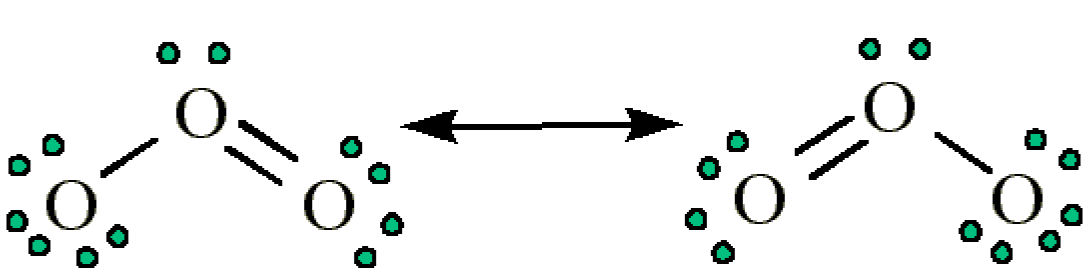
As



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 at right have three pairs each of unshared electrons. The P atom at right seems to have 10 electrons, and it does!

It breaks the octet rule.

As you might remember, I am from a place in Queens called <u>Ozone</u> 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 as below left. This switches around in an instant to the other side, which is also not stable. These bonding styles are both a bit unstable, so ozone RESONATES the bonds back and forth quickly. This shows <u>RESONTATING BONDS</u>.



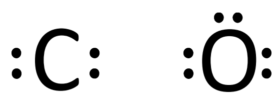
Neither side is "stable" and what ends up happening is that it forms as the left shows, realizes it's not stable so it reverts to the right side, but then realizes it's unstable that way too, so it reverts back to the first way (and over and over). This ozone exists, and the bonding makes no sense, unless you "agree" that this can flip back and forth. It ends up forming an approximate 1½ sized bond on both sides rather than a double/single as shown. That's because this resonance is FAST.

COORDINATE COVALENT BONDS

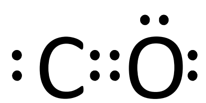
CO2 makes two double polar covalent bonds. Since the molecule has radial symmetry (balance) the molecule is nonpolar. When CO, carbon monoxide forms, there is NO WAY that you can get the electrons to balance unless you know a tricky bond called coordinate covalent.

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:

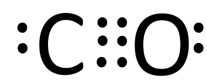


So what happens, strangely, is that CO makes a double polar covalent bond this way (first)



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

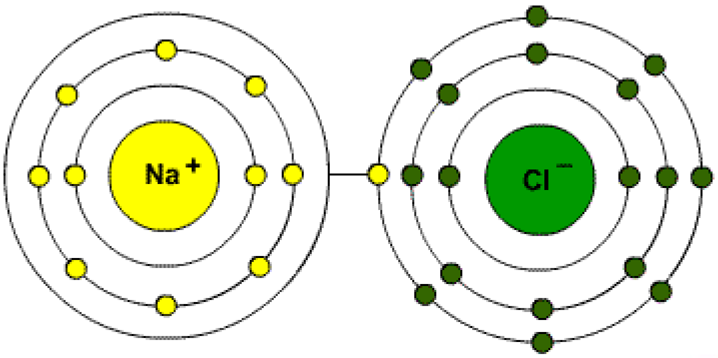
A coordinate covalent bond can be described as oxygen coordinating an octet for carbon by "lending" 2 of it's (top) unshared electrons to the double bond, making what LOOKS LIKE a triple bond, but is really a double polar covalent bond, PLUS a coordinate covalent bond. It looks like this:

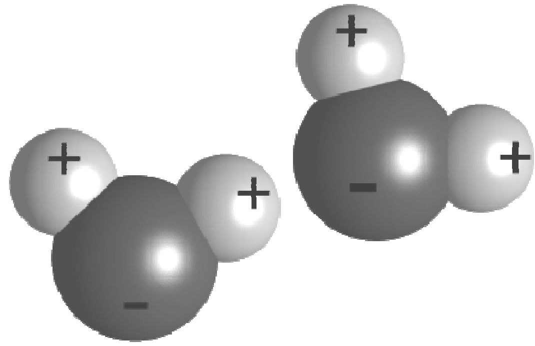


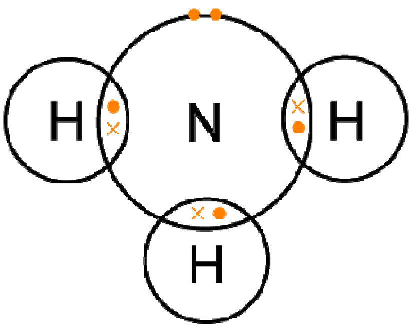
This "appears" to be a triple bond, but notice that carbon has 2 unshared electrons at left, and only has 2 more of it's electrons in the center. The extra electron carbon seems to "have" is from the oxygen atom.

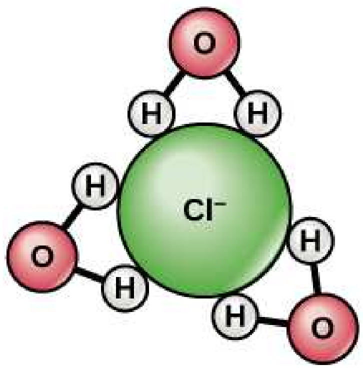
This is a double polar covalent bond PLUS two electrons just being put into the middle so carbon "feels" like it gets an octet too. Not normal, but CO exists in our life, we need to keep track of this bond.

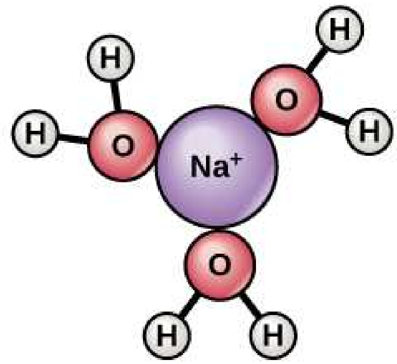
There are many ways to show bonding, many are here. These are not used often, but you should be able to figure them out if you care to. NaCl forms when Na transfers an electron to Cl. Two water molecules with polar bonds (hydrogen are +). One H is attracted to the neighboring oxygen, via hydrogen bonding.



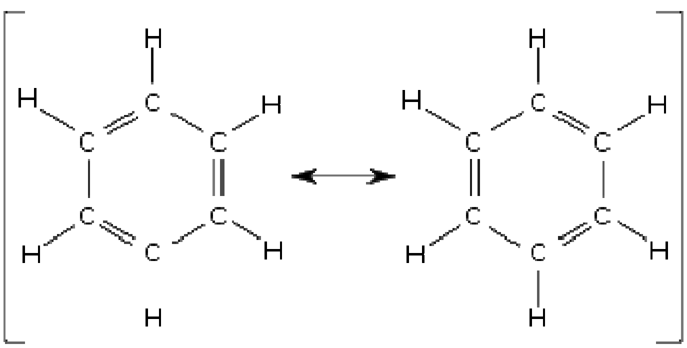








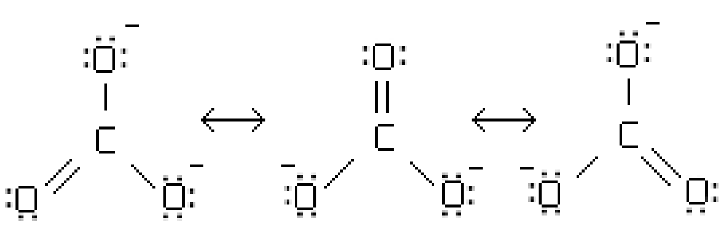
Ammonia forms when three H atoms bond to one N atom. The nitrogen has one unshared pair of electrons at top. Next is water molecules that surround sodium and chlorine ions ($NaCl_{(AQ)}$) Note the orientation of the water molecules, their polarity "point" them at the ions in a particular way.

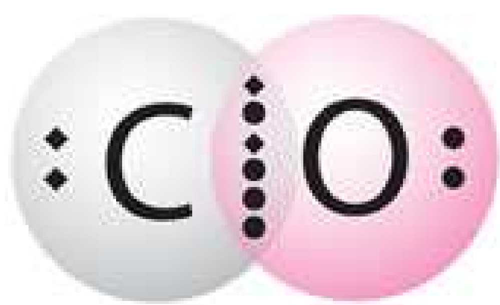


C_6H_6 resonates with 3 triple bonds and 3 single bonds between the carbon atoms. Neither is more stable than the other, so they will resonate back and forth.

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.

At bottom right is CO, showing the carbon electrons as diamonds, and the oxygen electrons as dots. Count them up. There are 6 in the center, 2 carbon diamonds, and 4 oxygen dots. Both atoms end up with an octet, only because oxygen coordinates this with the coordinate covalent bond.





<u>Alloys</u>

Alloys are mixtures of elements that contain at least one metal, often 2 or more metals. The 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 sterling silver (used in forks and knives) made from silver and copper melted together, stainless steel (for scalpels and tools) made from iron and chromium mixed together. Brass (trumpets, tubas, and French horns) is made from melting together zinc and copper.

Some alloys are made from metals plus nonmetals as well, such as: cast iron (used in plumbing pipes) made from iron and carbon.

These elements get mixed and when allowed to solidify, the atoms pack differently and often give stronger, more durable metallic mixtures with properties that make them better suited for certain uses.

It’s important to note again: these are mixtures, the metals, or metal-nonmetal mixtures are not compounds. These alloys have no set formulas, but certain quantities of each element make for somewhat stronger, or more lustrous, or increase some quality. They can be melted apart and the elements should separate by density.

The Bronze Age started about 5000 years ago. During that time humans learned to melt together the metals copper and tin. Together the tools they made were much stronger, and golden in color. Below are drawings of weapons from that period from all over Europe. Wooden handles, or even spears were attached to the metal points.

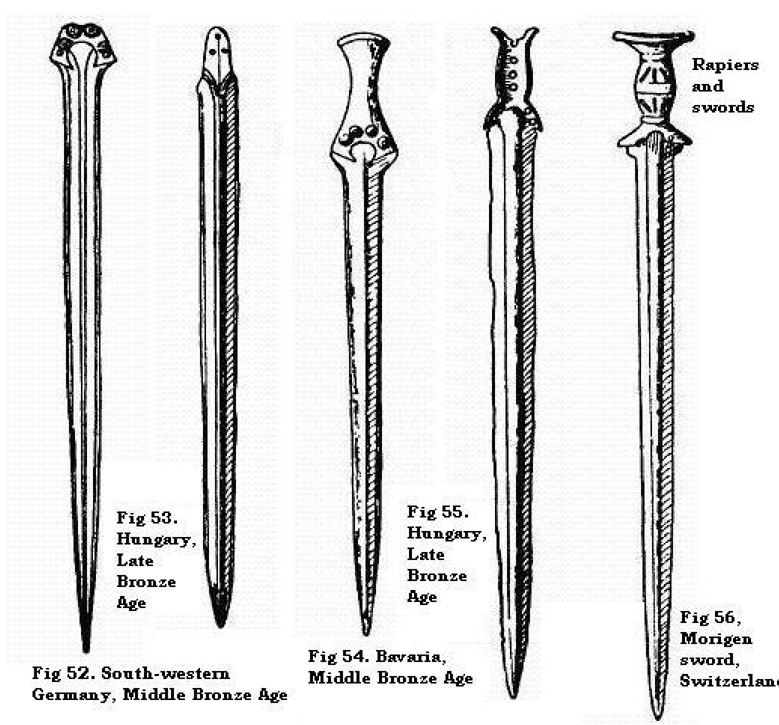


Fig 52. South-western Germany, Middle Bronze Age

Fig 53. Hungary, Late Bronze Age

Fig 54. Bavaria, Middle Bronze Age

Fig 55. Hungary, Late Bronze Age

Fig 56, Morigen sword, Switzerland

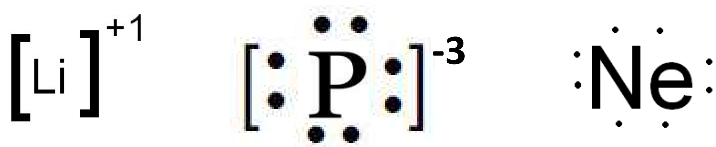
PRACTICE QUESTIONS Set 1 (answers on the last page)

1. What are the electron configurations for phosphorous and for magnesium? What are their ion electron configurations?
P has a 2-8-5 electron configuration, the Mg has a 2-8-2. The P^{-3} anion is 2-8-8, the Mg^{+2} cation is 2-8.

2. Which electrons are drawn in Lewis dot diagrams? **ONLY the outermost ones, the VALENCE ELECTRONS**

3. Draw Lewis Dot Diagrams for the lithium atom, boron atom, calcium cation, and the oxide anion.
The lithium atom has 1 valence electron, the boron atom has 3 valence electrons. The calcium cation transfers 2 electrons and ends up with an empty outermost orbital, oxygen gains 2 electrons to get its octet

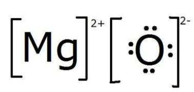
4. Will an atom of aluminum ever normally lose 1 or 4 electrons? Why not? **No, aluminum is a metal, it loses enough electrons to become isoelectric to a noble gas. Al start as 2-8-3, when it loses 3 electrons to become Al^{+3} that changes to 2-8 (like neon)**

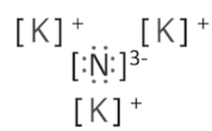


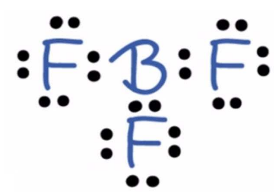
PRACTICE QUESTIONS Set 2

1. Draw the Lewis dot diagrams for Li^{+1}, P^{-3}, and for Neon.

2. Draw Lewis Dot diagrams for magnesium oxide and for potassium nitride.







3. Define ISOELECTRIC. **Isoelectric means obtaining the same electron configuration as a noble gas. This happens when metals lose exactly enough electrons to match up their electron configuration with a noble gas, or when a non-metals gain enough electrons to do this same thing. They DON'T become noble gases, they get similar electron configurations.**

PRACTICE Questions set 3

1. Draw larger sized structural diagrams for KCl, for boron trifluoride, carbon tetra bromide.
2. Carbon dioxide... Have double polar covalent bond.
3. Name the bond types in oxygen difluoride. This has 2 single polar covalent
4. Name the bond types in ammonium chloride. IONIC!