Chapter 6 Bonding Diary

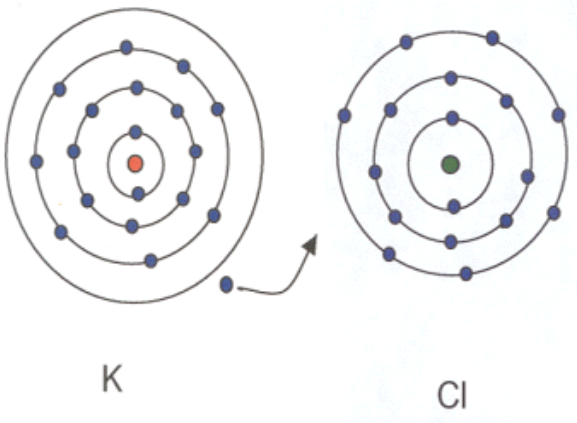
Bonding deals primarily with how ionic compounds and molecular compounds form. As you might remember from early on, it's all about the ELECTRONS. Are they present, or have they been lost forming a cation? Are they equal to the number of protons, or have some been added when an anion is made? Finally, are they being shared by more than one atom in a covalent bond (the "co"– meaning shared, and "valent"– meaning valence electrons), and how equally or how unequally are they being shared?

Ionic bonding

Atoms on the left side of the periodic table are metals. Metals tend to lose electrons to get that coveted complete outer orbital of electrons, just like the NOBLE GASES have. Metal atoms will lose one, two, or even three electrons to get that full valence orbital (stable octet), to become isoelectric to the noble gases. When they lose these valence electrons they lose negative charges, making them positively charged cations. Metals ONLY lose electrons to form cations, they do not gain electrons. Metal atoms can only lose these valence electrons when a potential electron acceptor is nearby and close enough to "catch" these electrons. Metals do not release these electrons to the universe, they can only give these electrons to willing non-metals, and only if they can all be put into a non-metal orbital.

Atoms on the right side of the periodic table are non-metals. Non-metals tend to gain electrons to get that coveted complete outer orbital of electrons, just like the NOBLE GASES have. Non-metal atoms gain one, two, or even three (!) electrons to get that full valence orbital (stable octet), to become isoelectric to the noble gases. When they gain these extra valence electrons they gain more negative charge, making them negatively charged anions. Non-metals ONLY gain electrons to form anions, they do not lose electrons. Non-metals only gain these electrons when a suitable metal atom is close enough to "give" these electrons away. Non-metals can not grab electrons from the universe, they can only come from "willing" metal atoms and only if all the electrons can be fit into the "unfull" orbitals.

On the left is an orbital model (Bohr) of potassium and chlorine. The valence electron from the 4th orbital of potassium is given to the third orbital of chlorine. That leaves potassium with a +1 charge, but a 2-8-8 electron orbital (it's isoelectric to argon). Chlorine gains an 8th electron, filling it's valence orbital, becoming 2-8-8, three full orbitals as well. Chlorine becomes a –1 anion, and the two ions are very attracted to each other due to difference in charges. They form an IONIC BOND.



On the right is the Lewis Dot Diagram of KCl. Potassium gives up its single valence electron in the 4th orbital, leaving it with a +1 charge. The Chlorine anion picked up this electron, giving it a FULL orbital, with 8 electrons, shown in red. The proper form for Lewis Dot diagrams and ions is to put the ions into BRACKETS and then show the charges of the ions. Also, show the original valence orbitals with the NEW valence electrons (potassium with ZERO left, and chlorine now with 8). Lewis dot diagrams can be used to represent all ionic compounds. The cations get brackets with the appropriate "+" charge, and the anions all get 8 electrons inside the brackets, with the appropriate "–"charge sign.

Ionic bonds are very polar, with EN differences nearly always over 2.0, and very strong. Ionic compounds become electrolytes in solution and when melted. The ionic compound will become ions in aqueous solution. These ions allow for electricity to flow freely. You can't get shocked in distilled water, pure water has NO IONS, and can't conduct electricity.

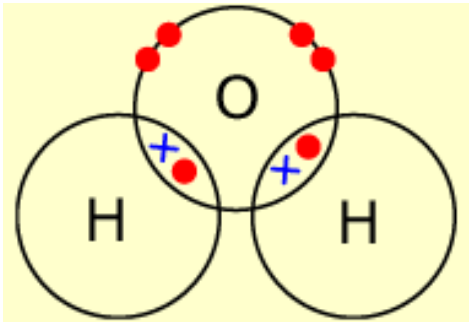
If you MELT an ionic crystal (very very hot), melted sodium chloride for example, the ions free from each other, just like if it were dissolved into water, so LIQUID ionic compounds, like $NaCl_{(L)}$, or $KCl_{(L)}$, are in fact electrolytes.

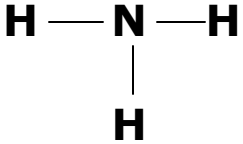
Covalent Bonds

When 2 or more NON-metals bond, by sharing electrons, the bonds are called covalent bonds. "Co" means shared, and "valent" refers to valence electrons. So, covalent bonds are when 2 or more atoms share valence electrons, all to get that special complete outer orbital (the octet rule).

The Lewis Dot diagrams do not get brackets, as all atoms retain their electrons, but they get to borrow some, or share some with other atoms.

On the right is WATER. The oxygen has the six RED dot valence electrons. 2 pair of electrons at top are unshared pairs, and the two at the bottom, they each bond with a hydrogen, giving oxygen the 8 electron full orbital. Each hydrogen gets to share back an oxygen electron, giving each a full small outer orbital.





Next is ammonia, another molecular compound. The nitrogen has one pair of unshared electrons at top, and three different hydrogen bonded to the nitrogen, making the expected formula of NH_3.

Below it is the STRUCTURAL FORMULA, showing each shared pair of electrons as a single line, or single bond. The unshared pair of electrons at the top of the nitrogen can be drawn in, or sometimes they are just left out of the diagram though it is understood that they are there.

All molecules can be drawn as Lewis Dot Diagrams, or as Structural Diagrams, which are less tiresome to draw, but require you to truly understand the octet rule, and be able to count to eight without having to SEE all 8 dots!

Double and Triple Covalent bonds

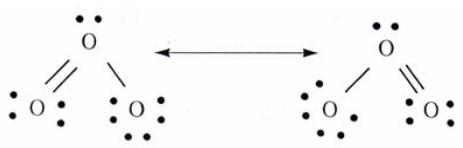
When oxygen exists as a diatomic molecule, it will look like this. Each oxygen can have an octet of electrons, only if a DOUBLE bond, or sharing of 2 pairs of electrons happens.

Carbon dioxide is similar, in that, the carbon and each oxygen make a double bond.

Nitrogen can only exist with 2 octets if the nitrogen share THREE PAIR of electrons with each other, hence a TRIPLE BOND, as shown here.

Nitrogen will have a pair of unshared electrons on either end, and the oxygen molecule just above, they have 2 pairs of unshared electrons each.
Bonds form in order to create a stable octet for each atom in the molecule.

Resonance is when a single and double bond can't decide where to be, so the double bond and single bond trade places quickly back and forth. Both sides of the molecule are single and double bonded, almost at the same time. Ozone, O_3 shows resonance. The bonds resonate back and forth, like sound waves or a pendulum. The double and single bonds trade places repeatedly and constantly. This creates "2 different, but the same" ozone molecules.



Bond Polarity

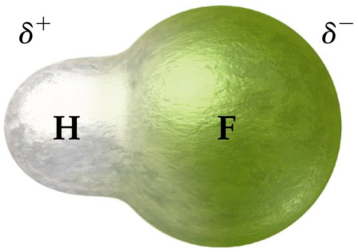
Linus Pauling's relative scale of Electronegativity can be applied to covalent bonds and molecules. As you may recall EN is defined as the ability of an atom to attract electrons in a chemical bond (covalent). So, when 2 atoms decide to share electrons, the one with the higher Electronegativity value will get the electron more than the other atom. It won't share fairly, rather, it will take what it can get.

In the diatomic HONClBrIF twins, since both atoms are identical to each other, the difference in Electronegativity is zero. That means that the bond is always going to be SINGLE COVALENT, and also NON-POLAR. Since neither of the two atoms gets an electron more of the time, neither side is negative or positive, so the **bond** is not polar.

A **Nonpolar bond** exists when valence electrons are **shared equally.**

But, with something like HF, hydrogen fluoride, the Electronegativity values (EN) are not the same, and there is a consequence that we need to watch out for.

HF has 2 atoms, hydrogen and fluorine. The EN values for each are 2.1 and 4.0 respectively. That is a difference of 1.9, which is huge. Since the F has such a high EN value, it's the fluorine that gains the electron most of the time, and that leaves the hydrogen "electron-less" most of the time. That means that the bond is now POLAR, there is a clear positive side (the hydrogen) and a negative side, the fluorine.



A **polar bond** exists when valence electrons are **shared unequally.**

The lower case Greek letter sigma means "a little bit", and in the picture the H side is a "little bit" positive" and the fluorine side is a "little bit" negative.
With any molecular compound we need to look at the atoms involved in the bond to determine if the electrons are being shared equally or unequally. Then we can tell if it is polar or nonpolar.

HCl has EN values of 2.1 and 3.2 for chlorine. This large EN difference means that chlorine gets the electron most of the time from hydrogen, giving it a negative end of that bond, leaving the hydrogen more positive.

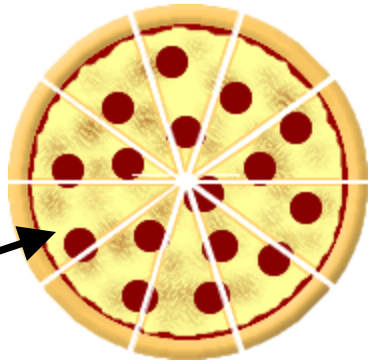
The atom with the HIGHER EN value gets that electron most of the time. Sometimes you may see a special arrow, showing where the electron "WAS" and where it "WENT". This arrow is called the DIPOLE ARROW. Dipole means 2 poles or electrostatic charged ends of a polar bond.

The "plus" sign shows that the Hydrogen is now more positive since it "lost" its electron to chlorine. The arrow head indicates where the electron went, which also means that the chlorine side of the bond is negatively charged. It is a polar bond.

Molecular polarity

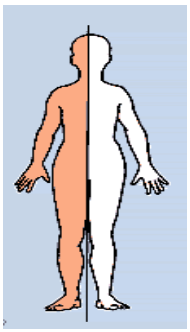
Just as bonds can be polar or non-polar, whole molecules can be polar or non-polar as well. To grasp molecular polarity we need to learn a new term, RADIAL SYMMETRY.

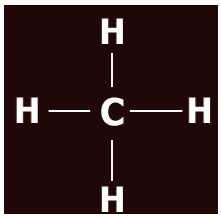
Symmetry means the same on both sides. Radial symmetry means that as long as you cut something in half through the middle, from ANY direction, both halves end up symmetrical. A circle has radial symmetry. A drawing of a valentine heart, or a stick figure of a man do NOT have radial symmetry.



The pizza is cut into 10 pieces. Any straight cut through the center will result in 2 identical halves. Molecules with radial symmetry are **NON-POLAR**. Without this type of symmetry, the molecules are polar.

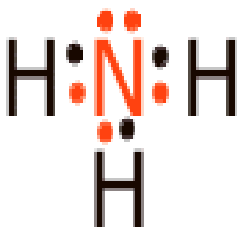
A human can be "cut" in half only in one particular plane, but no other cut will result in 2 symmetrical pieces. Bilateral symmetry is nice, but it will not make a molecule non-polar. Only a molecule with radial symmetry is non-polar.

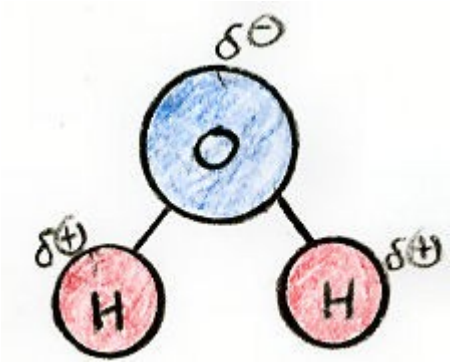




Methane, CH_4, has radial symmetry, no matter how you cut it in half, you get "half" a carbon atom, and 2 hydrogen atoms. Therefore the molecule is symmetrical. Each of the carbon-hydrogen bonds is polar, but since each hydrogen "loses" it's electrons to the more electronegative carbon, all the + charges on the hydrogen are balanced, or symmetrical, or, the whole molecule is NON-POLAR.

The Lewis structure of ammonia on the right shows those unshared electron atop nitrogen. Ammonia does not have radial symmetry, there fore the molecule is a **polar molecule**.



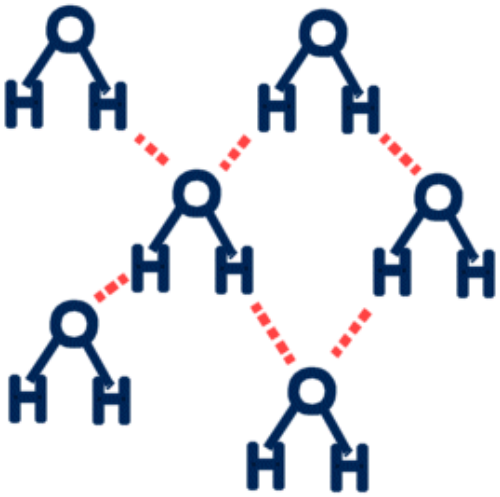


Here is water again. The lower case Greek sigma shows where the molecule is a "little bit" negative on the oxygen side and a "little bit" positive on the hydrogen side. The two pairs of unshared electrons of oxygen are not shown here, but they are atop the oxygen atom.

The H-O bonds are polar, and the molecule is polar as well.

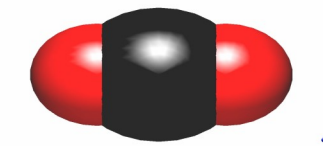
The dipole created by the bonds of H-O in water, or in HCl, or HF, all create what is known as a PERMANENT DIPOLE. Permanent dipoles cause the molecules to all have positive and negative sides. These polar molecules are therefore attracted to each other. The more polar the molecule, the more intermolecular attraction there is. These intermolecular attractions give rise to specific compound properties, such as higher boiling points.

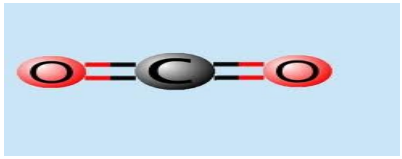
This diagram on the right shows the intermolecular attraction of the positive hydrogen atoms to the negative oxygen atoms in water. These intermolecular forces or attractions give water many of its basic properties.



Molecular shapes

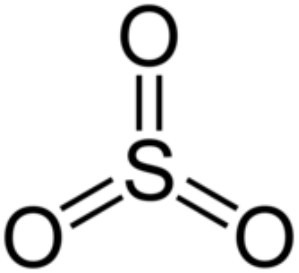
Molecules are so small that most cannot ever be seen. Still, due to their structures, they have shapes that are clear. These shapes are listed below.

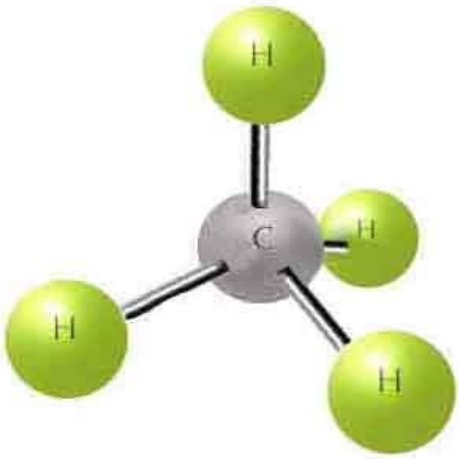




Carbon dioxide, in a space filling model, is a **LINEAR** or a straight line molecule. The double bonds are not shown, but you know they are there because it follows the octet rule. The bonds are polar and the molecule is non-polar due to radial symmetry.

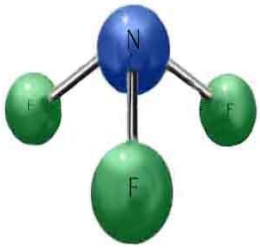
In this structural diagram of sulfur trioxide, the molecule is shown drawn as it exists in three dimensions, flat and **PYRAMIDAL.** The bonds are polar, the molecule is non-polar. It shows radial symmetry



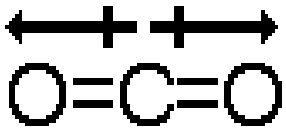


Methane is shown here in 3-D (almost), with the carbon central to the hydrogen surrounding the carbon in exactly 120 degree angles. They are all symmetrical. The C-H bonds are all mildly polar (C 2.6, H 2.1 electronegativity values), but the molecule has radial symmetry. It forms a **TETRAHEDRAL** shaped molecule.

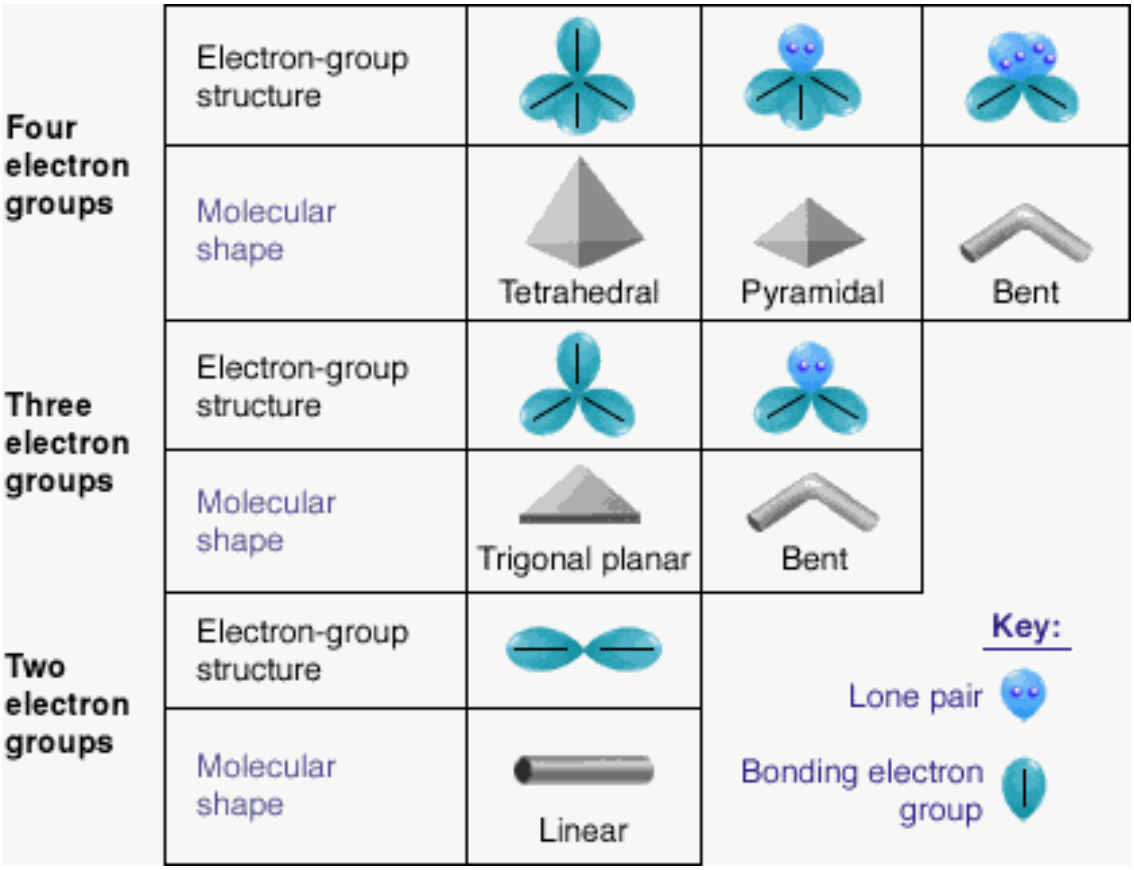
Water is a **BENT** molecule, and is a good example of how water really looks. The two pairs of un-shared electrons atop the red oxygen are not shown, but they are there. Water shows bilateral but NOT RADIAL SYMMETRY, water is a very polar molecule.



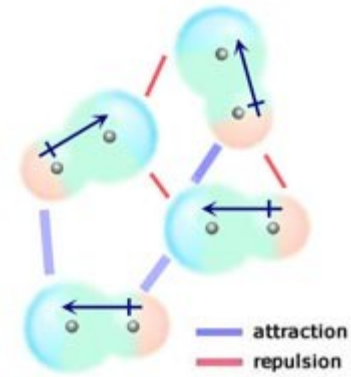
Nitrogen trifluoride, NF_3 , shows a **PYRAMIDAL** shape. **This is the same shape as AMMONIA (NH_3).** The 2 unshared electrons, or ONE PAIR of unshared electrons on top of the nitrogen are not shown (5 valence electrons total, 2 unshared, one each shared with each fluorine here).
These bonds are polar (3.0 to 4.0 EN values). The molecule shows NO RADIAL SYMMETRY, therefore it is a polar molecule.



Above is the carbon dioxide non-polar molecule, with polar double bonds, showing the DIPOLE ARROWS, indicating where the electrons were (with carbon) and where they went, (to the oxygen's). Polar bonds can result in non-polar molecules due to the symmetrical charge distribution. It is a **LINEAR** molecule.



Intermolecular Forces



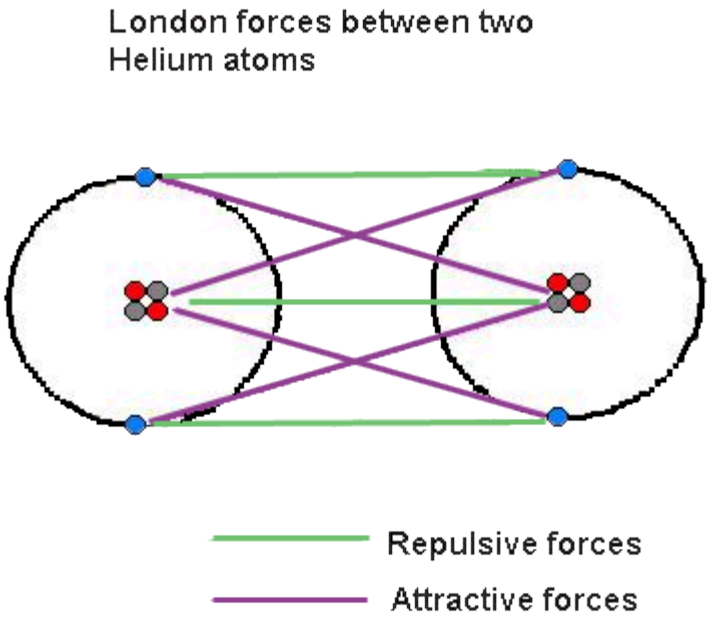
This diagram shows polar molecules, possibly HF, with dipole arrows, and dipole attractions. Permanent dipoles, due to large EN value differences, will make molecules attractive to each other (at least one side to the opposite side of the next molecule). There are three basic types of intermolecular attractive forces.

1. London dispersion forces (aka Van der Waals): interactive attraction due to the temporary movement of electrons in individual molecules

2. Dipole-dipole interactions: intermolecular attraction due to permanent dipoles(2 polar molecules)

3. Hydrogen bonds: strongest of these intermolecular forces, due to permanent dipoles that include hydrogen atoms (polar molecules).

Vander waals forces can be explained by looking at the halogens in group 17. Top to bottom they are F, Cl, Br, and I. Both F_2 and Cl_2 are gases at STP, that is because they have collectively in each molecule 18 and 34 electrons each. As these electrons fly around they can temporarily find themselves out of balance. This creates a short lived but real temporary dipole. This makes the molecules attractive to each other for short times. Of course, the kinetic energy in these gases is much too strong to allow such weak temporary dipoles to force them into a liquid. These dispersion forces exist, but at STP they are nearly irrelevant to the molecules.

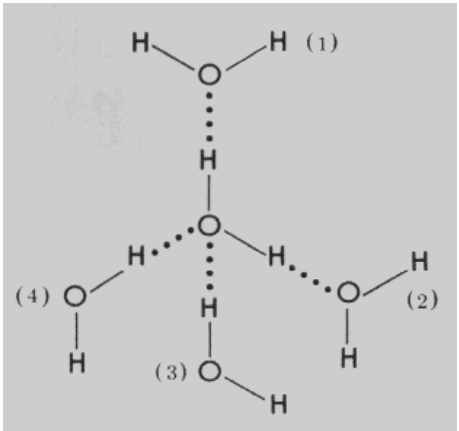


Bromine has 70 electrons in its molecule (Br_2). These electrons also fly around, and because there are so many more, these dispersion forces accumulate. At STP bromine is in fact a liquid! It's due to the fact that it has a lot of electrons that create enough short lived dipoles.

Iodine has 106 electrons (I_2). With this many electrons the dispersion forces are so strong that iodine is in fact a solid! at STP. The dispersion forces are weak, but when there are SO MANY, they have a large impact on the phase of matter.

Dipole-dipole interactions occur when polar molecules with polar bonds that force polar molecules to form have an intermolecular attraction to each other. The positive sides and the negative sides are attracted to each other inter-molecularly. This is just like dispersion forces, but those are temporary, these are permanent, due to large EN differences.

Hydrogen Bonding occurs when the dipole attractive forces are due to having a hydrogen atom bonded to a strongly electronegative atom such as oxygen, nitrogen or fluorine, that have their own unshared pairs of electrons. These are the strongest of the van der waals forces of intermolecular attraction.



Metallic Bonding

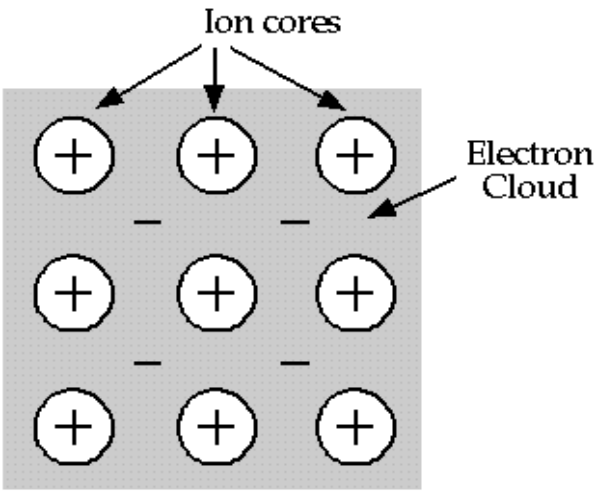
Metallic bonding occurs when metals bond to themselves. A hunk of copper has zillions of copper atoms bonded together, by metallic bonding. An ionic compound can NEVER have metallic bonding. A molecule (2 or more non-metals) can never have metallic bonding either.

The metals will form into cations (nucleus of neutrons plus protons, PLUS the inner electrons), and the VALENCE ELECTRONS will escape the nucleus and float around all the cations.

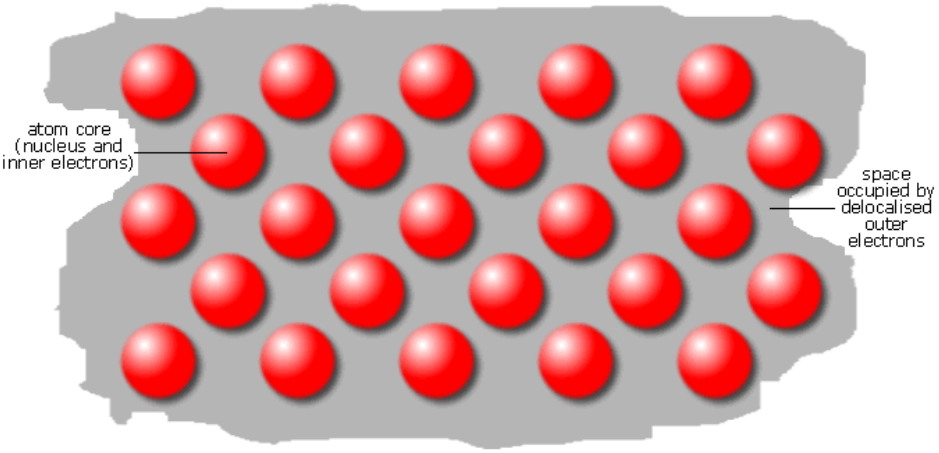
The cations are awash in a SEA OF MOBILE VALENCE ELECTRONS.

When metals are pounded hard, the cations squeeze together, which is a problem since they are all positively charged. They should force themselves apart rather than get close together. Instead, the valence electrons move to offset that positive charge, allowing the metal to physically take a new shape (malleable and ductile).

If electricity is run through a metal, since electricity is the FLOW OF ELECTRONS IN ONE DIRECTION, electrons enter a piece of metal and jump out the other end to keep the flow moving. Different electrons enter and leave, but the correct number of valence electrons remain moving around the cations, otherwise once the electricity was cut off, you'd be left with real cations rather than a neutral piece of metal.



ION CORES are the cations (nucleus plus inner electrons). The electron cloud is invisible of course, with the electrons flowing between the cations.



The gray zone is the space where the electrons are moving around. The red balls indicate the location of the metal cations.

Properties of Metals, non-metals, metalloids, ionic & molecular compounds

property	metals	non-metals	metalloids	ionic compounds	molecular compounds
luster	yes	no	some are, such as Si	no	not usually
malleable, ductile	yes, yes	no, no	no	no, no	no
conduct heat	yes	no	no	no	no
conduct electricity	yes	no	some, such as Si	no	no
form cations or anions?	cations only	anions only	follows metal non-metal rules	formed from ions but are neutral	none
smallest particle	atoms	atoms or molecules	atoms	formula units	molecules
density	fairly high	low	varies, check table S	medium high	varies
melting point	relatively high	relatively lower	varies	extremely high	low compared to metals or ionic compounds
phases at room temp	solid	solid, liquid, or gas	solid	solid	solid