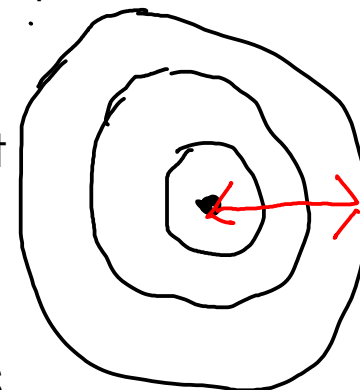


PERIODIC TRENDS

- Some properties of elements can be related to their positions on the periodic table.

ATOMIC RADIUS

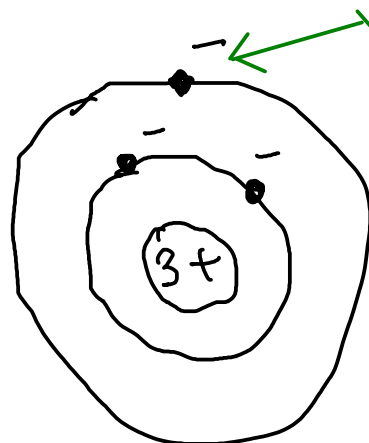
- The distance between the nucleus of the atoms and the outermost shell of the electron cloud.
- Relates to the size of the atom.
- As you go DOWN A GROUP (↓), the atomic radius INCREASES.



- Why? As you go down a period, you are ADDING SHELLS!

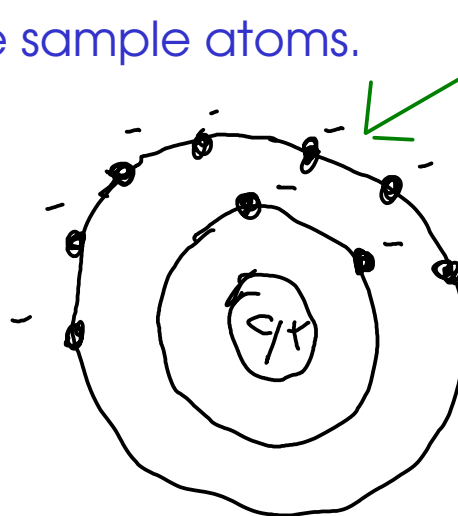
- As you go ACROSS A PERIOD (→), the atomic radius DECREASES

Why? Let's look at some sample atoms.



Outer electron sees an effective +1 charge (shielded from +3 nucleus by 2 electrons)

lithium $1s^2 2s^1$



Outer electrons see an effective +7 charge (shielded from +9 nucleus by 2 electrons)

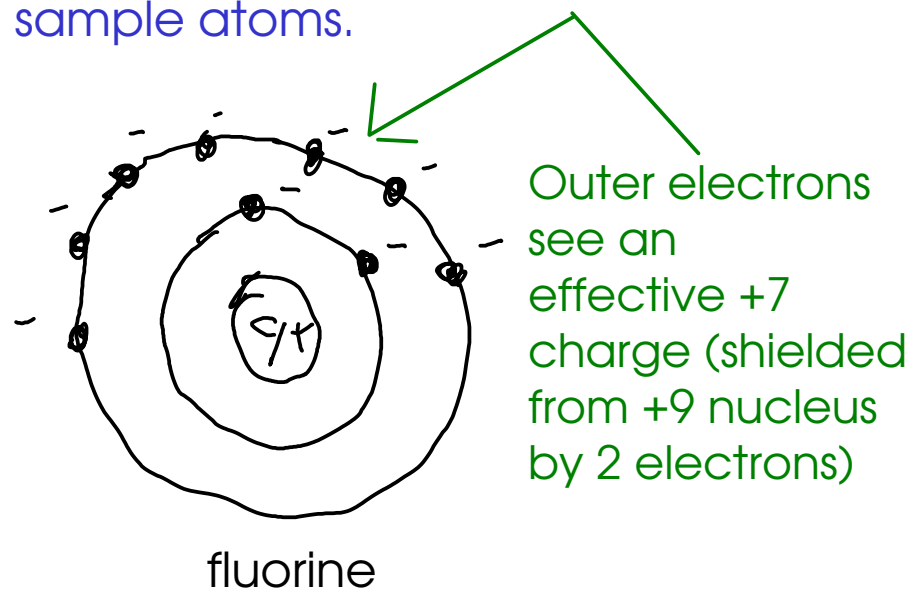
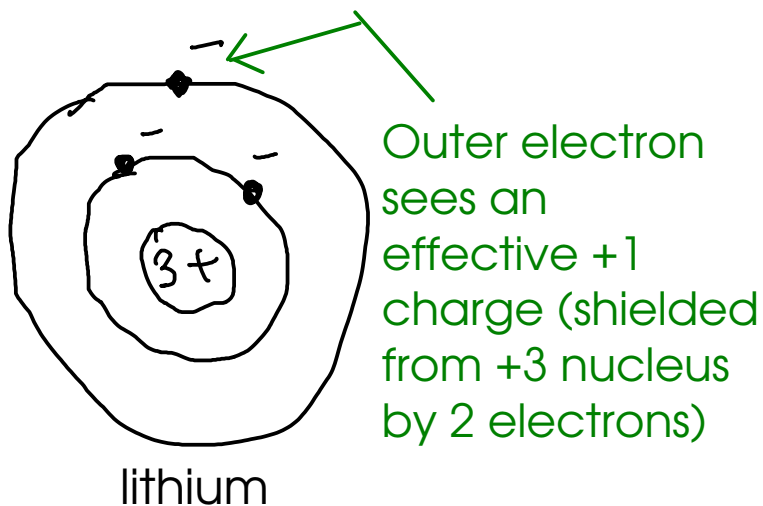
fluorine $1s^2 2s^2 2p^5$

... so fluorine's outer shell is pulled closer to the nucleus than lithium's!

(FIRST) IONIZATION ENERGY

- The amount of energy required to remove a single electron from the outer shell of an atom.
- Relates to reactivity for metals. The easier it is to remove an electron, the more reactive the metal.
- As you go DOWN A GROUP (↓), the ionization energy DECREASES.
 - Why? As you go down a period, you are ADDING SHELLS. Since the outer electrons are farther from the nucleus and charge attraction lessens with distance, this makes electrons easier to remove as the atoms get bigger!
- As you go ACROSS A PERIOD (→), the ionization energy INCREASES.

- Why? Let's look at some sample atoms.



... since fluorine's outer electrons are held on by a larger effective charge, they are more difficult to remove than lithium's.

THE FIRST TWO PERIODIC TRENDS IN A NUTSHELL

LARGER
IONIZATION
ENERGYSMALLER
RADIUS

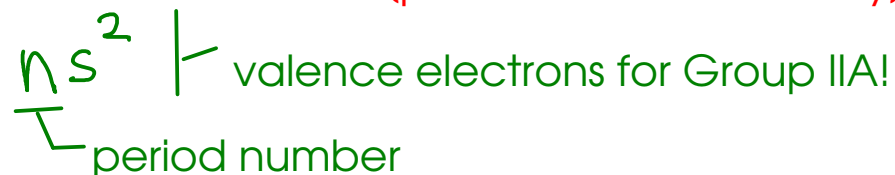
IA		IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA
H	Li	Be											B	C	N	O	F	Ne
Na	Mg	III B	IV B	V B	VI B	VII B	VIII B	IB	IIB			Al	Si	P	S	Cl	Ar	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
Fr	Ra	Ac*	Rf	Db	Sg	Bh	Hs	Mt	*"inner" transition metals go here									

LARGER
RADIUSSMALLER
IONIZATION
ENERGY

- the electron affinity is the ENERGY CHANGE on adding a single electron to an atom.
 - Atoms with a positive electron affinity cannot form anions.
 - The more negative the electron affinity, the more stable the anion formed!
- General trend: As you move to the right on the periodic table, the electron affinity becomes more negative.

EXCEPTIONS

- Group IIA does not form anions (positive electron affinity)!



- To add an electron, the atom must put it into a higher-energy (p) subshell.

- Group VA: can form anions, but has a more POSITIVE electron affinity than IVA



└── Half-full "p" subshell! To add an electron, must start pairing!

- Group VIIIA (noble gases) does not form anions



CHEMICAL BONDS

- A CHEMICAL BOND is a strong attractive force between the atoms in a compound.

3 TYPES OF CHEMICAL BOND

Type	Held together by...	Example
Ionic bonds	attractive forces between oppositely charged ions	sodium chloride
<u>Covalent bonds</u>	sharing of valence electrons between two atoms (sometimes more - "delocalized bonds")	water
* Metallic bonds	sharing of valence electrons with all atoms in the metal's structure - make the metal conduct electricity	any metal

*For CHM 110, you don't need to know anything more about metallic bonds than what's in this table. If you take physics, you may learn more about the characteristics of the metallic bond.

... so how can you tell what kind of bond you have? You can use the traditional rules of thumb:

- Metal-Nonmetal bonds will be ionic
 - Nonmetal-nonmetal bonds are usually covalent
- Metalloids act like NONMETALS, here.

... but for better information about bonding, you can use ELECTRONEGATIVITY.

ELECTRONEGATIVITY:

-A measure of how closely to itself an atom will hold shared electrons

p346:
chart of
electroneg.
values

p352,
10th

... in other words, how ELECTRON-GREEDY an atom is!

Bonds with ...	are ...	Examples
Little or no difference in electronegativity between atoms	NONPOLAR COVALENT	C-C, C-H, etc.
Larger differences in electronegativity between atoms	* POLAR COVALENT	H-F, C-F, C-Cl, etc.
Very large differences in electronegativity between atoms	IONIC	NaCl, KBr, etc.

* A POLAR bond is a bond where electrons are shared unevenly - electrons spend more time around one atom than another, resulting in a bond with slightly charged ends

- You may look up electronegativity data in tables, but it helps to know trends!

	IA	IIA										IIIA	IVA	VA	VIA	VIIA	
2	Li	Be										B	C	N	O	F	
3	Na	Mg	IIIB	IVB	VB	VIB	VIIB	VIIIB	IB	IIB		Al	Si	P	S	Cl	
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I
6	Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At
7	Fr	Ra	Ac*	Rf	Db	Sg	Bh	Hs	Mt	*"inner" transition metals go here							

Notes:

- ① - FLUORINE is the most electronegative element, while FRANCIUM is the least!
- ② - All the METALS have low electronegativity
- ③ - HYDROGEN is similar in electronegativity to CARBON

(p 346)

... so C-H bonds are NONPOLAR

DESCRIBING CHEMICAL BONDING

"octet rule"

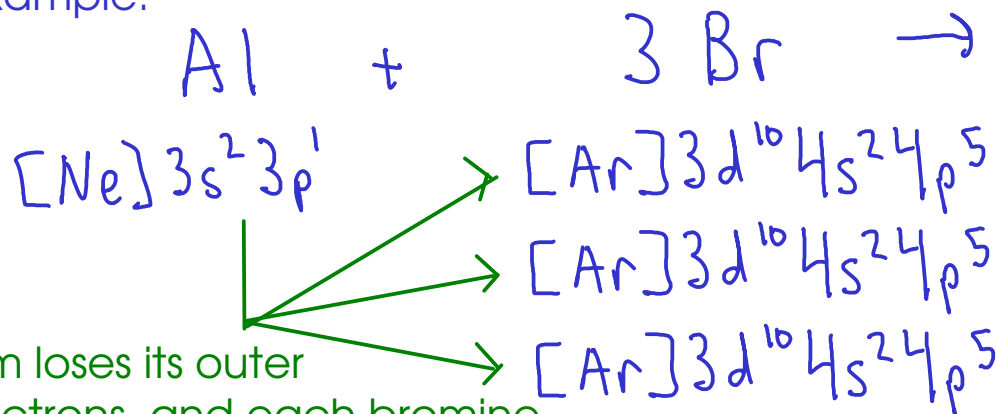
- a "rule of thumb" (NOT a scientific law) predicting how atoms will exchange or share electrons to form chemical compounds
- atoms will gain, lose, or share enough electrons so that they end up with full "s" and "p" subshells in their outermost shell.

- Why "octet"? An "s" subshell can hold two electrons, while a "p" subshell can hold six. $2+6 = 8$

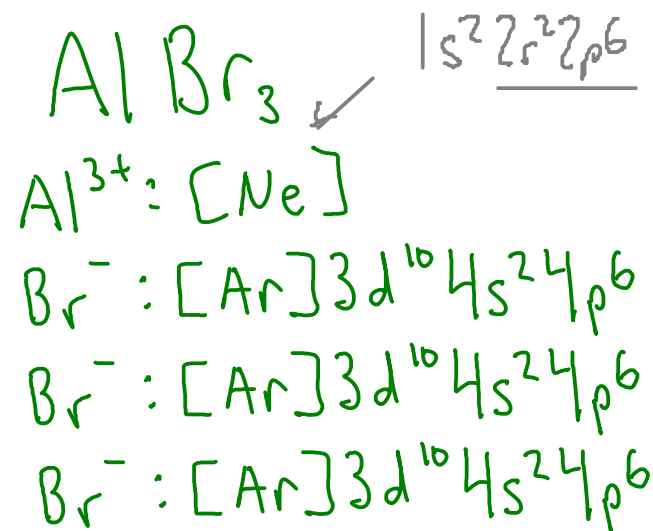
IONIC COMPOUNDS

- When atoms react to form IONS, they GAIN or LOSE enough electrons to end up with full "s" and "p" subshells.

example:



Aluminum loses its outer three electrons, and each bromine gains one!



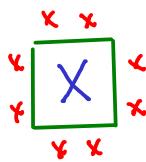
... but using electron configurations to describe how aluminum bromide forms is a bit cumbersome! Can we simplify the picture a bit?

LEWIS NOTATION / ELECTRON-DOT NOTATION

- Lewis notation represents each VALENCE electron with a DOT drawn around the atomic symbol. Since the valence shell of an atom contains only "s" and "p" electrons, the maximum number of dots drawn will be EIGHT.

- To use electron-dot notation, put a dot for each valence electron around the atomic symbol. Put one dot on each "side" of the symbol (4 sides), then pair the dots for atoms that have more than four valence electrons.

examples:



More examples



Which "side" you draw the dots on isn't important, as long as you have the right number of electrons and the right number of "pairs"



To draw a dot structure for an atom, you need to know HOW MANY valence electrons it has! You can determine this simply from the periodic table, WITHOUT writing the whole electron configuration!

The number of valence electrons equals the group number in the A/B group numbering system FOR "A" GROUPS!

	IA												VIII A								
1	H	IIA											III A	IVA	VA	VIA	VIIA	He			
2	Li	Be											B	C	N	O	F	Ne			
3	Na	Mg	IIIB	IVB	VB	VIB	VIIB	VIIIB	IB	IIB						Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr			
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe			
6	Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn			
7	Fr	Ra	Ac*	Rf	Db	Sg	Bh	Hs	Mt												

2 valence electrons

1 valence electron

3 valence electrons

4 valence electrons

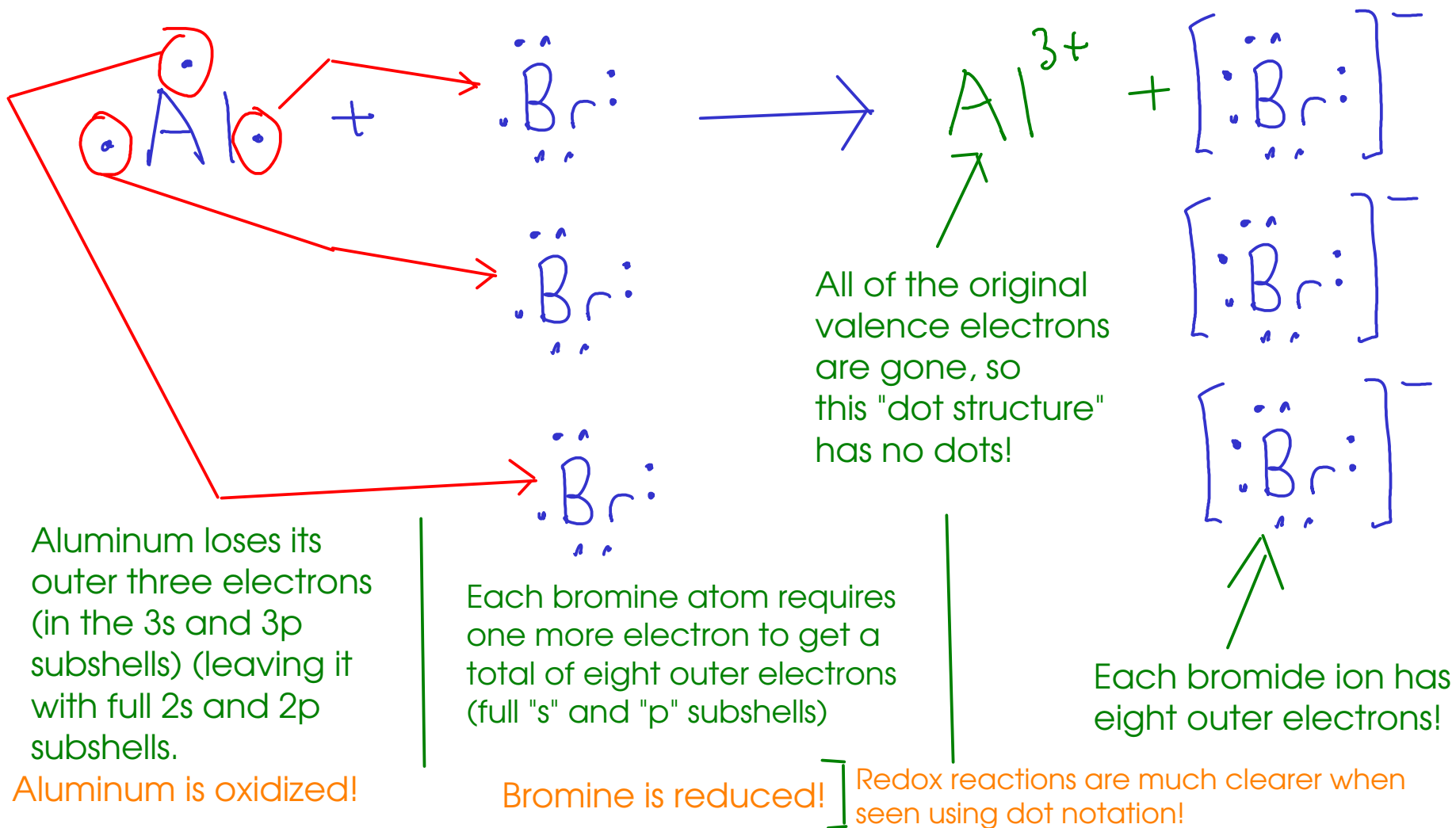
5 valence electrons

6 valence electrons

7 valence electrons

8 valence electrons (except helium!)

... but how do we use this to describe a reaction that produces ions? Let's look at our previous example!

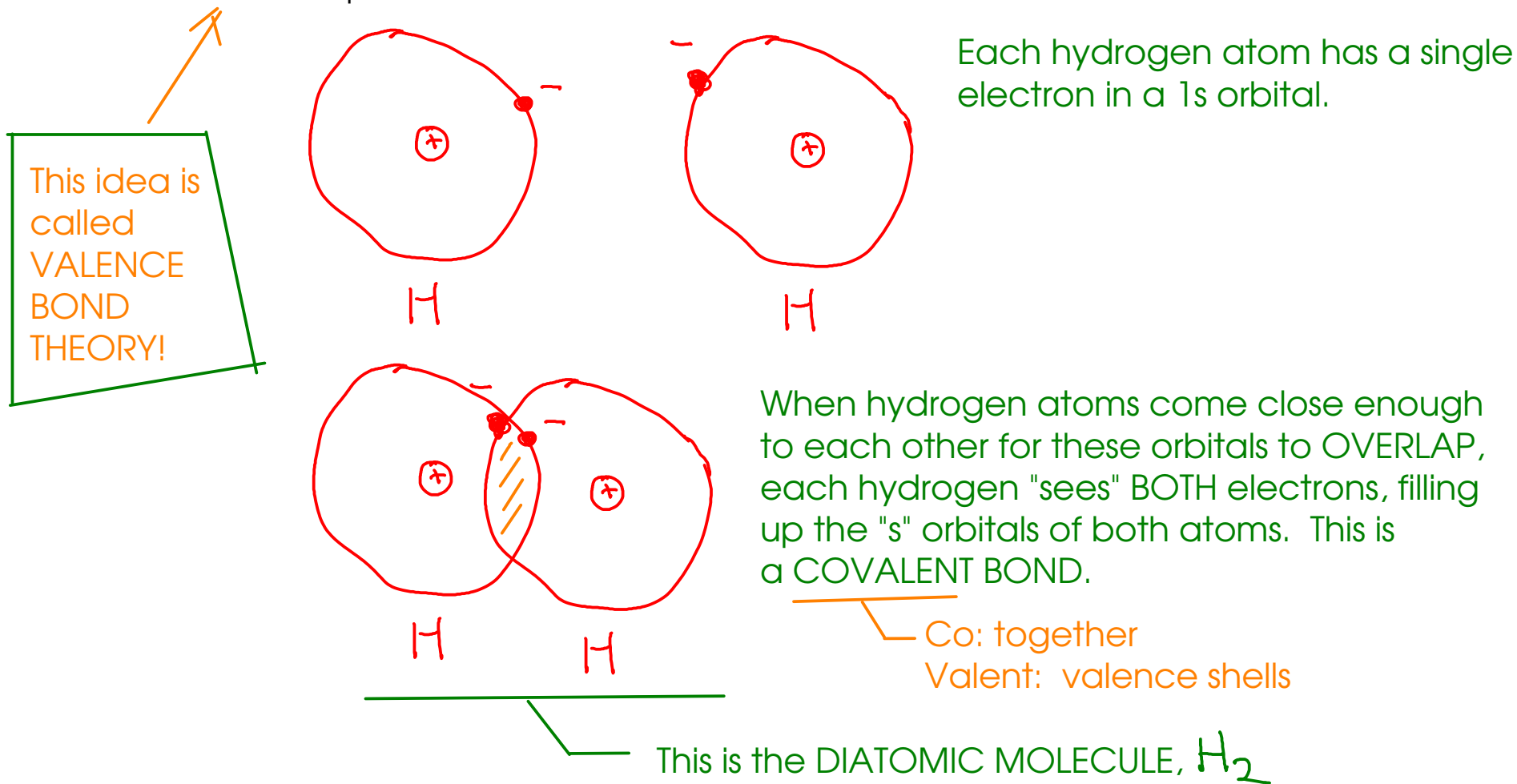


... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

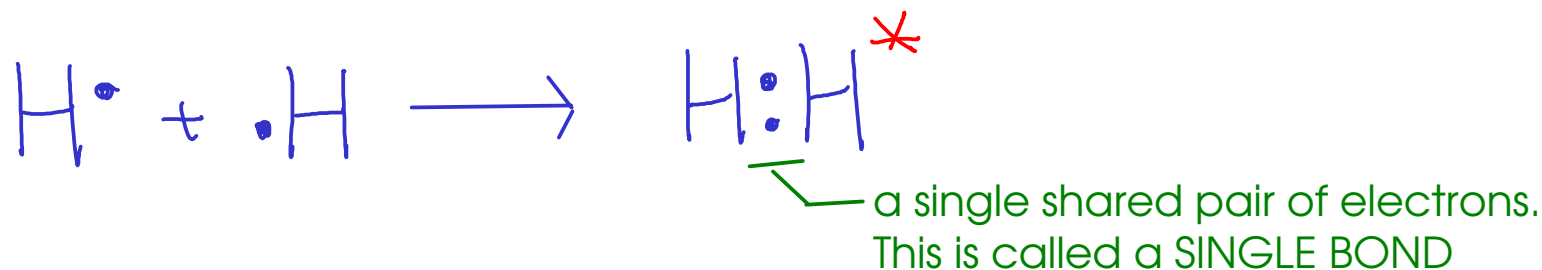
MOLECULAR COMPOUNDS

- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

How might atoms SHARE electrons? By coming together close enough so that their atomic ORBITALS overlap each other:



... so how would this look using dot notation?

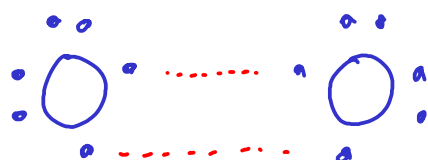


In dot structures, SHARED PAIRS of electrons are often written as DASHES to make the structures look neater.



* Why doesn't hydrogen end up with eight electrons? Because hydrogen has only the first shell, which contains only a single "s" subshell (NO "p" subshell). This "s" subshell is full with two electrons, and that's all hydrogen needs to get.

Let's look at OXYGEN ...



We know that oxygen exists in air as the diatomic molecule O_2



The oxygen atoms share TWO pairs of electrons. This is called a DOUBLE BOND

OR

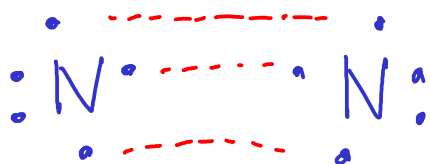


Each oxygen atom has a share in eight electrons!

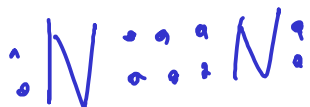
A few notes on the double bond:

- ① - For atoms to share more than one pair of electrons, they have to move closer to one another than they would if they were only sharing one pair of electrons. This BOND DISTANCE is measurable!
- ② - It takes more energy to break a double bond between two atoms than it would to break a single bond between the same two atoms. This BOND ENERGY is also measurable!

Let's look at NITROGEN ...



We know that nitrogen exists in air as the diatomic molecule N_2



The nitrogen atoms share THREE pairs of electrons. This is called a TRIPLE BOND

OR



Nitrogen gas is fairly inert ... it's hard to break the triple bond in nitrogen gas apart!

A few notes on the triple bond:

- ① - For atoms to share three pairs of electrons, they have to move closer to one another than they would if they were sharing one or two pairs of electrons. Triple bonds have the shortest BOND DISTANCE of all covalent bonds.
- ② - It takes more energy to break a triple bond between two atoms than it would to break either a single or double bond between the same two atoms. The triple bond has the largest BOND ENERGY of all three kinds of covalent bonds.

SO FAR, we've seen that ...

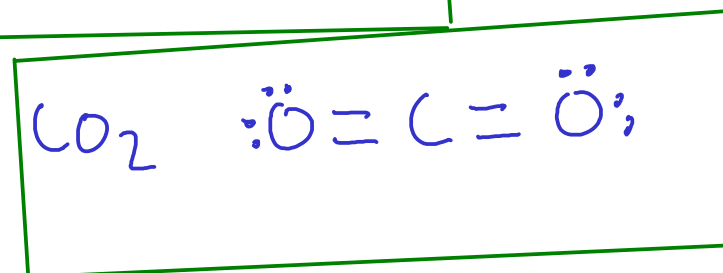
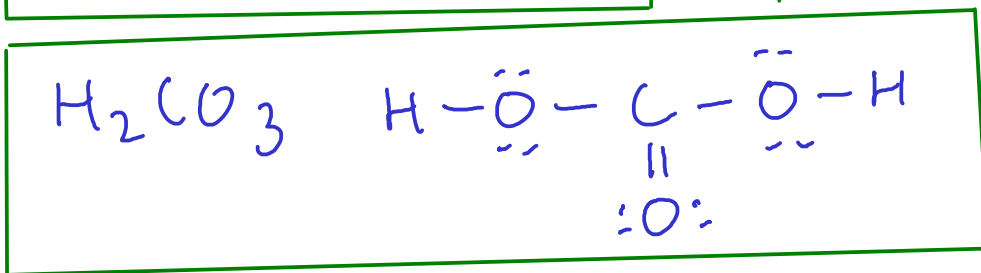
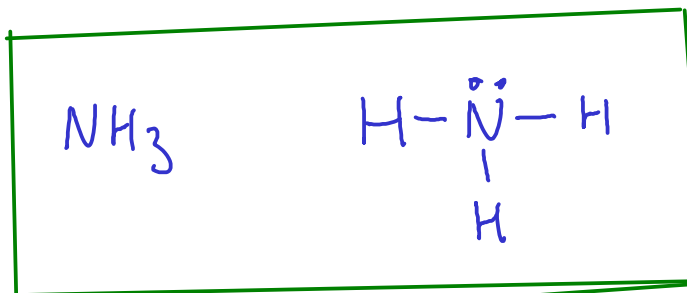
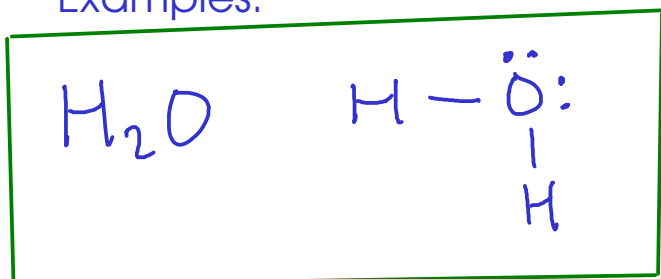
- ① Atoms may share one, two, or three pairs of electrons with each other.
- ② Atoms will usually share enough electrons so that each atom ends up with a share in EIGHT electrons - the "octet rule"

- HYDROGEN will only end up with two electrons!

- Some other atoms may end up with more or less than eight electrons. Exceptions to the octet rule are covered in Chapter 9.

NOW, how could we come up with dot structures for some more complicated (and therefore, more interesting) molecules?

Examples:



DRAWING DOT STRUCTURES FOR SIMPLE MOLECULES

① Count valence electrons

② Pick central atom and draw skeletal structure

- central atom is usually the one that needs to gain the most electrons!

- skeletal structure has all atoms connected to center with single bonds

③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

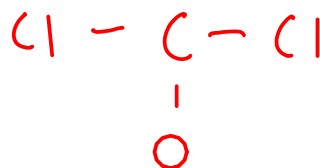
④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.



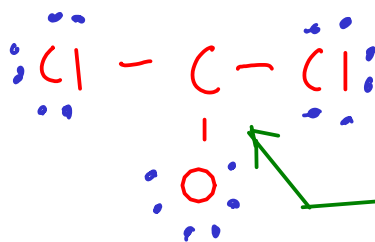
$$\begin{aligned} \text{C} &: 1 \times 4 \\ \text{O} &: 1 \times 6 \\ \text{Cl} &: 2 \times 7 = 14 \end{aligned}$$

24 valence e^-

Pick C as central atom, since it needs to gain more electrons than either of the other elements!



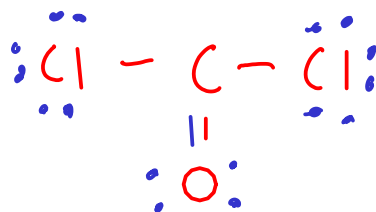
Distribute the other electrons ...
(We've used six so far for bonds!)



... stop when you've used all 24 electrons!

Problem: The central carbon atom has a share in only SIX electrons (from the bonds)! We need to fix that!

We'll need to make one of our pairs of unshared electrons (lone pairs) into an additional bond ... but which one to use? Use OXYGEN, since it needs to gain two electrons anyway and is likely to form two bonds to do so!



Making a double bond "fixes" this structure.

① Count valence electrons

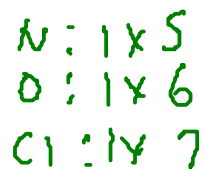
② Pick central atom and draw skeletal structure

- central atom is usually the one that needs to gain the most electrons!

- skeletal structure has all atoms connected to center with single bonds

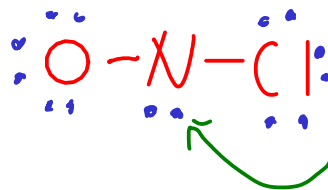
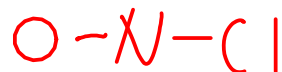
③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.



$$18 e^-$$

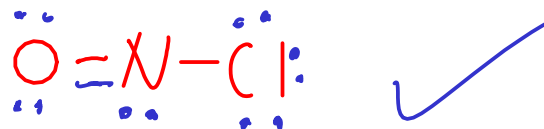
Pick NITROGEN for central atom. (Needs more electrons than O or Cl). Draw skeleton ...



Distribute remaining electrons ...

Last two electrons go on central NITROGEN...

Even with the lone pair, nitrogen still has a share in only SIX electrons ... Fix that with a double bond? Use oxygen again ... for the same reason as the last example.



① Count valence electrons

② Pick central atom and draw skeletal structure

- central atom is usually the one that needs to gain the most electrons!

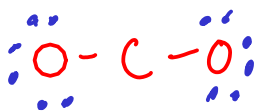
- skeletal structure has all atoms connected to center with single bonds



$$\begin{array}{r} \text{C: } 1 \times 4 \\ \text{O: } 2 \times 6 \\ \hline 16 e^- \end{array}$$



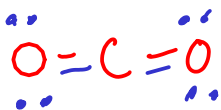
Choose CARBON for central atom.



Distribute remaining valence electrons. But CARBON has a share in only four?



Now six ...



A double bond with each oxygen atom "fixes" carbon dioxide!

③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.



What about this? This is not the likely structure because it makes two identical atoms (oxygen) in the same situation (bonding with carbon and no other atom) bond DIFFERENTLY. If elements are in fact chemically identical, that shouldn't happen. We should instead use the structure with two double bonds.

① Count valence electrons

② Pick central atom and draw skeletal structure

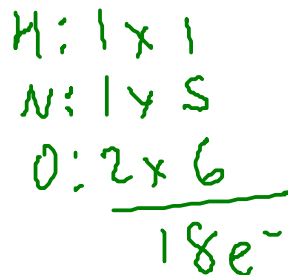
- central atom is usually the one that needs to gain the most electrons!
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③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). If not, make double or triple bonds.

HNO_2 "nitrous acid"

In oxyacids, the acidic hydrogen atoms are attached to OXYGEN atoms in the structure!



For any OXYACID, we MUST have a hydrogen atom bonded to an oxygen atom, so we put the hydrogen atom here!



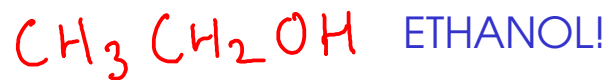
... but NITROGEN only has a share in six electrons. Fix!



Use the oxygen atom on the left to make the double bond. Why? The oxygen atom on the right has already formed two bonds (one with N, one with H). These two oxygen atoms are in DIFFERENT environments, so we don't expect them to necessarily bond the same way with nitrogen!

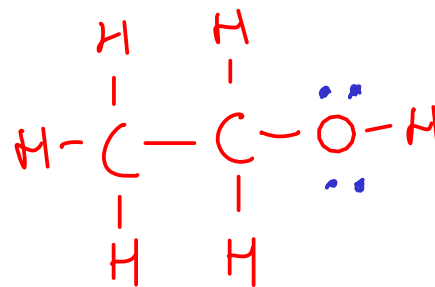
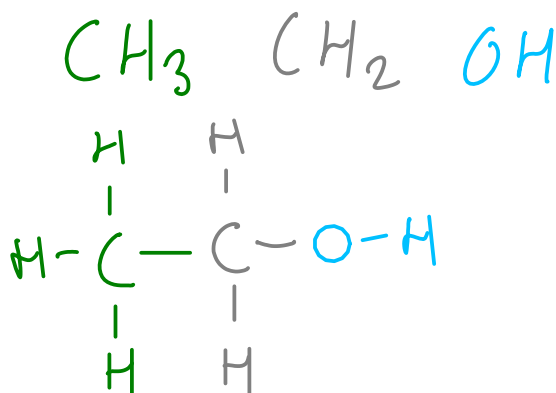
A DOT STRUCTURE FOR A LARGER MOLECULE

- ① Count valence electrons
- ② Pick central atom and draw skeletal structure
 - central atom is usually the one that needs to gain the most electrons!
 - skeletal structure has all atoms connected to center with single bonds
- ③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.
- ④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). If not, make double or triple bonds.



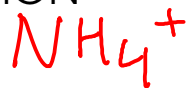
$$\begin{array}{l|l} \text{C} : 4 \times 2 = 8 & \\ \text{H} : 1 \times 6 = 6 & 20 \\ \text{O} : 6 \times 1 = 6 & \end{array}$$

This formula gives us a hint to the structure of ethanol. Ethanol has THREE central atoms chained together.



A DOT STRUCTURE FOR A POLYATOMIC ION

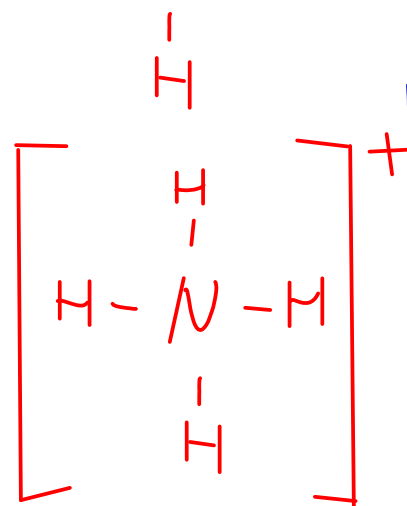
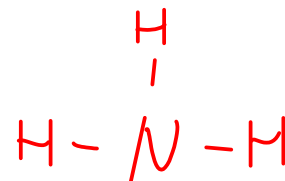
- ① Count valence electrons
- ② Pick central atom and draw skeletal structure
 - central atom is usually the one that needs to gain the most electrons!
 - skeletal structure has all atoms connected to center with single bonds
- ③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.
- ④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.



NINE valence electrons?
But we're used to using PAIRS for both bonding pairs and lone pairs of electrons. What's going on?

$$\begin{array}{r} \text{N: } 1 \times 5 \\ \text{H: } 4 \times 1 \\ \hline 9 \text{ valence } e^- \\ - 1 e^- \text{ (+1 charge)} \\ \hline 8 \text{ valence } e^- \end{array}$$

We need to subtract an electron from the total to get a +1 charge. (If the charge were negative, we'd need to add more electrons!)



We need to indicate the charge of this ion. Easy way - Draw brackets around structure and put charge in upper right.