

## PERIODIC TRENDS

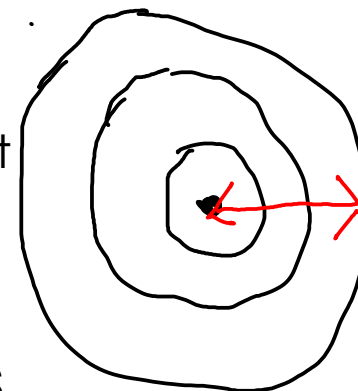
- Some properties of elements can be related to their positions on the periodic table.

### ATOMIC RADIUS

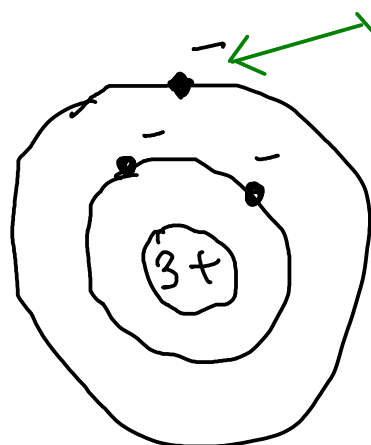
- The distance between the nucleus of the atoms and the outermost shell of the electron cloud.
- Relates to the size of the atom.
- As you go DOWN A GROUP ( ↓ ), the atomic radius INCREASES.

- Why? As you go down a group (from one period to the next) , you are ADDING SHELLS!

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

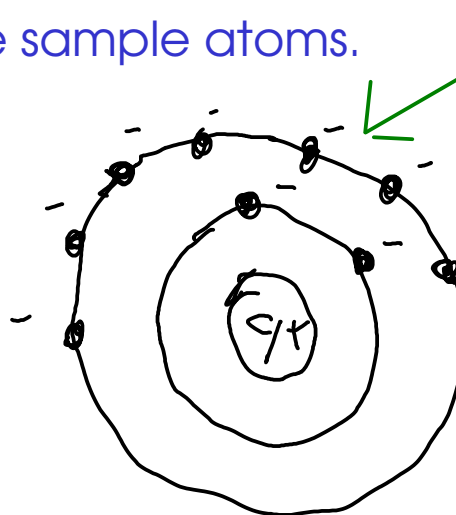


Why? Let's look at some sample atoms.



lithium

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



fluorine

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

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

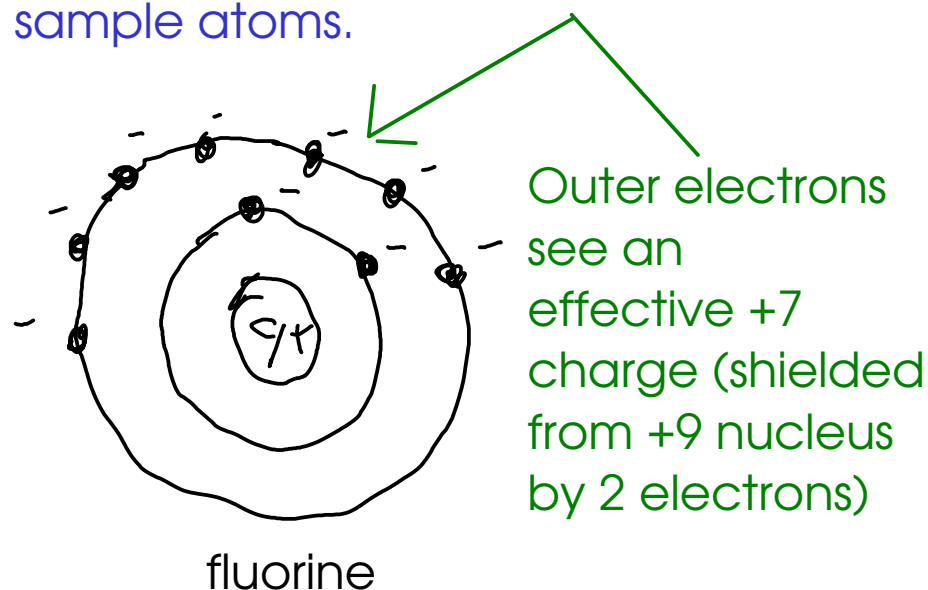
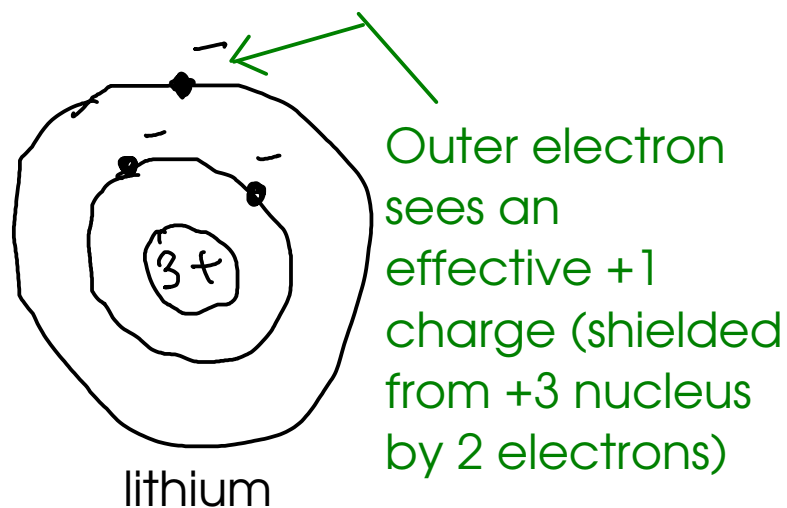
## IONIZATION ENERGY (or FIRST IONIZATION ENERGY)

- 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 ( $\downarrow$ ), 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 ( $\rightarrow$ ), 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.

## PERIODIC TRENDS IN A NUTSHELL

LARGER  
IONIZATION  
ENERGY

SMALLER  
RADIUS

IA																			VIIIA
H	IIA											IIIA	IVA	VA	VIA	VIIA			He
Li	Be											B	C	N	O	F			Ne
Na	Mg	IIIB	IVB	VB	VIB	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  
RADIUS

SMALLER  
IONIZATION  
ENERGY

## "octet rule"

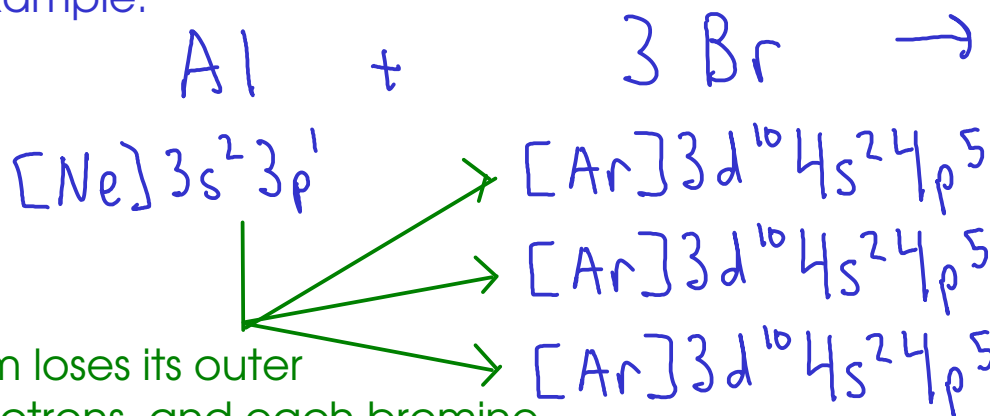
- a "rule of thumb" (NOT a scientific law) predicting how atoms will exchange or share electrons to form chemical compounds
- atoms will gain, lose, or share enough electrons so that they end up with full "s" and "p" subshells in their outermost shell.

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

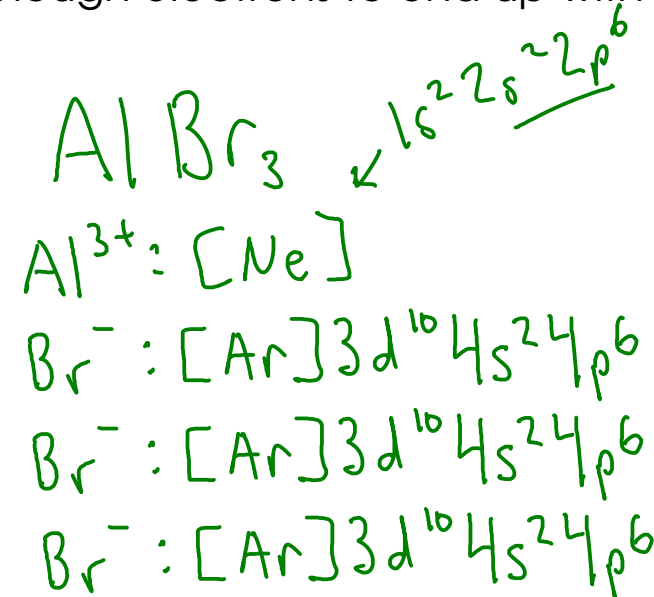
## IONIC COMPOUNDS

- 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!



↑ To save space, these electron configurations have been written with the "noble gas core" shortcut. Bromine's electron configuration is exactly like argon's - with the addition of some 3d, 4s, and 4p electrons!

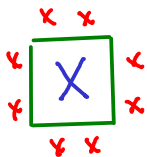
... but using electron configurations to describe how aluminum bromide forms is a bit cumbersome! Can we simplify the picture a bit?

## 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:



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"



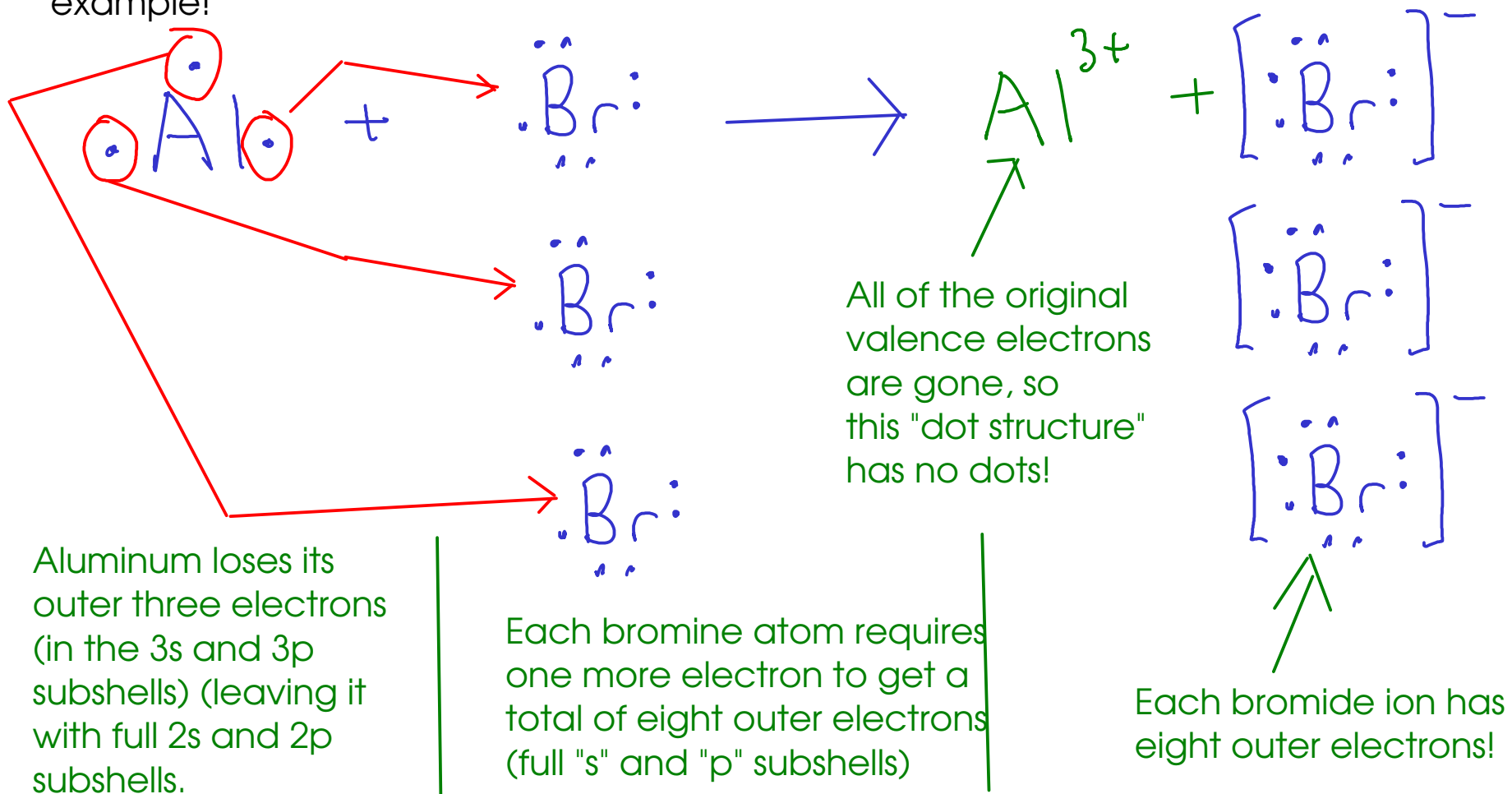
To draw a dot structure for an atom, you need to know HOW MANY valence electrons it has! You can determine this simply from the periodic table, WITHOUT writing the whole electron configuration!

The number of valence electrons equals the group number in the A/B group numbering system

1	H																		He
2	Li	Be										B	C	N	O	F			Ne
3	Na	Mg										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										

1 valence electron (Group IA)  
 2 valence electrons (Group IIA)  
 3 valence electrons (Group IIIA)  
 4 valence electrons (Group IVA)  
 5 valence electrons (Group VA)  
 6 valence electrons (Group VIA)  
 7 valence electrons (Group VIIA)  
 8 valence electrons (except helium!) (Group VIIIA)

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



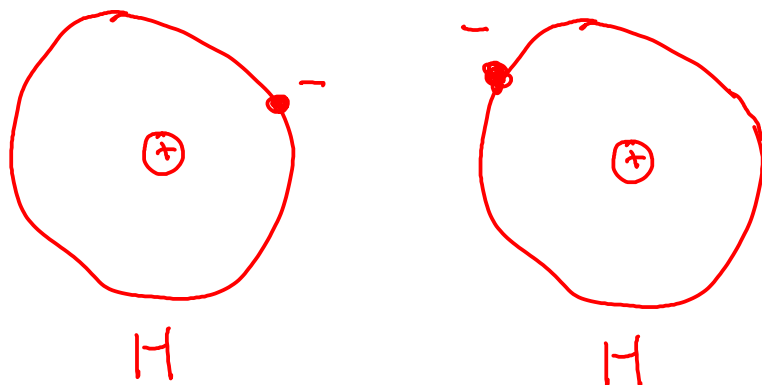
... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

This is an OXIDATION-REDUCTION (electron transfer) reaction. Dot notation makes the transfer of electrons very obvious.

## MOLECULAR COMPOUNDS

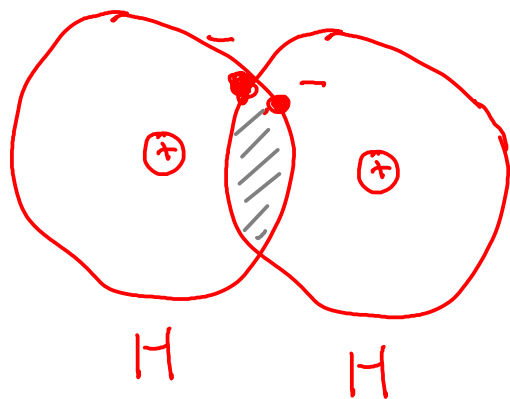
- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

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



Each hydrogen atom has a single electron in a 1s orbital.

This idea is called  
VALENCE BOND THEORY

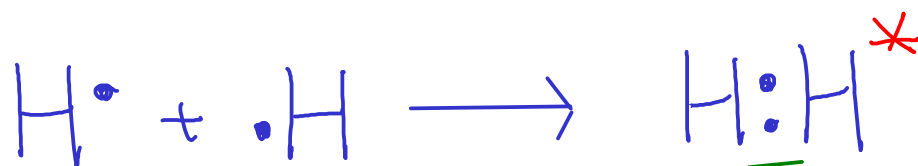


When hydrogen atoms come close enough to each other for these orbitals to OVERLAP, each hydrogen "sees" BOTH electrons, filling up the "s" orbitals of both atoms. This is a COVALENT BOND.

This is the DIATOMIC MOLECULE,  $H_2$



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



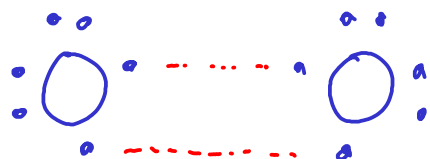
a shared pair of electrons. This is called a SINGLE BOND

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

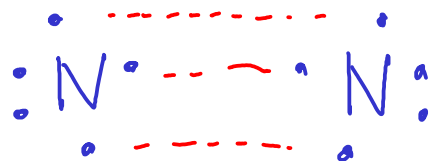


Each oxygen atom has a share in eight electrons!

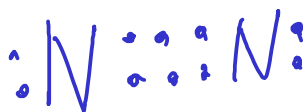
A few notes on the double bond:

- 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



The STABILITY of the nitrogen molecule (in other words, its relative inertness compared to molecules like hydrogen and oxygen) is probably due to the triple bond.

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.



## DRAWING DOT STRUCTURES FOR SIMPLE MOLECULES

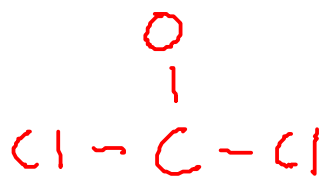
① Count valence electrons

② Pick central atom and draw skeletal structure

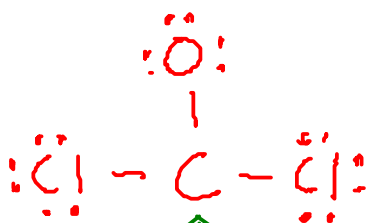
- central atom is usually the one that needs to gain the most electrons!
- skeletal structure has all atoms connected to center with single bonds

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

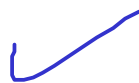
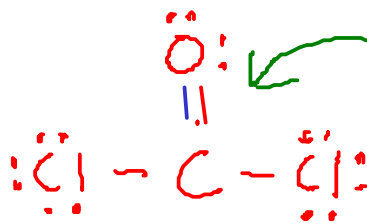
④ Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.



We choose CARBON as the central atom because it needs to gain four more electrons (more than oxygen or chlorine)



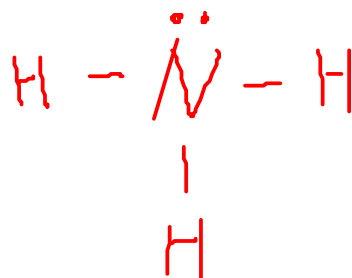
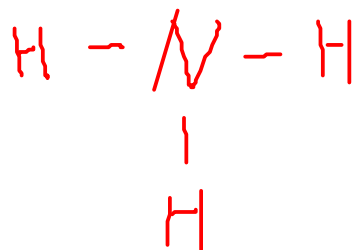
Problem: This carbon atom only has a share in SIX electrons.



Where to put the double bond? OXYGEN needed to gain TWO more electrons, so it's more likely to share two pairs of electrons than chlorine is (chlorine only needs one more electron)

$$\begin{array}{l} \text{C} : 4 \\ \text{O} : 6 \\ \text{Cl} : 2 \times 7 = 14 \\ \hline 24 \text{ electrons} \end{array}$$

- 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.



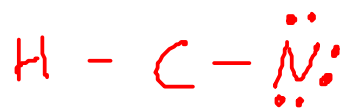
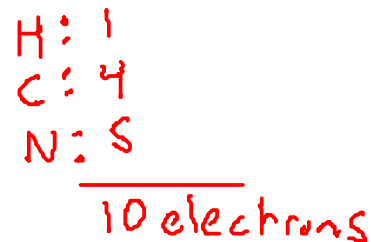
$$\begin{array}{r} \text{N: } 5 \\ \text{H: } 3 \times 1 = 3 \\ \hline 8 \end{array}$$

HYDROGEN will never be a central atom, since all it can do is gain one more electron by forming a single bond!

We put the remaining pair of electrons on NITROGEN, since hydrogen is "full" with only two electrons.

(Hydrogen's outer shell is the first shell - which has only a single "s" subshell: two electrons)

- 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.



... but CARBON has a share in only FOUR electrons.



... now carbon has a share in SIX electrons



Finally, with the triple bond between nitrogen and carbon, carbon has a share in eight electrons.



## A DOT STRUCTURE FOR A LARGER MOLECULE

① Count valence electrons

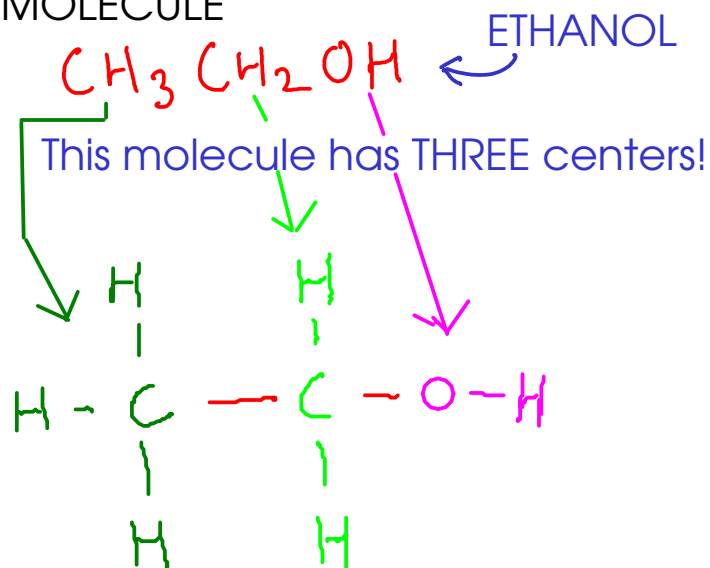
② Pick central atom and draw skeletal structure

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

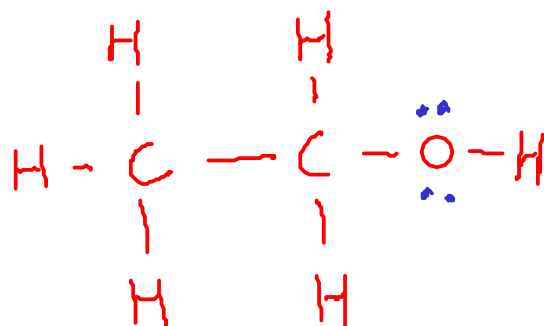
- skeletal structure has all atoms connected to center with single bonds

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.



$$\begin{array}{r} \text{C } 2 \times 4 = 8 \\ \text{H } 6 \times 1 = 6 \\ \text{O } 1 \times 6 = 6 \\ \hline 20 \text{ electrons} \end{array}$$



The remaining four electrons go onto the oxygen atom, since the hydrogens and carbons are "full".

Compare to WATER:



The ALCOHOLS (like ethanol, methanol, and isopropanol) are similar in structure to WATER.

Small-molecule alcohols dissolve very well in water due to this structural similarity.



## A DOT STRUCTURE FOR A POLYATOMIC ION

① Count valence electrons

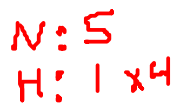
② Pick central atom and draw skeletal structure

- 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.

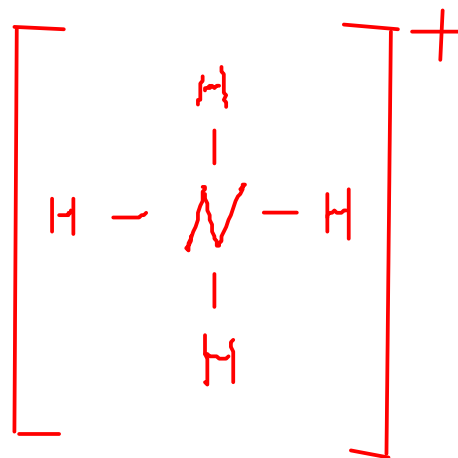
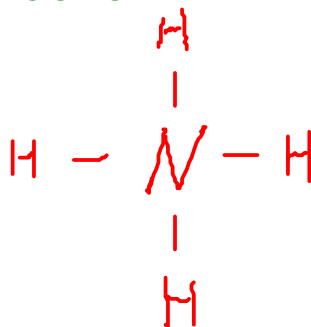


$$\underline{9}$$

$$-1$$

$$\underline{8 \text{ electrons}}$$

To get a +1 charge, the ammonium ion had to lose a single electron



We typically draw the structure of an ion in brackets, then indicate the charge at the upper right just like we do with a regular formula for an ion.