CHEMICAL BONDS

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

TWO TYPES OF CHEMICAL BOND

Type	Held together by...	Example
Ionic bonds	attractive forces between oppositely charged ions	sodium chloride
Covalent bonds	sharing of valence electrons between two atoms (sometimes more - "delocalized bonds")	water

Some compounds are held together by one type of bond, others (such as ionic compounds containing polyatomic ions) are held together by both!

Na_2CO_3

... 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 number describing how tightly an atom will hold bonded electrons.	Openstax p 346: Chart of electronegativities

... 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!

INCREASING ELECTRO-NEGATIVITY

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA
2	Li	Be											B	C	N	O	F
3	Na	Mg											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

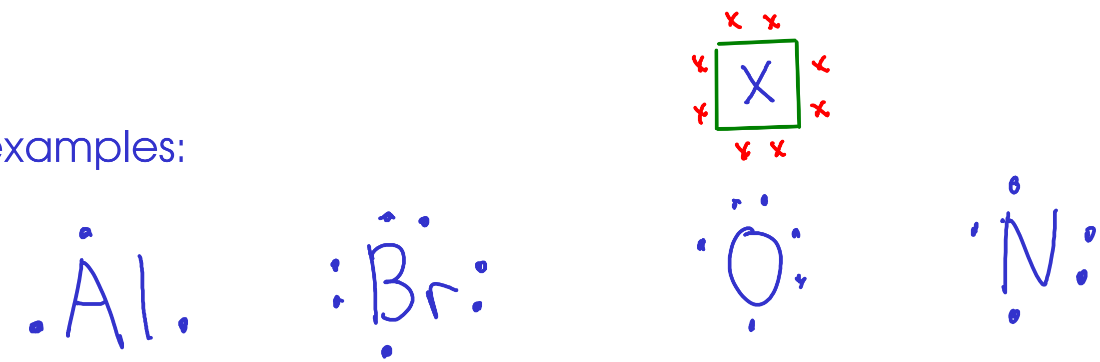
... so C-H bonds are NONPOLAR

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.

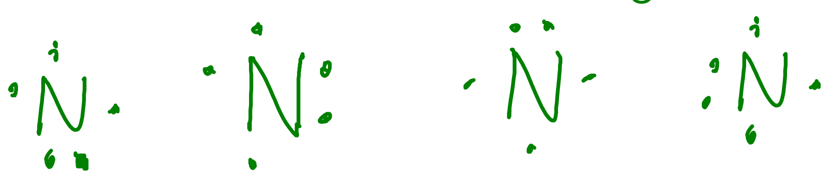
- To use electron-dot notation, put a dot for each valence electron around the atomic symbol. Put one dot on each "side" of the symbol (4 sides), then pair the dots for atoms that have more than four valence electrons.

examples:



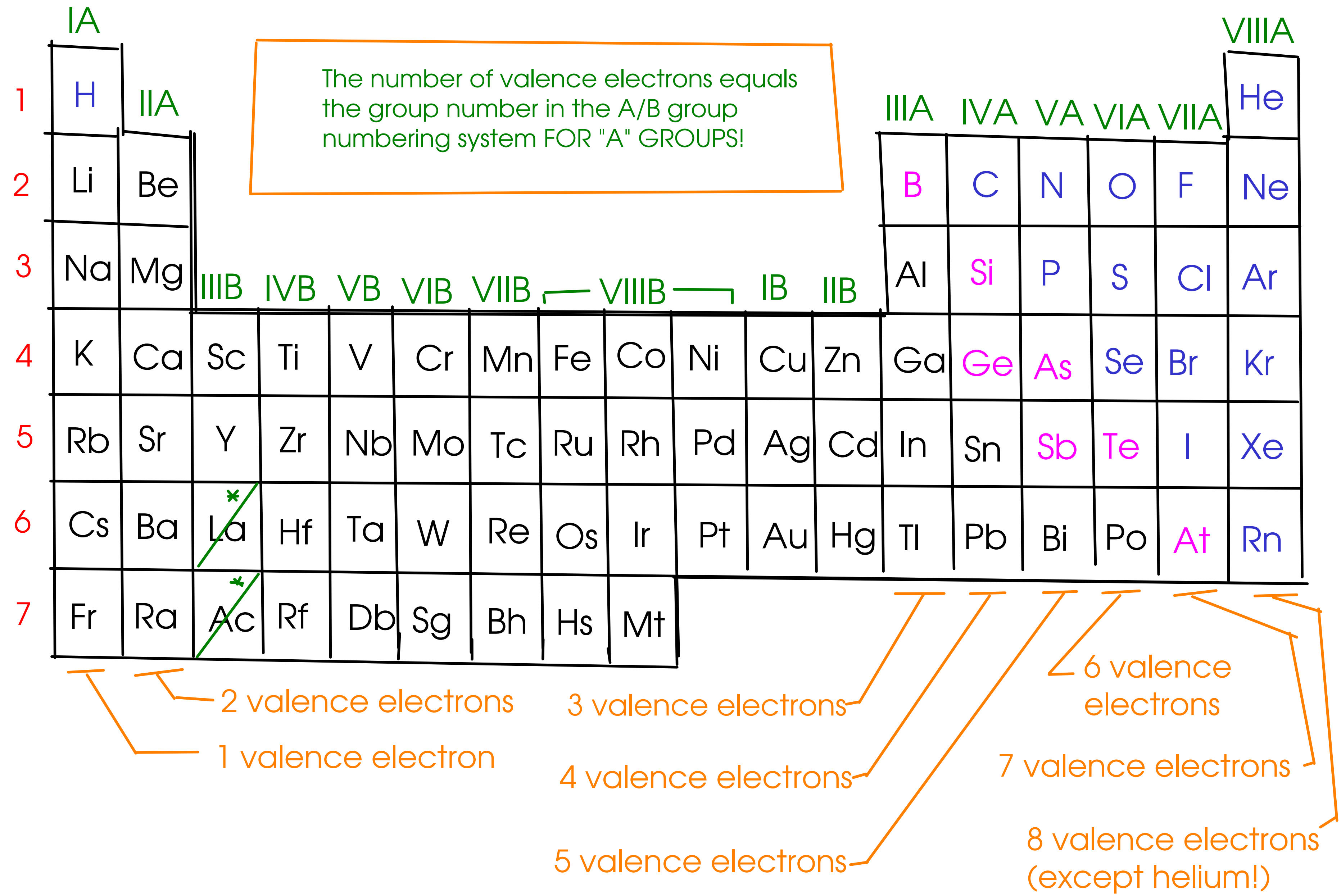
More examples

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"



. . . are all equivalent!

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.



DESCRIBING CHEMICAL BONDING

"octet rule"

- a "rule of thumb" (NOT a scienfitic 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 often GAIN or LOSE enough electrons to end up with full "s" and "p" subshells.

example:

$$Al + 3\,Br \rightarrow AlBr_3$$

$$\cdot\dot{Al}\cdot + \begin{matrix}\cdot\ddot{Br}\!:\\ \cdot\ddot{Br}\!:\\ \cdot\ddot{Br}\!:\end{matrix} \longrightarrow Al^{3+} + \begin{matrix}[:\ddot{Br}:]^-\\ [:\ddot{Br}:]^-\\ [:\ddot{Br}:]^-\end{matrix}$$

All of the original valence electrons are gone, so this "dot structure" has no dots!

Aluminum loses its outer three electrons (in the 3s and 3p subshells) (leaving it with full 2s and 2p subshells.

Each bromine atom requires one more electron to get a total of eight outer electrons (full "s" and "p" subshells)

Each bromide ion has eight outer electrons!

Aluminum is oxidized!

Bromine is reduced!

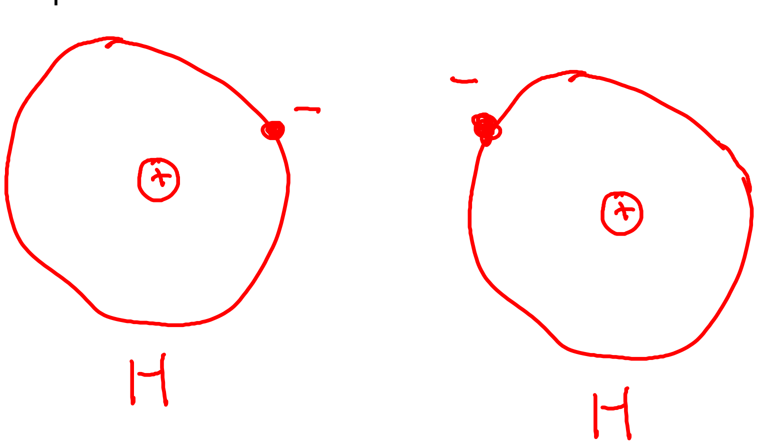
Redox reactions are much clearer when seen using dot notation!

MOLECULAR COMPOUNDS

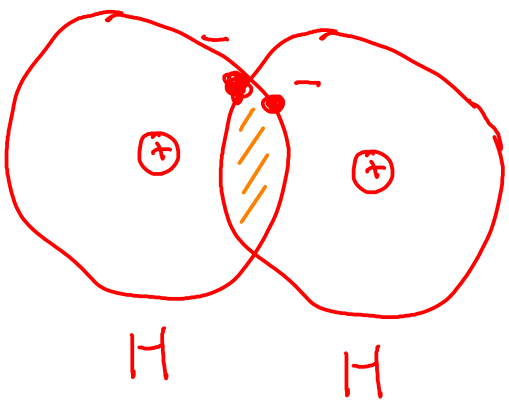
- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

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

This idea is called VALENCE BOND THEORY!



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



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.

Co: together
Valent: valence shells

This is the DIATOMIC MOLECULE, H_2

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

$$H\cdot + \cdot H \longrightarrow H:H\ ^{*}$$

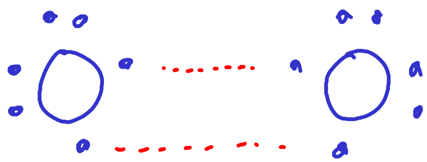
a single shared pair of electrons.
This is called a SINGLE BOND

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

$H:H$ becomes $H-H$

* 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.

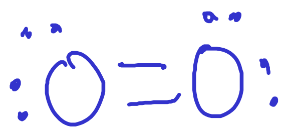
Let's look at OXYGEN ...



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



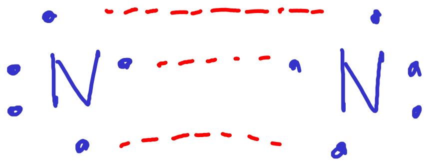
Each oxygen atom has a share in eight electrons!

A few notes on the double bond:

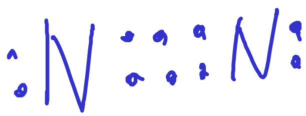
1 - For atoms to share more than one pair of electrons, they have to move closer to one another than they would if they were only sharing one pair of electrons. This BOND LENGTH is measurable!

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 STRENGTH 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 LENGTH 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 STRENGTH 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 a single other atom.

2. Atoms will usually share enough electrons so that each atom ends up with a share in EIGHT electrons - the "octet rule"

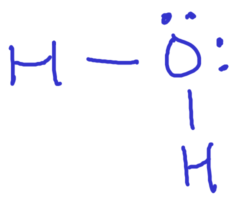
   - HYDROGEN will only end up with two electrons!

   - Some other atoms may end up with more or less than eight electrons.

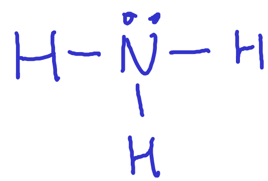
NOW, how could we come up with dot structures for some more complicated (and therefore, more interesting) molecules?

Examples:

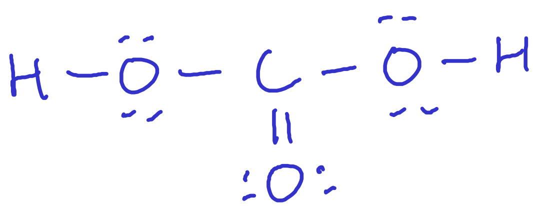
H_2O



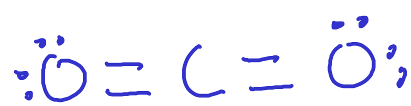
NH_3



H_2CO_3



CO_2



DRAWING DOT STRUCTURES FOR SIMPLE MOLECULES

$COCl_2$

C: 1 x 4
O: 1 x 6
Cl: 2 x 7 = 14

24 valence electrons

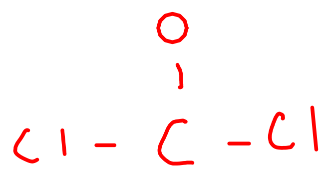
1. Count valence electrons

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

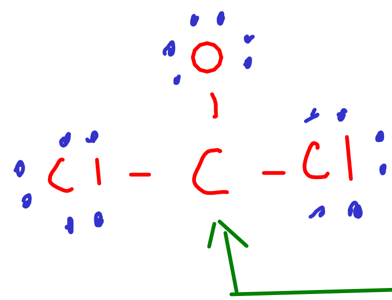
3. Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

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

Choose CARBON as the central atom since it needs to gain more outer shell electrons than either O or Cl. Since it needs more electrons, it will probably form more bonds!

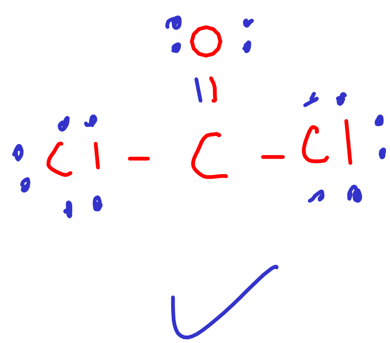


<- Skeletal structure



<- distributed remaining electrons, stopped at 24 (our count of total valence electrons)

Only 6 valence electrons!



Make a double bond, but with WHICH other atom? O or Cl? Pick by looking at the original count of valence electrons. Atoms that need more electrons form more bonds! O needs more, should form more 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!

- 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.

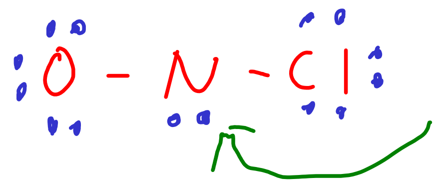
NOCl

N : 1 x 5
O : 1 x 6
Cl : 1 x 7

18 e^-

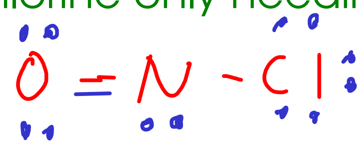
Pick NITROGEN as the central atom, since it needs more valence electrons that O or Cl.

O - N - Cl <- Skeletal structure



Outer atoms were full, so last pair goes on NITROGEN.

Problem: NITROGEN still has a share in only six valence electrons. Make a double bond, but where?
Pick OXYGEN since it needed two electrons compared to chlorine only needing one.



Adding a double bond between O and N 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.

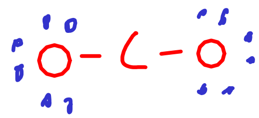
(4) Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.

CO_2

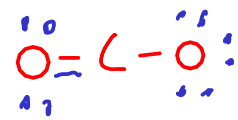
C: 1 x 4
O: 2 x 6

$16e^-$

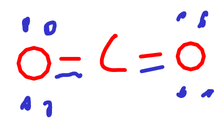
Pick C as central atom. Needs more valence electrons than O!



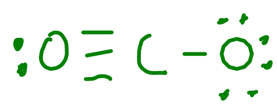
C only has a share in FOUR valence electrons!



Now C has six!



Making a second double bond with the other O fixes this structure!



What about this one?

This structure shows two atoms of the same element put into the same chemical situation (each bonding to one carbon atom and nothing else), yet bonding very differently - sharing different numbers of electrons! This shouldn't happen because the two oxygen atoms are the same element and should have bonded the same way!

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

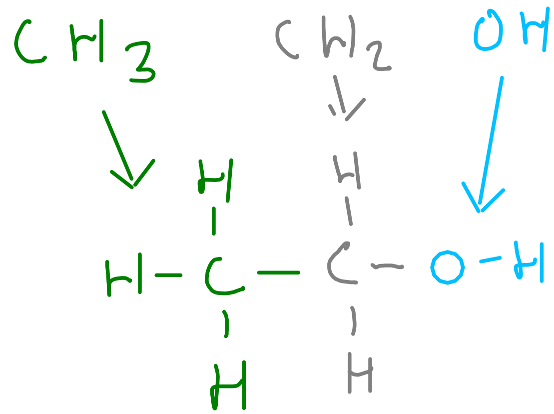
3. Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

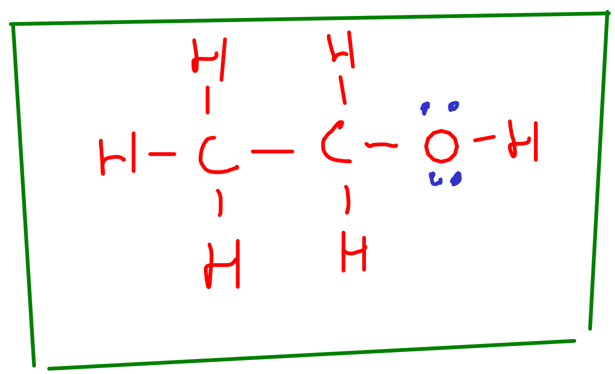
4. Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.

CH_3CH_2OH ETHANOL!

C : 4 x 2 = 8
H : 1 x 6 = 6
O : 6 x 1 = 6
| 20

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.

NH_4^+

N: 1x5
H: 4x1

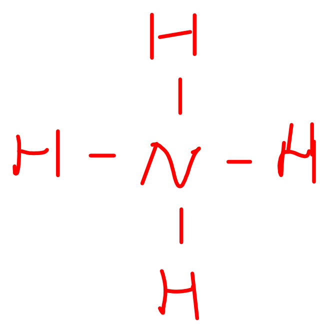
9 valence e^-

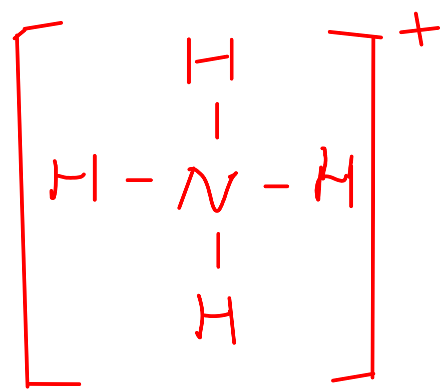
$-1\ e^-$ (+1 charge)

$8\ e^-$

Since ammoniun ion has a +1 charge, that means it has lost an electron so we must adjust the valence electron count!

(If there's a negative charge, add electrons! One electron per unit charge.)





To indicate the charge, draw brackets around the structure and put the charge in the upper right corner!