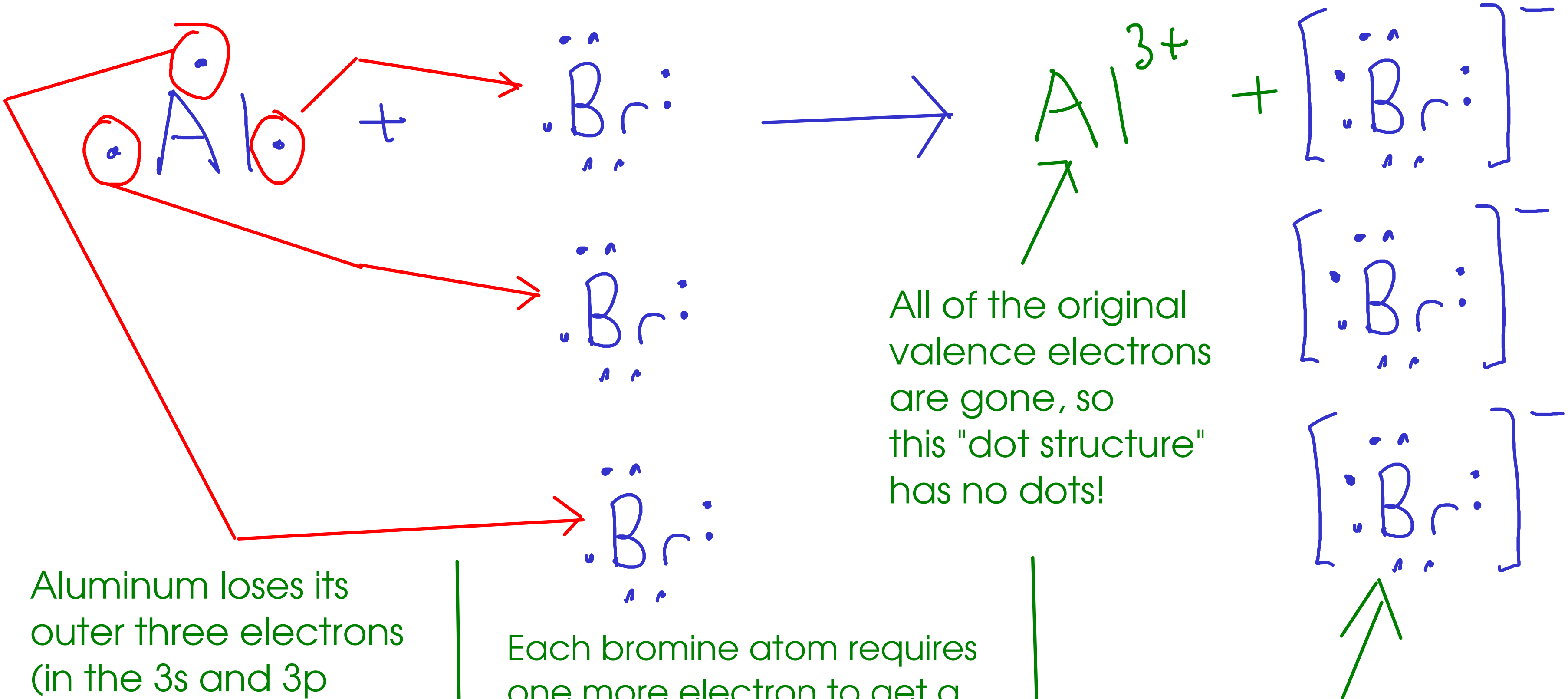
... but how do we use this to describe a reaction that produces ions? Let's look at our previous example!



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.

Aluminum is oxidized!

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

Bromine is reduced!

Each bromide ion has eight outer electrons!

Redox reactions are much clearer when seen using dot notation!

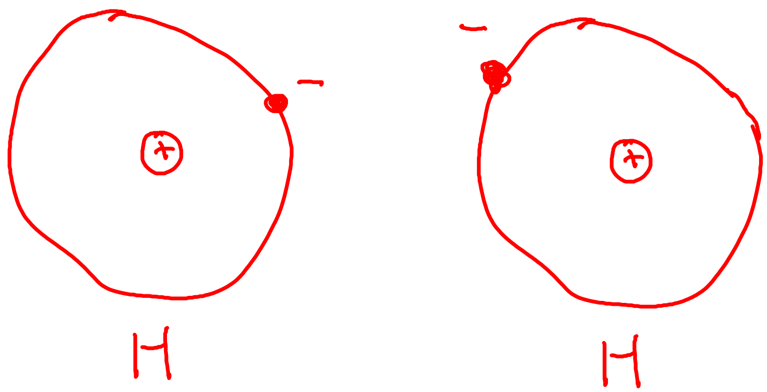
... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

MOLECULAR COMPOUNDS

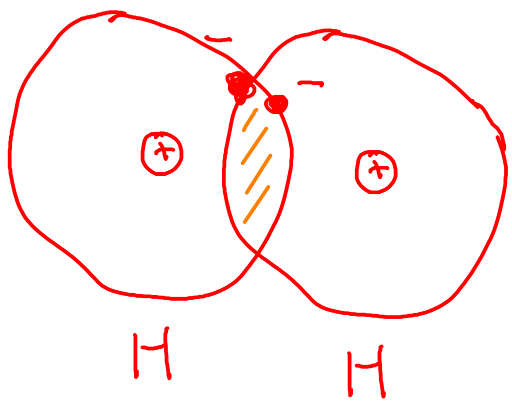
- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

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

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?

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

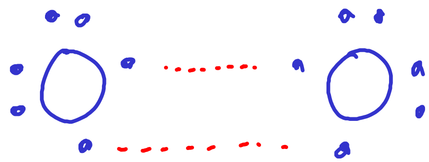
a single shared pair of electrons. This is called a SINGLE BOND

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.

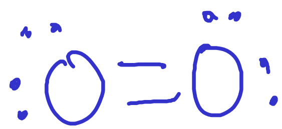
Let's look at OXYGEN ...



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



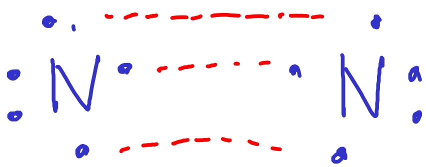
Each oxygen atom has a share in eight electrons!

A few notes on the double bond:

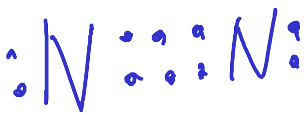
(1) - For atoms to share more than one pair of electrons, they have to move closer to one another than they would if they were only sharing one pair of electrons. This BOND DISTANCE is measurable!

(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 ...

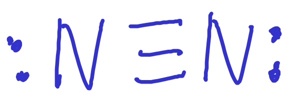


We know that nitrogen exists in air as the diatomic molecule N_2



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.

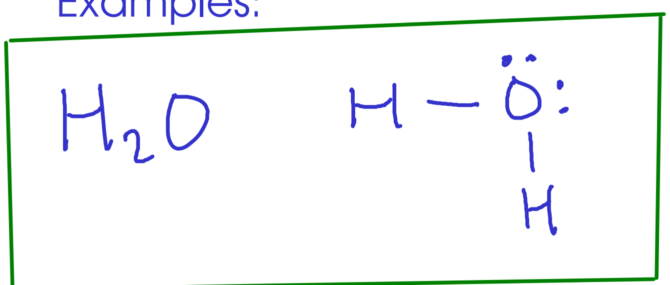
(2) Atoms will usually share enough electrons so that each atom ends up with a share in EIGHT electrons - the "octet rule"

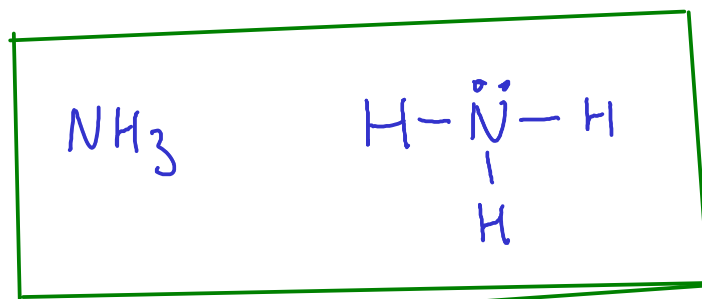
- HYDROGEN will only end up with two electrons!

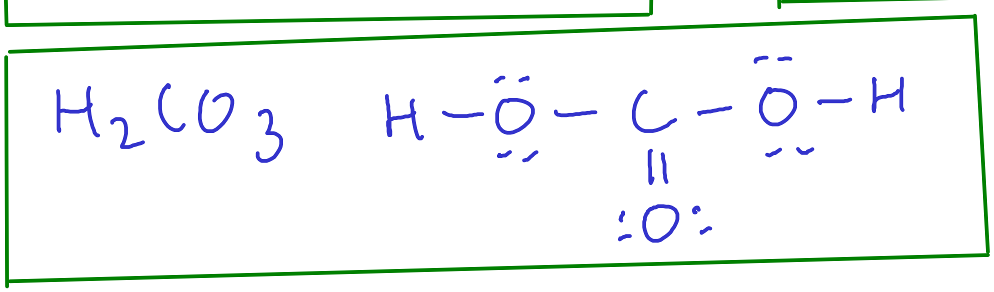
- Some other atoms may end up with more or less than eight electrons. Exceptions to the octet rule are covered in Chapter 9.

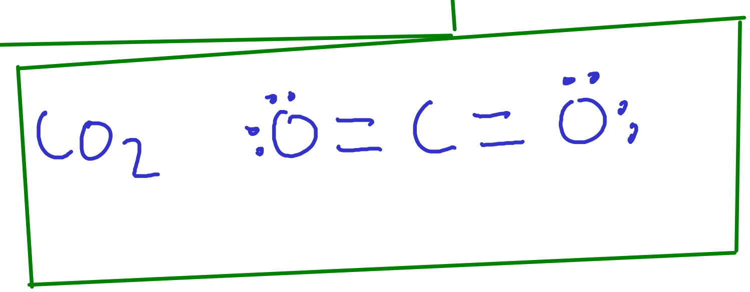
NOW, how could we come up with dot structures for some more complicated (and therefore, more interesting) molecules?

Examples:









(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

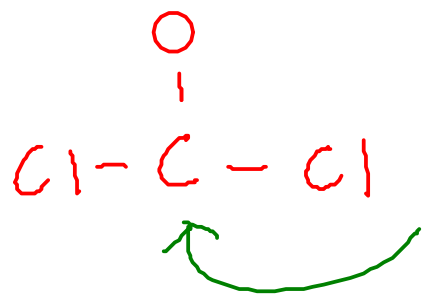
(3) Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

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

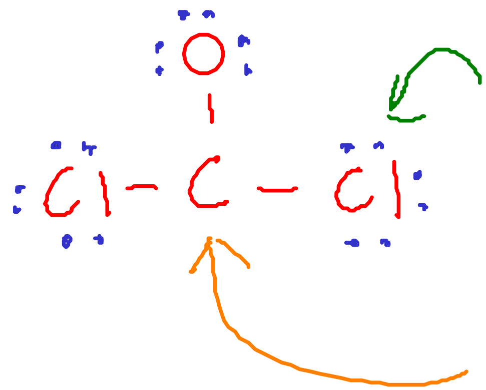
$COCl_2$

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

24 electrons



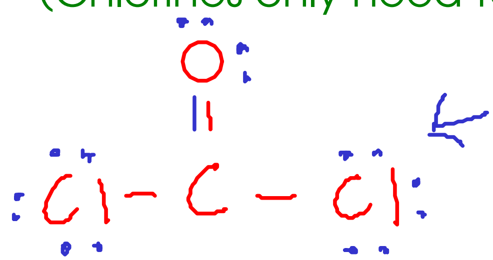
Choose CARBON as the central atom, since it needs to gain more electrons than the others.



Distribute remaining electrons. We stop here because we have used all 24 electrons.

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

We'll pick OXYGEN to share two pairs of electrons. Remembe that oxygen needed to gain two more electrons to get a share in eight, so oxygen should be able to share two electrons of its own to get those two more electrons! (Chlorines only need to gain one more electron...)



This structure looks better. All atoms have a share in eight 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

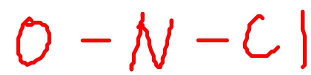
3. Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

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

NOCl

N : 1x5
O : 1x6
Cl : 1x7

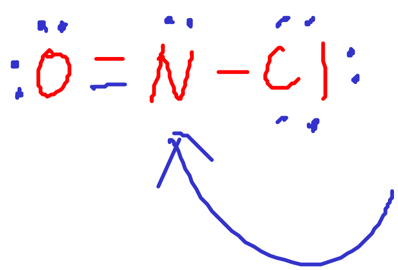
18 electrons



We pick NTROGEN as our central atom, since it needs three more electrons (more than O or Cl)

We put the last two electrons onto the central atom, since the outer atoms already have a share in eight...

... but NITROGEN has a share in only six electrons. We solve this the same way we solved the issue with the previous structure. Make a double bond with oxygen.



Using a doubl;e bond between oxygen and carbon 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.

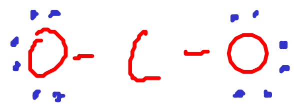
(4) Check octet rule - each atom should have a share in 8 electrons (H gets 2). if not, make double or triple bonds.

CO_2

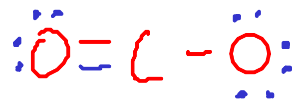
C : 4

O : 2×6 = 12

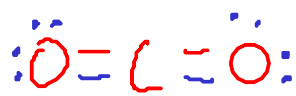
16 electrons



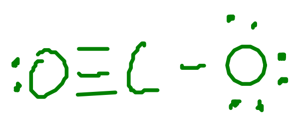
... but the carbon atom only has a share in four electrons!



... and now the carbon has a share in six electrons.



Adding a second double bond with the other oxygen atom fixes this structure and gives all atoms a share in eight electrons.



The two oxygen atoms are in identical chemical environments and SHOULD bond to the carbon in the same way, NOT in the arrangement we pictured on the left.

EXPERIMENTALLY, we find that the bond distances between the carbon and the oxygen atoms are the same. This supports the double-bond structure we drew originally, and does NOT agree with the triple/single-bond structure!

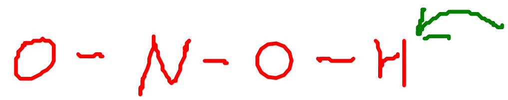
1. Count valence electrons
2. Pick central atom and draw skeletal structure
   - central atom is usually the one that needs to gain the most electrons!
   - skeletal structure has all atoms connected to center with single bonds
3. Distribute remaining valence electrons around structure, outer atoms first. 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.

HNO_2 "nitrous acid"

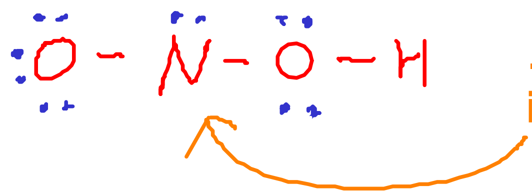
In oxyacids, the acidic hydrogen atoms are attached to OXYGEN atoms in the structure!

H : 1 x 1
N : 1 x 5
O : 2 x 6

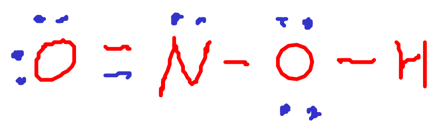
18 electrons



In an OXYACID, the acidic hydrogen must be attached to OXYGEN, so we draw the H here



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



Here, the oxygen on the left is in a different chemical environment than the one on the right, so we're not surprised that the bonding is different to the nitrogen atom. (Compare to the carbon dioxide structure we did earlier!)