

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</u> bonds	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

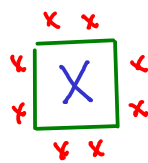
... but using electron configurations to describe how aluminum bromide forms is a bit cumbersome! Can we simplify the picture a bit?

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	IV A	V A	VIA	VII A	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg	IIIB	IVB	VB	VIB	VII B	VIII B	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

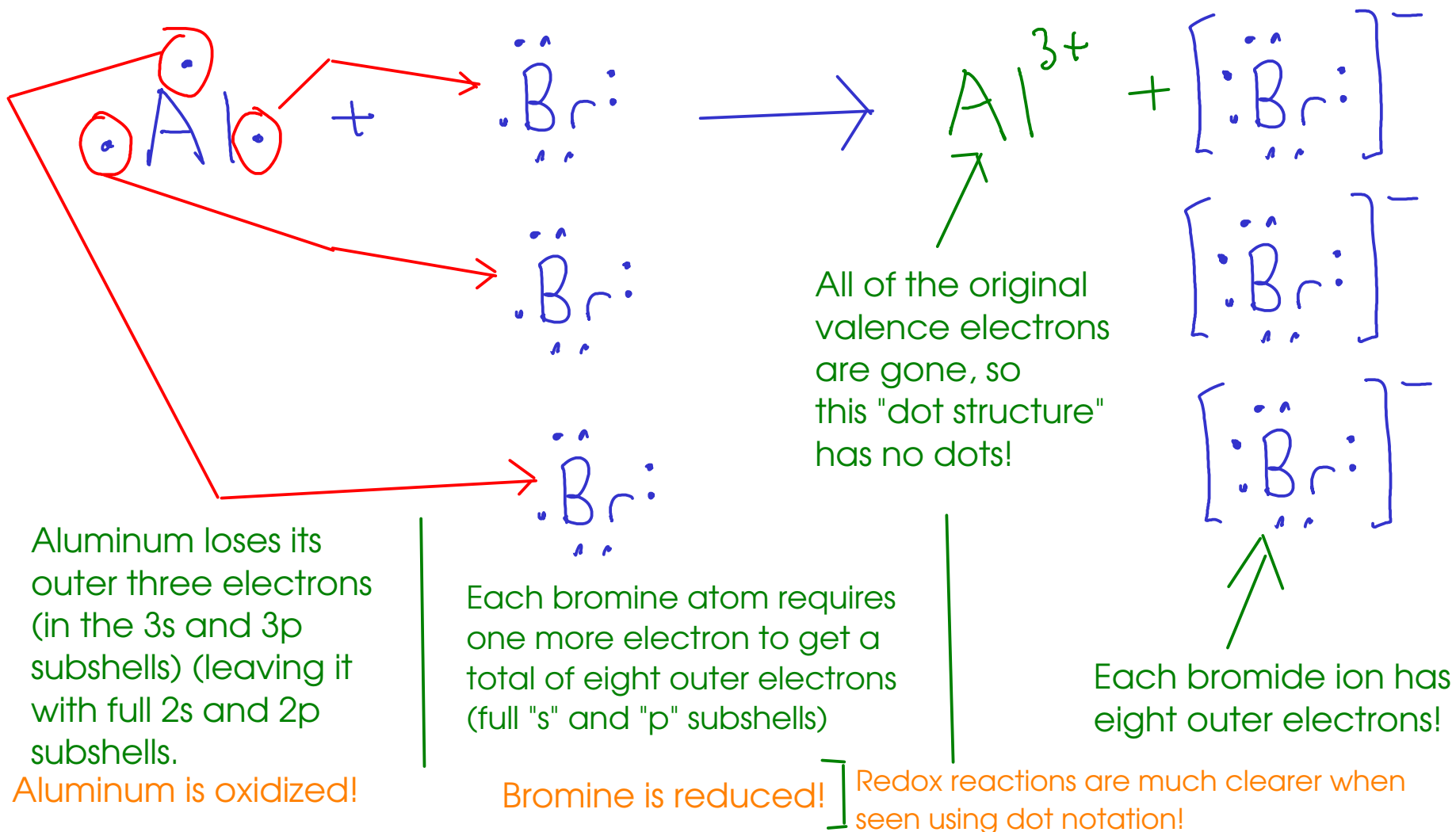
5 valence electrons

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!



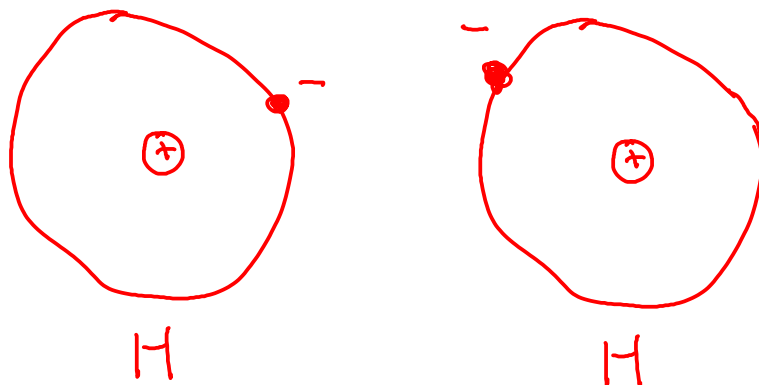
... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

MOLECULAR COMPOUNDS

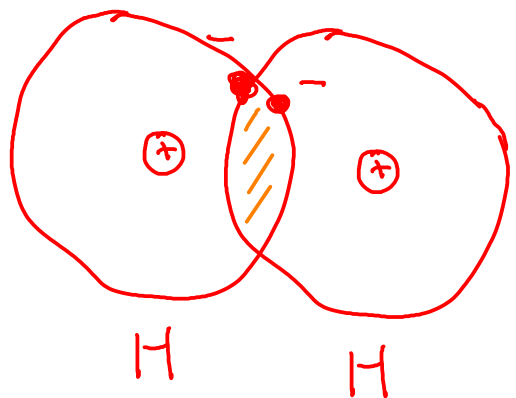
- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

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

This idea is called
VALENCE
BOND
THEORY!



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

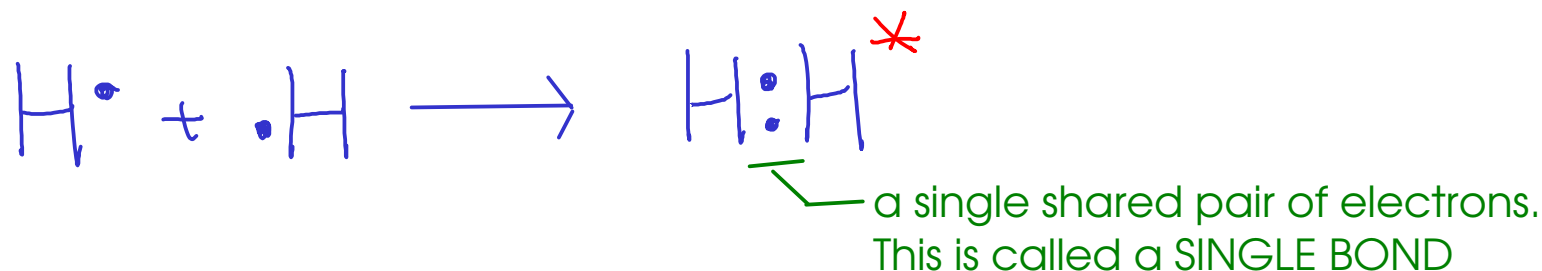


When hydrogen atoms come close enough to each other for these orbitals to OVERLAP, each hydrogen "sees" BOTH electrons, filling up the "s" orbitals of both atoms. This is a COVALENT BOND.

Co: together
Valent: valence shells

This is the DIATOMIC MOLECULE, H_2

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

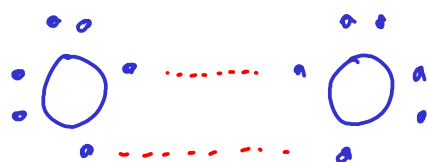


In dot structures, SHARED PAIRS of electrons are often written as DASHES to make the structures look neater.



* 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

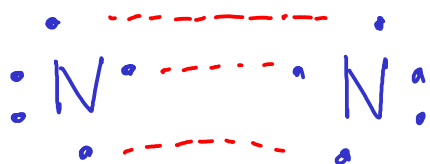


Each oxygen atom has a share in eight electrons!

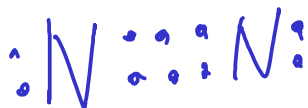
A few notes on the double bond:

- 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.

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.

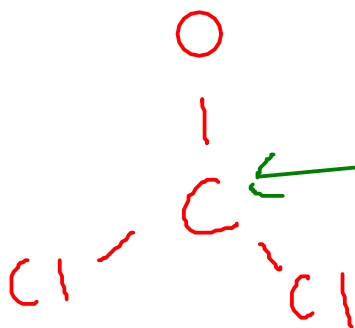


$$\text{C} : 1 \times 4$$

$$\text{O} : 1 \times 6$$

$$\text{Cl} : 2 \times 7 = 14$$

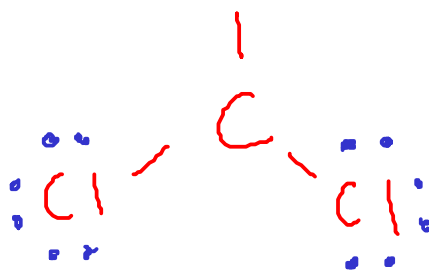
24 electrons



Choose CARBON as the central atom, as it needs to gain more electrons than either O or Cl.

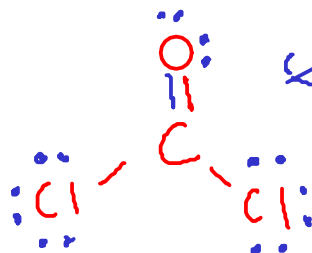


Distribute remaining electrons, stop when we run out!



... but the CARBON ATOM doesn't have enough electrons! How do we get more electrons for carbon?

Make a double bond! But with which other atom? Since oxygen originally needs two more electrons, it's more likely to form a double bond than chlorine.



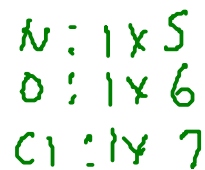
This structure will work ... as each atom has a share in eight electrons!

- Count valence electrons
- 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

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

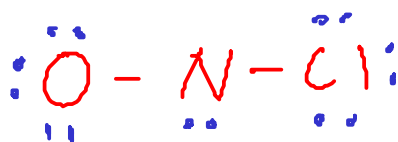
- Check octet rule - each atom should have a share in 8 electrons (H gets 2). If not, make double or triple bonds.



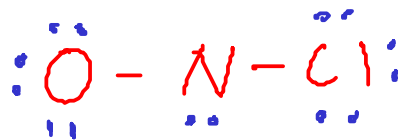
$$18 e^-$$



We'll pick NITROGEN as our central atom, since it needs more electrons than O or Cl.



We ran out of space on the outer atoms, so the last pair of electrons will go on the central nitrogen.



Even with the last pair on nitrogen, the atom still has a share in only six electrons. So we'll need to try a double bond!



As before, we choose oxygen for the double bond. Now all the atoms have a share in eight electrons!

① Count valence electrons

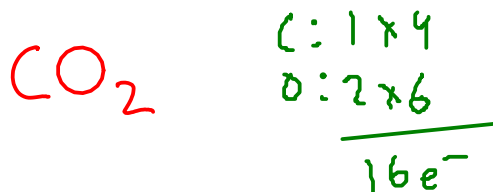
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③ Distribute remaining valence electrons around structure, outer atoms first. Follow octet rule until you run out of electrons.

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



$\text{O}-\text{C}-\text{O}$ Carbon is the central atom.

$\begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array} - \text{C} - \begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array}$... but carbon only has a share in four electrons!

$\begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array} = \text{C} - \begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array}$... now six.

$\begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array} = \text{C} = \begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array}$ Adding a second double bond gives carbon a share in eight valence electrons!

$\begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array} \equiv \text{C} - \begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \end{array}$ Why not this one? The two oxygen atoms are in the same environment and should bond the same way!

This structure says something that we can test - that the two oxygen atoms are different distances from the central carbon! Experimentally, we find that the oxygens are the same distance from the central carbon. Consistent with the double bond structure drawn earlier.