

- 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"

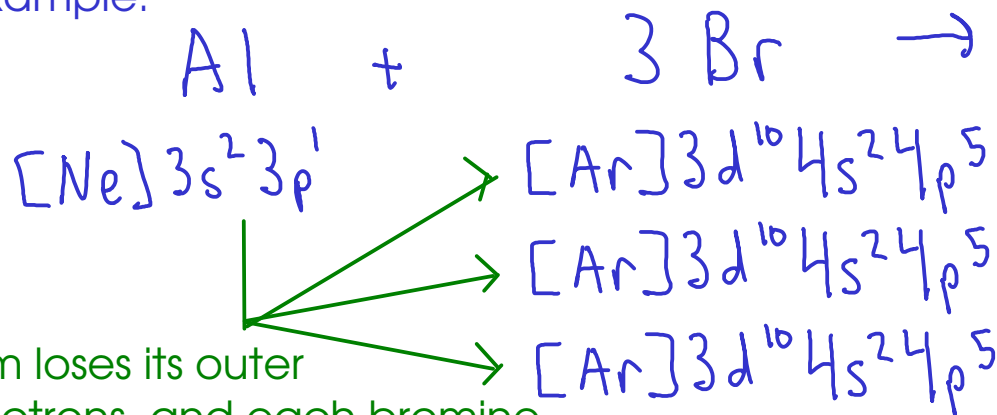
- a "rule of thumb" (NOT a scientific law) predicting how atoms will exchange or share electrons to form chemical compounds
- atoms will gain, lose, or share enough electrons so that they end up with full "s" and "p" subshells in their outermost shell.

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

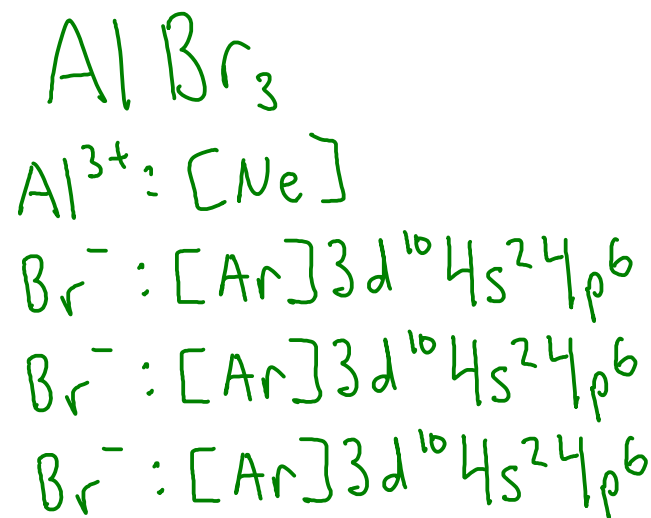
IONIC COMPOUNDS

- 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!



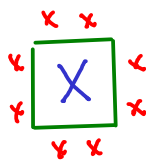
... but using electron configurations to describe how aluminum bromide forms is a bit cumbersome! Can we simplify the picture a bit?

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"



... 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, 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!

1	IA												VIII A					He
2	H	IIA											III A	IVA	VA	VIA	VIIA	Ne
3	Li	Be											B	C	N	O	F	Ar
4	Na	Mg	IIIB	IVB	VB	VIB	VII B	VIII B	IB	IIB	Al	Si	P	S	Cl	Ar		
5	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
6	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
7	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

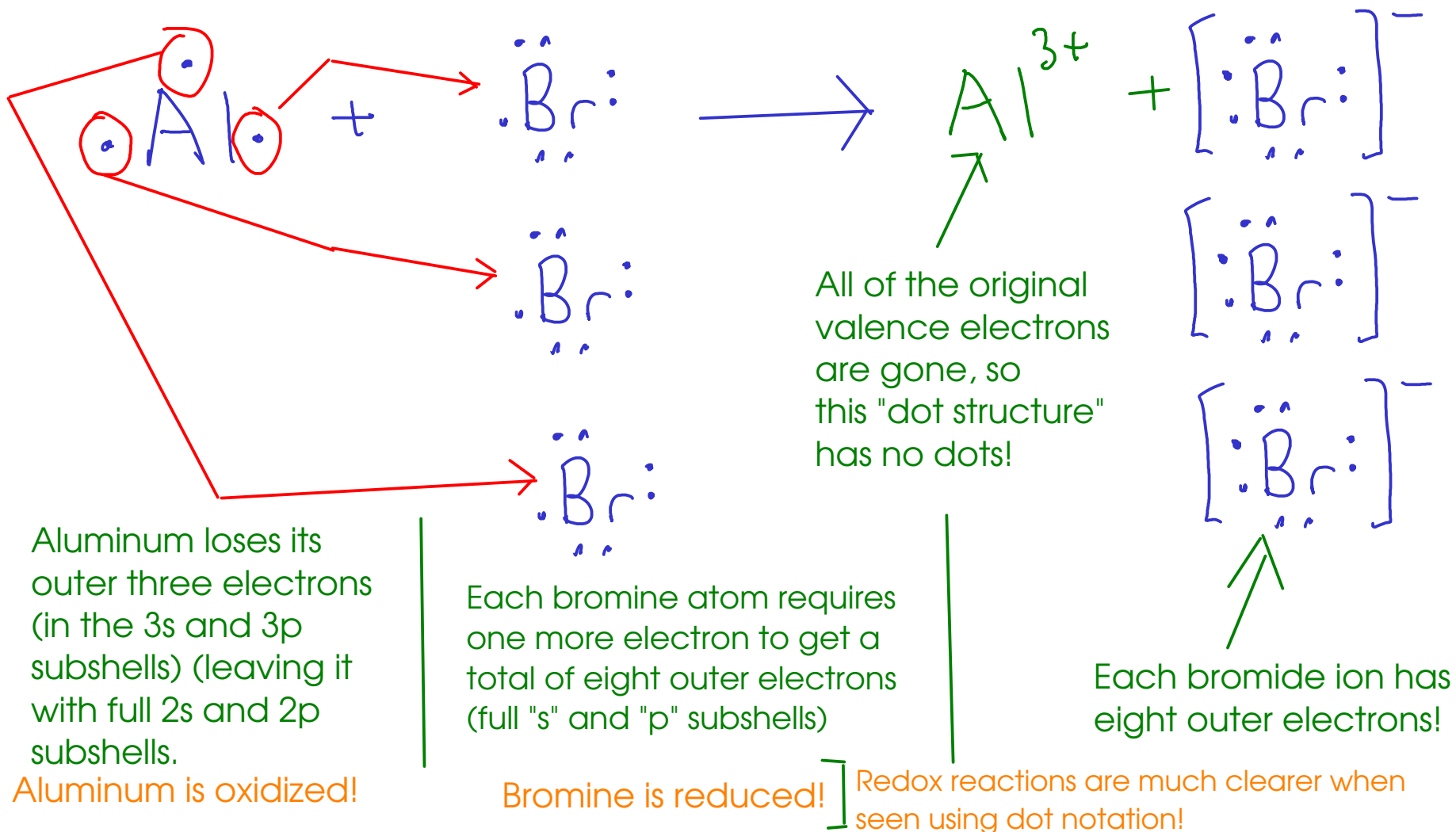
5 valence electrons

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!



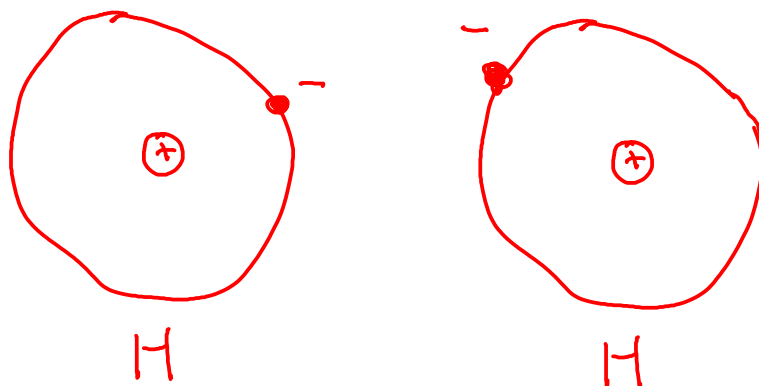
... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

MOLECULAR COMPOUNDS

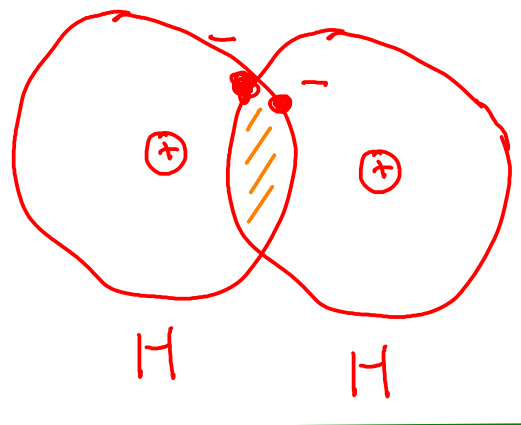
- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

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.

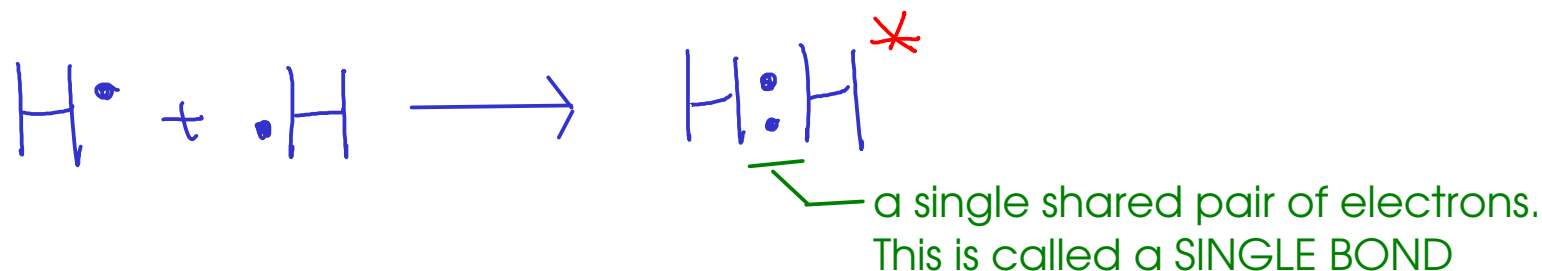


When hydrogen atoms come close enough to each other for these orbitals to OVERLAP, each hydrogen "sees" BOTH electrons, filling up the "s" orbitals of both atoms. This is a COVALENT BOND.

Co: together
Valent: valence shells

This is the DIATOMIC MOLECULE, H_2

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

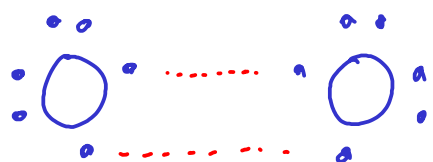


In dot structures, SHARED PAIRS of electrons are often written as DASHES to make the structures look neater.



* 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

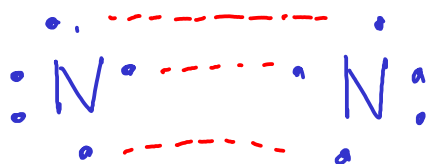


Each oxygen atom has a share in eight electrons!

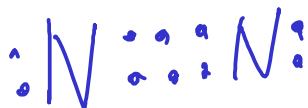
A few notes on the double bond:

- 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.

SO FAR, we've seen that ...

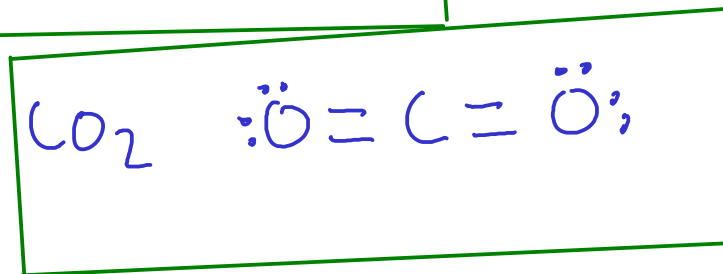
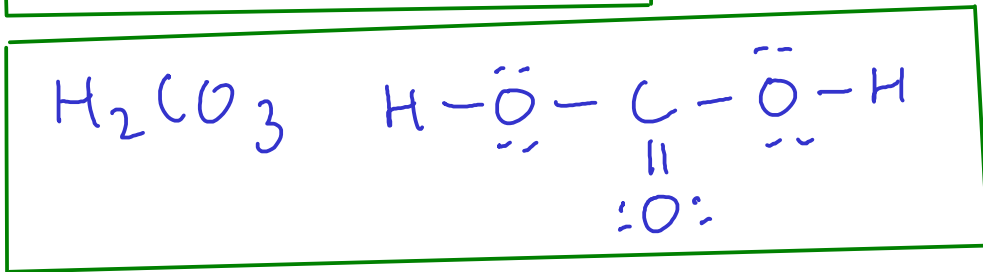
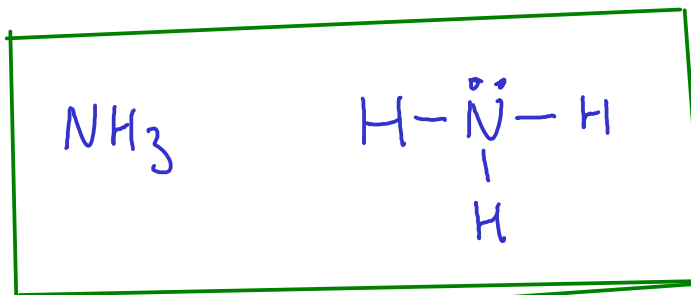
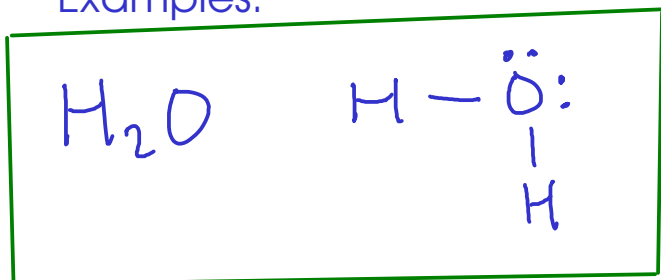
- ① Atoms may share one, two, or three pairs of electrons.
- ② Atoms will usually share enough electrons so that each atom ends up with a share in EIGHT electrons - the "octet rule"

- 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.

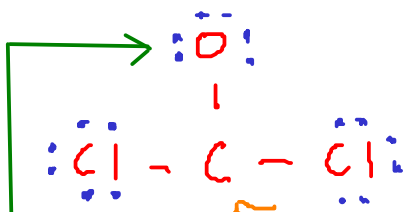


$$\begin{array}{l} \text{C: } 4 \\ \text{O: } 6 \\ \text{Cl: } 2 \times 7 = 14 \\ \hline \end{array}$$

24 electrons



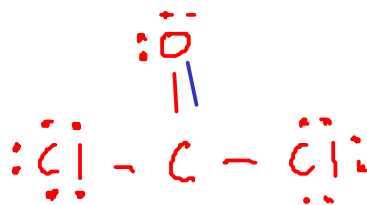
Choose CARBON as central atom.



Distribute the remaining electrons. We stop here because we ran out of electrons.

Carbon has a share in only SIX electrons!

We pick OXYGEN to share two pairs of electrons with carbon, since it needed two additional electrons from the beginning!



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