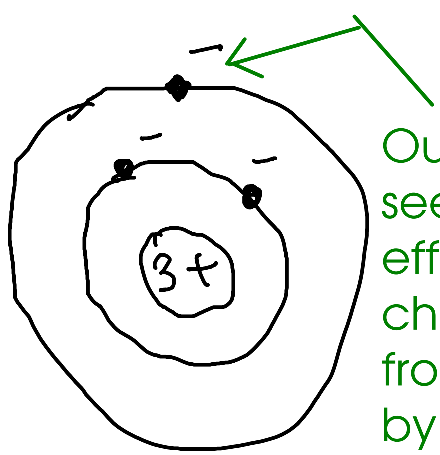
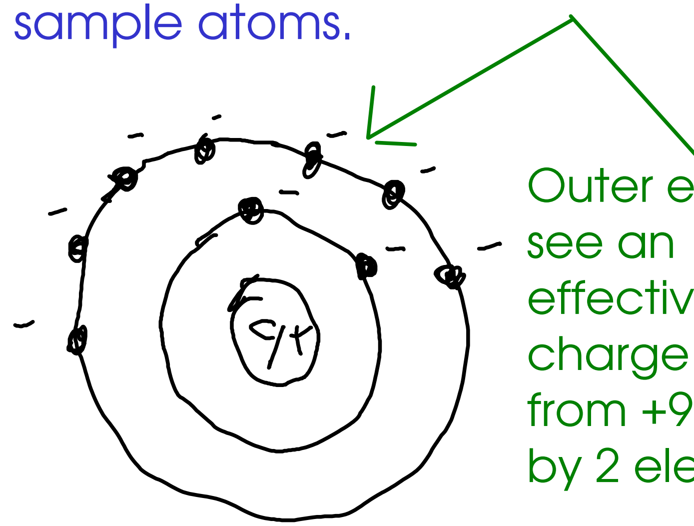
(FIRST) IONIZATION ENERGY

- 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 (↓), 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 (⟶), the ionization energy INCREASES.
    - Why? Let's look at some sample atoms.



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

lithium



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

fluorine

... since fluorine's outer electrons are held on by a larger effective charge, they are more difficult to remove than lithium's.

THE FIRST TWO PERIODIC TRENDS IN A NUTSHELL

LARGER IONIZATION ENERGY

SMALLER RADIUS

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

ELECTRON AFFINITY

- the electron affinity is the ENERGY CHANGE on adding a single electron to an atom.
  - Atoms with a positive electron affinity cannot form anions.
  - The more negative the electron affinity, the more stable the anion formed!

- General trend: As you move to the right on the periodic table, the electron affinity becomes more negative.

EXCEPTIONS

- Group IIA does not form anions (positive electron affinity)!

$\underbrace{n}_{\text{period number}}s^2$ valence electrons for Group IIA!

- To add an electron, the atom must put it into a higher-energy (p) subshell.

- Group VA: can form anions, but has a more POSITIVE electron affinity than IVA

ns^2np^3 valence electrons for Group VA!

$\underbrace{np^3}$ Half-full "p" subshell! To add an electron, must start pairing!

- Group VIIIA (noble gases) does not form anions

$\underbrace{ns^2np^6}$ full "s" and "p" subshells!

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
Covalent 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

ELECTRONEGATIVITY TRENDS

- 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

(p 346)

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 GAIN or LOSE enough electrons to end up with full "s" and "p" subshells.

example:

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

$1s^2 2s^2 2p^6$

[Ne] $3s^2 3p^1$

[Ar] $3d^{10} 4s^2 4p^5$

[Ar] $3d^{10} 4s^2 4p^5$

[Ar] $3d^{10} 4s^2 4p^5$

Al^{3+}: [Ne]

Br^-: [Ar] $3d^{10} 4s^2 4p^6$

Br^-: [Ar] $3d^{10} 4s^2 4p^6$

Br^-: [Ar] $3d^{10} 4s^2 4p^6$

Aluminum loses its outer three electrons, and each bromine gains one!

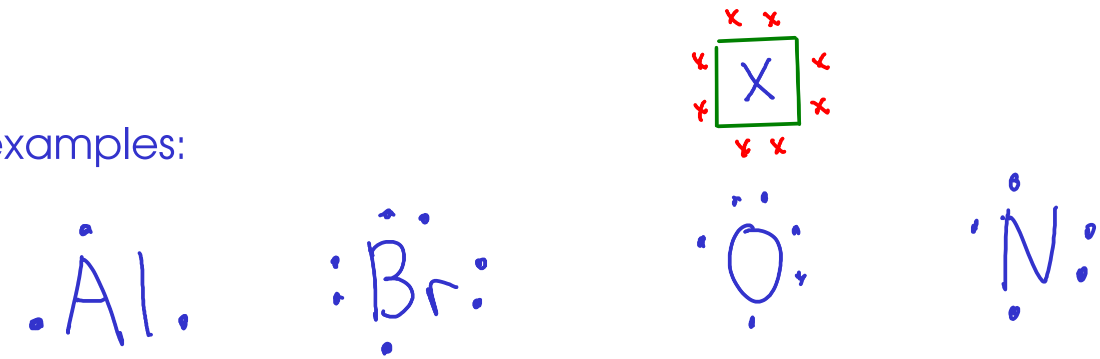
... but using electron configurations to describe how aluminum bromide forms is a bit cumbersome! Can we simplify the picture a bit?

LEWIS NOTATION / ELECTRON-DOT NOTATION

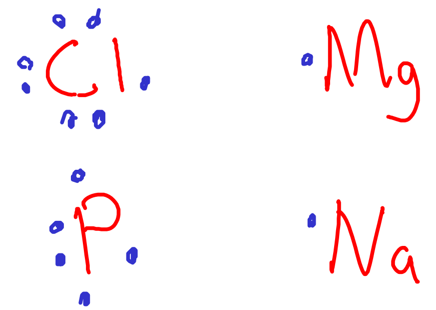
- Lewis notation represents each VALENCE electron with a DOT drawn around the atomic symbol. Since the valence shell of an atom contains only "s" and "p" electrons, the maximum number of dots drawn will be EIGHT.

- 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.

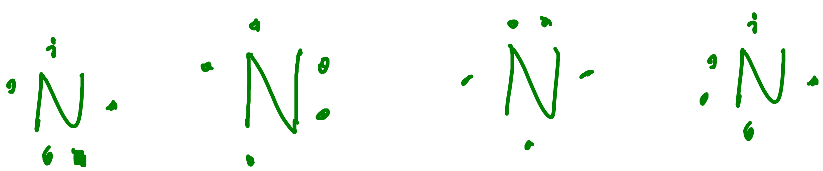
examples:



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!

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
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

2 valence electrons

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!

$$\cdot\dot{Al}\cdot \; + \; \cdot\ddot{\underset{\cdot\cdot}{Br}}: \;,\; \cdot\ddot{\underset{\cdot\cdot}{Br}}: \;,\; \cdot\ddot{\underset{\cdot\cdot}{Br}}: \;\longrightarrow\; Al^{3+} + \left[:\ddot{\underset{\cdot\cdot}{Br}}:\right]^- \;,\; \left[:\ddot{\underset{\cdot\cdot}{Br}}:\right]^- \;,\; \left[:\ddot{\underset{\cdot\cdot}{Br}}:\right]^-$$

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!

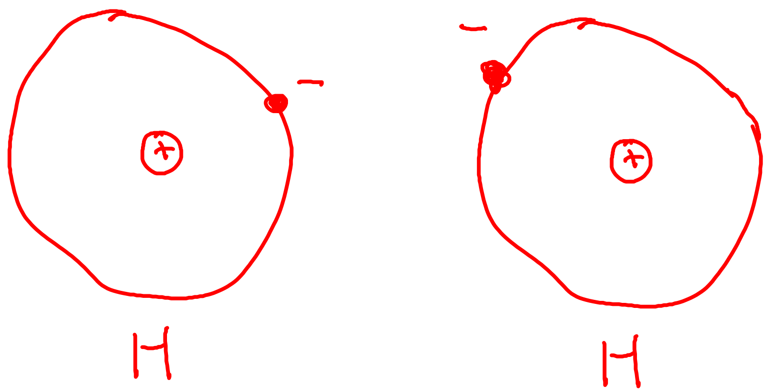
... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

MOLECULAR COMPOUNDS

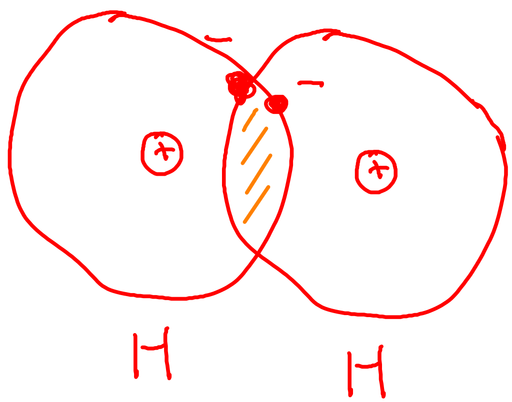
- Form when atoms SHARE electrons instead of transferring them. This results in the formation of MOLECULES ... groups of atoms held together by electron-sharing.

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?

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

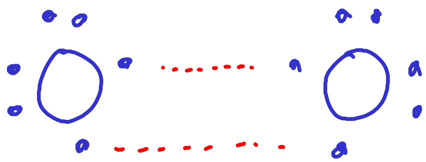
a single shared pair of electrons.
This is called a SINGLE BOND

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.

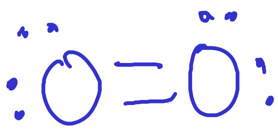
Let's look at OXYGEN ...



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



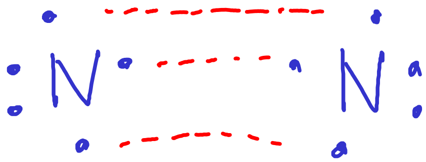
Each oxygen atom has a share in eight electrons!

A few notes on the double bond:

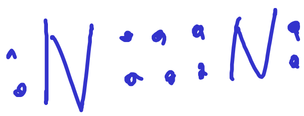
1. - For atoms to share more than one pair of electrons, they have to move closer to one another than they would if they were only sharing one pair of electrons. This BOND DISTANCE is measurable!

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 with each other.

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

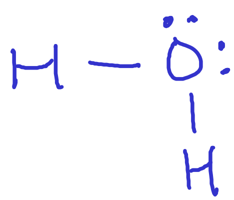
    - HYDROGEN will only end up with two electrons!

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

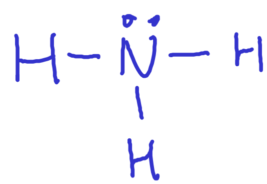
NOW, how could we come up with dot structures for some more complicated (and therefore, more interesting) molecules?

Examples:

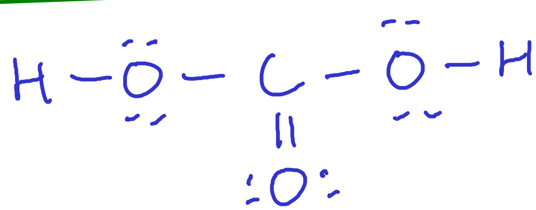
H_2O



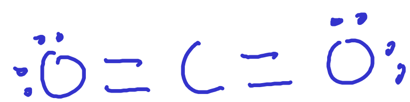
NH_3



H_2CO_3



CO_2



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.

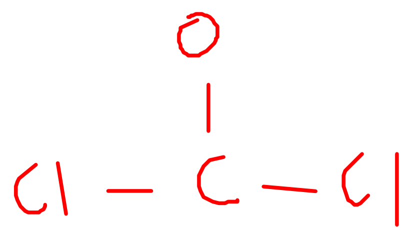
$COCl_2$

$C - 1 \times 4$

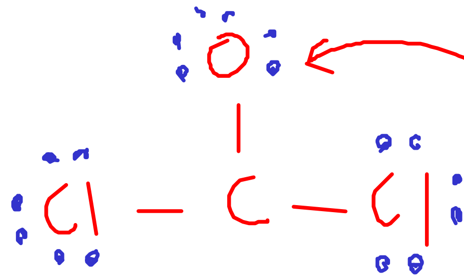
$O - 1 \times 6$

$Cl - 2 \times 7 = 14$

$24\ e^-$

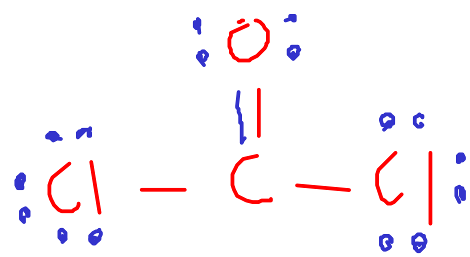


CHoose CARBON as central atom since it needs the most outer shell electrons to get 8 (4 more!)



Distribute electrons. Stop when you reach the total available (24 here)

... but CARBON has a share in only six valence electrons!



Create a double bond by using one of the "lone pairs" on OXYGEN.

Why oxygen? It needed to gain two more electrons, so was more likely to form two bonds. (Each bond "gains" an atom one more electron!)