Structure of Molecules

Bonding and Shape

Electron dot symbols

- Electron dot symbols are a way of showing the valence electrons for bonding
- Valence electrons are the s and p electrons in the outermost energy level.
- To draw the electron dot symbol, we imagine a square around the chemical symbol and place the valence electrons around the square. One electron per side initially and them start to pair them.

Hydrogen would look like: H·
Helium would look like: He:

Helium breaks the rule because it has a filled outer shell.

Carbon would look like: ·C· (with one dot on each of its four sides)

Draw the electron dot symbols for:

- Oxygen
- Nitrogen
- Fluorine

Electron Dot Structures

Also called Lewis Structures

- How atoms connect together to make compounds.
- Two types:
    - Ionic compounds
    - Covalent compounds
- Structure can give us other properties of the compound

Electron configuration of ions

- Metals lose electrons
    - Charge will equal group number for main group metals (tall columns)
- Non-metals gain electrons
    - Charge will equal group number minus eight
- Charge is the number of electrons gained or lost.

Lithium:
$1s^2\ 2s^1$
Lithium Ion (Li^+) :
$1s^2$

Lithium loses its one valence electron (2s) and has no more valence electrons.

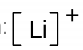
Fluorine:
$1s^2\ 2s^2\ 2p^5$
Fluoride ion (F^-):
$1s^2\ 2s^2\ 2p^6$

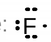
Fluorine gains one electron and has eight valence electrons.

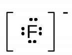
Notice that the electron configuration of the ions formed is the same as that of a noble gas.

We just saw that the atoms will gain or lose electrons so that it has an electron configuration that looks like a noble gas.
How does this affect the electron dot symbol?

Lithium: 

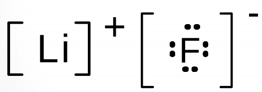
Lithium Ion: 

Fluorine: 

Fluoride ion 

Electron dot structure of ionic compounds

- Write the symbols for the ions.
- Place them next to each other



That's all there is to it for ionic compounds.

Electron dot structures for covalent compounds

- Determine the skeletal structure of the compound.
- Count the number of valence electrons in the compound.
- Connect all the outer atoms to the central atom by single bonds (use a single line).
- Give all of the outer atoms, except for H, an octet by placing the appropriate number of electron pairs around each atom. Any left over electron are placed in pairs on the central atom. Hydrogen can only have two electrons around it.
- If the central atom does not have an octet, form double or triple bonds until the central atom has an octet.

Example

CCl_4

Determine skeletal structure

- Atom that is able to form the most number of bonds is the central atom. In this case, carbon.

Cl

Cl C Cl

Cl

Count valence electrons

- Carbon has 4
- Each chlorine has 7 (7x4=28)
- Total is 32 (28+4=32)

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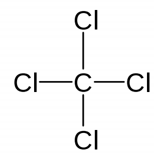
Connect outer atoms with single bonds

- Each single bond is two electrons. There are a total of 8 electrons used here.



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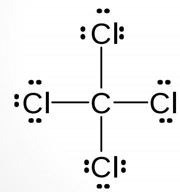
Fill octets on outer atoms

- Each of the chlorines needs six more electrons for an octet.



This structure uses all of the available electrons.

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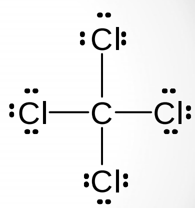
Check structure for octets

- All the atoms have an octet so this is the completed structure.
- No multiple bonds are needed.



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More Lewis Structures

Polyatomic ions and shapes of molecules

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Polyatomic ions

- A group of atoms, usually non-metals, that are covalently bonded together but have an overall charge.
- The Lewis structure takes this charge into account
  - Cations (positive charge) have fewer electrons
  - Anions (negative charge) have more electrons

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Structure of the carbonate ion (CO_3^{2-})

- Determine central atom:

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  Carbon because it can from up to 4 bonds
- Count valence electrons:

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  4 + 3x6 + 2 = 24
- Draw strucuture:

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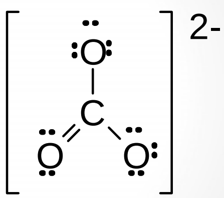

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Resonance Structures

- Occur when there is more than one correct way to draw the structure.
- Involves only a shifting of electrons, not atoms!
- All of the structures are equally correct, and they all exist at the same time.
- Carbonate ion is an example.

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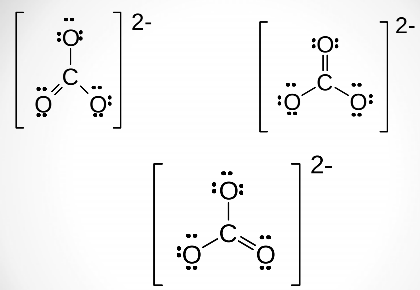

All of these are equally correct!

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Shapes of molecules

- Shape depends on the arrangement of electrons around the central atoms.
- The electrons will arrange themselves so that they are as far from each other as possible.
- This arrangement of electrons is what we perceive as the shape of the molecule.

VSEPR

- Count the number of groups of electrons around the central atom
- Determine the number of lone pairs out of that number of groups
- Follow the table (which must be memorized!).

VSEPR

Number of groups	Number of lone pairs	EPG	MG
2	0	Linear	Linear
	1		Linear
3	0	Trigonal Planar	Trig. Planar
	1		Bent
4	0	Tetrahedral	Tetrahedral
	1		Trig. Pyramidal
	2		bent

Draw Lewis structures and give electron pair and molecular geometries

- NI_3
- SO_3
- CO
- N_2O

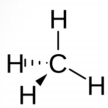
Electronegativity

- Electronegativity is the tendency of an element to pull electrons when it's in a bond with another element.
- Fluorine is the most electronegative, Cesium is the least.
- Electronegativity decreases as you move away from Fluorine in the periodic table
- The three most electronegative elements are N, O and F (in increasing order).

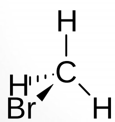
Molecular Polarity

- Some molecules have an uneven distribution of electrons due to a difference in electronegativity between the elements present in the compound.
- If the molecule is asymmetric, it is polar. For this class, if it contains a lone pair it is polar.
- If the molecule is symmetric, it's non-polar.

What do we mean by symmetry?



Methane (CH_4) is highly symmetric. No matter how we look at it, it looks the same.



Bromomethane (CH_3Br) is asymmetric because it looks different from different directions.

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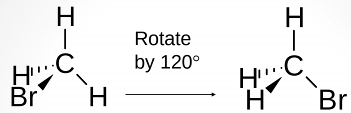

Because this molecule is asymmetric, it is polar. There is an uneven distribution of electrons in the molecule (they spend most of the time around the bromine).

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Draw Lewis structure, give geometries and polarity.

- FCN
- H_2S
- O_3
- CO_2

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