

Review

Valance Electron: Group Number.

Compound: A substance composed of two or more elements in fixed, definite proportions.

Forming Ions

Atoms of metals have few valance electrons (1-2) thus they tend to lose electrons to form a positive ion (cations).

Atoms of non-metals have many valance electrons (4-7) thus they tend to gain electrons to form negative ions (anions).

They do this to become stable in their outer shell.

Ionic Bonding: Type I

Format:	Name of Cation (metal)	Base Name of Anion (non-metal) + <i>ide</i>
Example:	NaCl	Sodium Chloride
	MgBr ²	Magnesium Bromide

Roman Numerals

1 = I	3 = III	5 = V	7 = VII
2 = II	4 = IV	6 = VI	8 = VIII

Ionic Bonding: Type II

Format:	Name of Cation (metal)	(Charge of cation (metal) in roman numerals)	Base name of Anion (non-metal) + <i>ide</i>
Example:	Copper	(I)	Chloride
	CuCl		

Ionic Bonding: Type II (cont)

CuCl² Copper (II) Chloride

VSEPR Theory

VSEPR: A theory based on the idea that electron groups (lone pairs, single bonds, or multiple bonds) repel each other.

VSEPR Ther

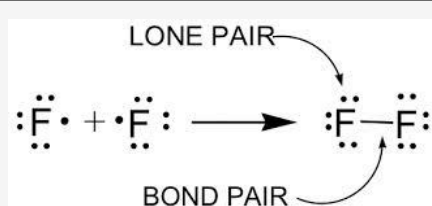
Drawing the Lewis Structure/Bonding

Step One: Draw the lewis structure for each covalent compound.

Step Two: Identify the bonds as single, double, or triple.

Step Three: Label the bonding and non-bonding electrons.

Example



Bonding - Why?

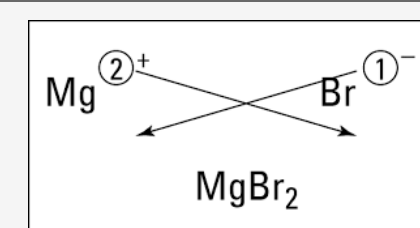
Octet Rule: Atoms bond in such a way as to obtain a full outer shell (8).

Bonding involved valance electrons only.

In general, atoms either transfer or share electrons to obtain a full outer shell (8).

Valance electrons are responsible for the chemical properties of an atom.

Ionic Bonding: Dot and Cross



Naming Compounds: Nomenclature

Is it Ionic? (Metal + One or more non-metals)

If so go to Type I and Type II.

OR

Is it Covalent? (All non-metals)

If so go to Type III.

Electron Groups

To determine the shape of a molecule, count only electron groups around the central atom.

Each of the following is consider *one electron group*.

Non-Bonding Pair - (A lone *pair* of electrons)

Bonding Electrons - (single, double, or triple)

Example: CH⁴ has 4 electron groups (4 single bonds, 0 lone pairs)

Drawing Molecular Geometries

Straight Line: Bond in plane of paper.

Hashed Line: Bonding going into paper.

Wedge: Bond coming out of the paper.



By chantalcramm

cheatography.com/chantalcramm/

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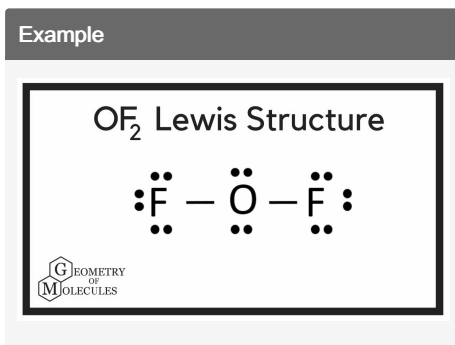
Terms	
Single Bond:	One pair of electrons shared between two atoms (Cl^2)
Double Bond:	Two pairs of electrons shared between two atoms. (O^2)
Triple Bond:	Three pairs of electrons shared between two atoms. (N^2)
Bonding Electrons:	Electrons shared between atoms.
Non-Bonding Electrons:	Electrons only found on one atom. (Lone pairs)
Overall:	Draw the lewis structure and determine how they will bond with one another to have full outer shells (8).
Identify the bonding and non-bonding electrons.	

Summary	
Ionic Bonding:	Covalent Bonding:
Metal + One or more non-metals.	All non-metals.
Electrons are transferred.	Electrons are shared.
Ions are formed.	Ions are not formed.
Ex. NaHCO_3 or NaCl	Ex. F^2 or CO^2

Prefixes				
1 =	3 =	5 =	7 =	9 =
Mono	Tri	Penta	Hepta	Nona
2 = Di	4 =	6 =	8 =	10 =
	Tetra	Hexa	Octa	Deca

Covalent Bonding: Type III				
Format:	Prefix	Base name element	Prefix	Base name of element
		1		2 + ide
Example:	Di	nitrogen	Mono	xide
		N^2O		
IF^3	---	Iodine	Tri	Fluoride
B^2H^8	Di	boron	Octa	hydride
CS^2	---	Carbon	Di	sulfide

Drawing Lewis Structures (2 Atoms or more)	
Step One:	Draw the lewis structure for each atom separately.
Step Two:	The atom that has the most unpaired electrons is the central atom.
Step Three:	The other atoms will share electrons with the central atom.



Possible Geometries

Molecular Geometries

Linear	Bent (V-shaped)	Trigonal planar	Trigonal pyramidal
Bond angles: 180°	<180°	120°	<120°
CO_2	SO_2	BCl_3	NH_3

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