

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 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

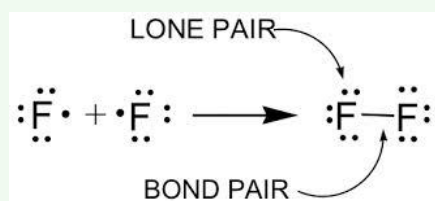
Drawing the Lewis Structure/Bonding

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?

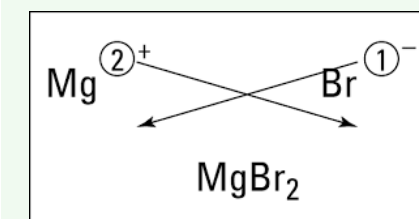
Octet Rule: Atoms bond in such a way as to obtain a full outer shell (8).

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.



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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	Prefix	Base name of element
		1		2 + <i>ide</i>

Example: Di nitrogen Mono *xide*
 N_2O

IF_3 --- Iodine Tri Fluoride

B_2H_8 Di boron Octa hydride

CS_2 --- Carbon Di sulfide

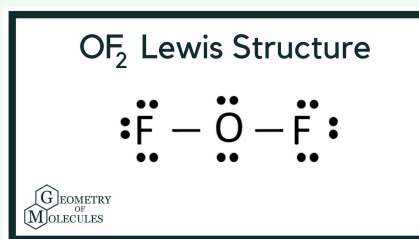
Drawing Lewis Structures (2 Atoms or more)

Step One: Draw the Lewis structure for each atom separately.

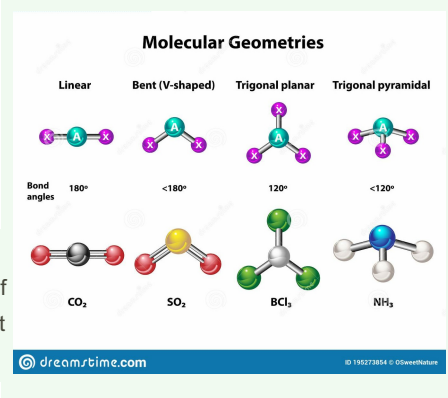
Step Two: The atom that has the most unpaired electrons is the central atom.

Step Three: The other atoms will share electrons with the central atom.

Example



Possible Geometries



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