Cheatography

Chemistry 1030: Unit 5 - Bonding and Nomenclature Cheat Sheet

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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) + *ide*
Example:	NaCl	Sodium Chlor*ide*
	$MgBr^2$	Magnesium Brom*ide*

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) + *ide*
Example: CuCl	Copper	(I)	Chlor*ide*

Ionic Bonding: Type II (cont)

$CuCl^2$	Copper	(II)	Chlor*ide*

VSEPR Theory

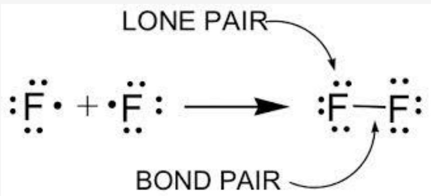
VSEPR:	A theory based on the idea that electron groups (lone pairs, single bonds, or multiple bonds) repel each other.

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?

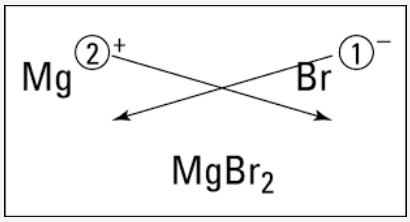
Octet Rule: Atoms bond in such a way as to obtain a full outer shell (8).

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: CH^4 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.

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

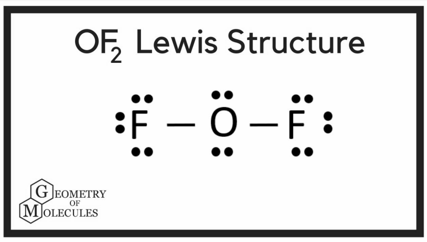
Covalent Bonding: Type III

Format:	Prefix	Base name element 1	Prefix	Base name of element 2 + *ide*
Example: N^2O	Di	nitrogen	Mono	x*ide*
IF^3	---	Iodine	Tri	Fluor*ide*
B^2H^8	Di	boron	Octa	hydr*ide*
CS^2	---	Carbon	Di	sulf*ide*

Drawing Lewis Structures (2 Atoms or more)

Step One:	Draw the lewis structure for each atom separetly.
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