Remember our discussion about extra stable e^- configs?

full shell > full subshell > ½ full subshell > others

Well those guidelines are going to enable us to predict if & what ions will form.

For example: Chlorine

adding one e^- will give us $3p^6$ → super good!

3p ↑↓ ↑↓ ↑↓
3s ↑↓
2p ↑↓ ↑↓ ↑↓
2s ↑↓
1s ↑↓

p^6 = Noble gas configuration - VERY stable!

$Cl: 1s^2 2s^2 2p^6 3s^2 3p^5$

$Cl^-: 1s^2 2s^2 2p^6 3s^2 3p^6$

$Na: 1s^2 2s^2 2p^6 \cancel{3s^1}$

$Na^+: 1s^2 2s^2 2p^6$

$3s^2$ ~~↑~~
$2p^6$ ↑↓ ↑↓ ↑↓
$2s^2$ ↑↓
$1s^2$ ↑↓

losing one e^- gives us a noble gas config.

So - if you can add or subtract e^- to create a noble gas config, this is a stable ion

Group IA → [Noble gas] _ s^1 → (+1) cation
Group IIA → [Noble gas] _ s^2 → (+2) cation
Group IIIA → [Noble gas] _ s^2p^1 → +3 cation (full s subshell isn't especially great)

Group VA → s^2p^3 → (-3) anion
Group VIA s^2p^4 → (-2) anion
Group VIIA s^2p^5 → (-1) anion

Predicting Charge on Transition Metals

We always fill the n^{th} s-subshell before the $(n-1)$ d subshell. However, once the e^- are there, the energy of those orbitals adjusts such that the s-electrons come out first (higher energy)

Nickel:
$3d \; \upharpoonleft\downharpoonright \; \upharpoonleft\downharpoonright \; \upharpoonleft\downharpoonright \; \upharpoonleft \; \upharpoonleft$
$4s \; \upharpoonleft\downharpoonright$
$\longrightarrow$
$4s \; \_\_$
$3d \; \upharpoonleft\downharpoonright \; \upharpoonleft\downharpoonright \; \upharpoonleft\downharpoonright \; \upharpoonleft \; \upharpoonleft$
Ni^{+2}

18 electron rule $\rightarrow$ $ns^2 \, nd^{10} \, np^6 = 18 \, e^-$

$Ni \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^2} 3d^{10}$ $\rightarrow$ if we remove these e^-, we have a stable e^- config

$Ni^{2+} \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} = 18 \, e^-$

$Cu \rightarrow 1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^1} 3d^{10} \rightarrow Cu^{+1} = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

This works well to predict lots of charges (often +2) on Transition metals.

- Also works for some p block elements: Ga^{3+}, In^{3+}, Tl^{3+}, Sn^{4+}, Pb^{4+}

Make sure to keep an eye out for other stable configs as well!

Fe: $[Ar] \, 4s^2 3d^6$ $\quad$ Fe^{+2} $[Ar] \, 3d^6 \rightarrow \upharpoonleft\downharpoonright \; \upharpoonleft \; \upharpoonleft \; \upharpoonleft \; \upharpoonleft$

Fe^{3+} $[Ar] \, 3d^5$ $\quad \underbrace{\upharpoonleft \; \upharpoonleft \; \upharpoonleft \; \upharpoonleft \; \upharpoonleft}_{\text{1/2 full subshell}}$

Nomenclature is a fancy word for names. So today we will focus on learning the rules associated with naming compounds.

Naming Ionic compounds → remember, they form between a metal and non-metal

Polyatomic Ions → KNOW THEM + their charges!

OH^-

CN^-

SCN^-

HCO_3^- + CO_3^{-2}

HSO_4^- + SO_4^{-2}

NO_2^- + NO_3^-

MnO_4^-

ClO^-, ClO_2^-, ClO_3^-, ClO_4^-

PO_4^{3-}

NH_4^+

Strategy: cation is said first

anion → drop last syllable and add "-ide"
or two

Cl^- → Chlor~~ine~~ → Chloride

N^{3-} → Nit~~rogen~~ → Nitride

S^{2-} → Sulf~~ur~~ → Sulfide

$NaCl$ → Sodium chloride

MgS → Magnesium sulfide

AlN → Aluminum nitride

$MgCl_2$ → Magnesium chloride

*do NOT indicate number of Cl^- because there is ONLY one way that Mg + Cl^- can interact to form a compound

-that is, the number of atoms is implied because of charges

Mg_2N_3 → Magnesium Nitride

Beryllium Oxide → Be^{2+} O^{2-} = BeO

Beryllium Nitride → Be^{2+} N^{3-} = Be_3N_2

Aluminum carbonate = Al^{3+} CO_3^{2-} = $Al_2(CO_3)_2$

Know your polyatomic ions!

- the parentheses indicate that there a two of everything inside

- So total 2 Al
2 C
6 O

Sodium Phosphate → Na^+ PO_4^{3-}

Na_3PO_4 no ( ) needed if only 1

Things change if we run into variable charge metals.

Iron Chloride → How do we know if this is Fe^{+2} or Fe^{+3}

$FeCl_2$ or $FeCl_3$

We Absolutely must indicate charge for variable charge cations

Iron(II) Chloride = $FeCl_2$ Iron(III) Chloride = $FeCl_3$

Tin(IV) oxide = Sn^{4+} O^{2-} = SnO_2

Tin(II) oxide = Sn^{2+} O^{2-} = SnO

$CuCl$ vs. $CuCl_2$

copper(I) chloride vs. Copper(II) chloride

Predicting Ionic compound Formulas:

Golden Rule → (+) MUST EQUAL (−)

1. Determine charges on each ion
2. Adjust number so that (+) + (−) = ∅ ← electrostatic neutrality

Sodium and Chlorine

↖ Group 1A = +1 ↖ Group 7A = −1

$+1 + -1 = \emptyset$ ✓

NaCl

Magnesium and chlorine

↖ +2 ↖ −1

$+2 + -1 \neq \emptyset$

$+2 + 2(-1) = \emptyset$ ✓

$MgCl_2$

Potassium and Oxygen

↖ +1 ↖ −2

$\underset{K^+}{2(+1)} + \underset{O^{2-}}{-2} = \emptyset$

K_2O

Aluminum and Oxygen

↖ +3 ↖ +2

$2(+3) + 3(-2) = \emptyset$

Al_2O_3

Predicting Ionic compound Formulas :

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Naming Ionic compounds → remember, these form between a metal and non-metal

Polyatomic Ions → KNOW THEM + their charges!

OH^-
CN^-
SCN^-
HCO_3^- + CO_3^{-2}
HSO_4^- + SO_4^{-2}
NO_2^- + NO_3^-
MnO_4^-
ClO^-, ClO_2^-, ClO_3^-, ClO_4^-
PO_4^{3-}

NH_4^+

Strategy: cation is said first

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Cl^- → Chlor~~ine~~ → Chloride
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Molecular compounds ⟶ formed through <u>covalent</u> bonds

(also called covalent compounds)

↳ shared electrons

form between two non-metals

-these differ from ionic

↳ electrons $\boxed{\text{NOT}}$ shared

form between a metal and a non-metal

Naming Covalent Compounds is easier. We still use the approach that the 1st element in the formula gets said 1st and the 2nd gets -ide

- the big difference is that we don't have cations or anions here so we can't assign charge. w/o charge, how can we know numbers of atoms?

number	Greek prefix
1	mono
2	di
3	tri
4	tetra
5	penta
6	hexa
7	hepta
8	octa
9	nona
10	deca

CO ⟶ Carbon monoxide

CO_2 ⟶ Carbon dioxide

N_2O_4 ⟶ dinitrogen tetraoxide

The order that atoms are named in a binary compound follows the same principle that ionic compounds do. Most (+) 1st.

For covalent compounds, this can be determined from Electronegativity. → this is how strongly one atom attracts electrons in a covalent bond. Very related to Electron Affinity

→ increasing Electronegativity

Name the compound formed between one sulfur atom and six chlorines.

- chlorine is more electronegative, so it's the "anion" and goes last

SCl_6 → sulfur hexachloride

Acids → recognize the formula because "H" comes first!

Naming boils down to 3 categories... based on name of anion that H^+ partners with (later, we'll learn these are called conjugate bases)

$HX \rightarrow H^+ + \boxed{X^-}$ ← this dictates how you name the acid!

1. "IDE rule": $HCl \rightarrow H^+ + \boxed{Cl^-}$ ← chloride

"hydro" + root + "ic acid"

hydrochloric acid   HBr : hydrobromic acid

this rule applies to all halogens!

2. "ATE rule" $HNO_3 \rightarrow H^+ + \boxed{NO_3^-}$ ← nitrate

root + "ic acid"   Nitric acid

$HClO_4$

perchlorate → perchloric acid

③ "ITE rule = $HNO_2 \rightarrow H^+ + \boxed{NO_2^-}$ ← nitrite

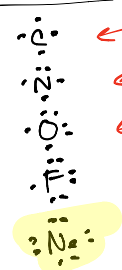
root + "ous acid = Nitrous acid

HClO hypochlorite ⟹ hypochlorous acid

Lewis Symbols

- convenient way to show valence electrons
  - dot ("·") represents one e^-
  - placed around the outside of an atom symbol

element	valence e^-	Lewis symbol
carbon	4	← needs 4 more e^- for octet
nitrogen	5	← 3 more
oxygen	6	← 2 more
fluorine	7	← 1 more
neon	8	← super stable with 8 V.E.



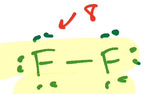
Covalent bonds form so that each atom has 8 V.E.
, we call this an octet

F_2 ← two F atoms covalently bonded to form a stable molecule

1. Draw Lewis symbol for each atom
2. Share e^- to form a pair
   - these shared e^- are "felt" by both nuclei
3. rewrite molecule showing ALL covalent bonds as a line ("—")
4. check to make sure all atoms have an octet AND correct # of e^- used

un paired = not stable

shared = covalent bond

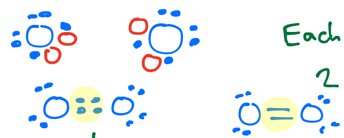


8 8

14 e^- used ✓

Molecular Oxygen O_2

O has 6 valence e^-
(v.e.)



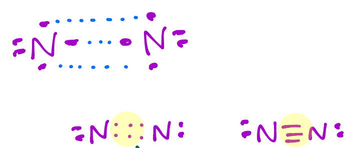
Each atom needs 2 electrons

double bond
2 pairs of shared e^-

* double bonds are shorter and stronger than single bonds *

Molecular Nitrogen N_2

$N \rightarrow 5$ v.e.



* Triple bonds are shorter and stronger than double bonds! *

Triple bond
3 pairs of shared e^-

	pm length	energy (aJ)
C−C	153	0.58
C=C	134	1.02
C≡C	120	1.35

N_2 is VERY unreactive because of the triple bond
– it's very hard to break

Lets move on to more complicated molecules:

$H_2O \rightarrow$ Note that H is written 1^{st} $\rightarrow$ water is an acid

$H_2O \rightarrow H^+ + OH^-$

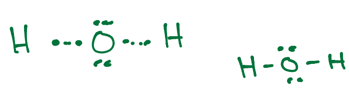
hydroxide

hydroxic acid

Hydrogen is in the 1^{st} shell, so a duplet (1 pair) is STABLE

* NEVER give H more than 1 bond! *

So O must be the central atom because it is the only one that can form multiple bonds



* Always show Lone pairs! *
(LP)

# of ve	
7	Cl
7	Cl
6	O
20	

Cl_2O

20 ← Lewis structure needs to have exactly 20 ve

2(10) ✓

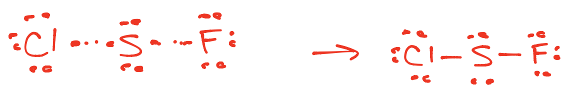
SClF

E.N $S = 2.5$

$Cl = 3.0$

$F = 4.0$

how are the atoms arranged? The general rule is that the least electronegative atom is the central atom.



General strategies for drawing Lewis structures:

① Use valence e- as a guide →

← needs to make 1 covalent bond

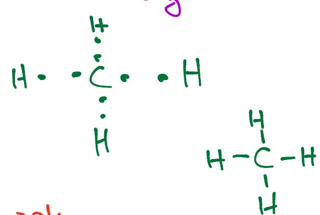
← needs to form 2 bonds

← needs to form 3 bonds

← 4 bonds

I find this method to be the most useful

CH_4



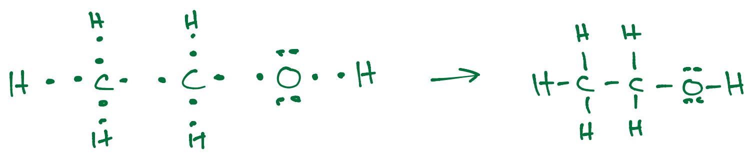
remember H can only form 1 bond, so C must be connected to C

CH_3CH_2OH

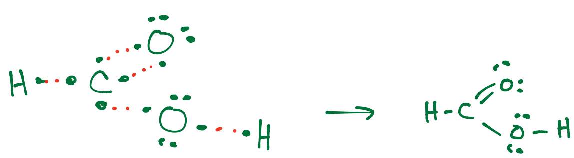
Things to ALWAYS check for:

① Do all atoms have an octet? (except H)

② Did you use the correct # of VE?

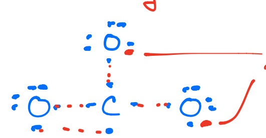


CHO_2H



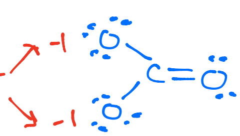
CO_3^{2-} ← we'll be needing 2 extra electrons

EN: 2.5 3.5



2 extra electrons on the molecule because of the -2 charge

extra e- went here!



octets? yes

VE? $C = 4$

$O = 3(6) = 18$

$-2 = 2$

24 → 24 used ✓