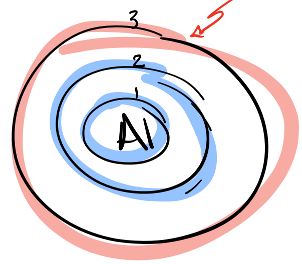
Writing Names & Formulas

Electronic Configurations

Al $\quad 1s^2 2s^2 2p^6 \mid 3s^2 3p^1$

13e⁻ $\quad$ Core e⁻ $\quad$ Highest level Valence e⁻

Valence e⁻

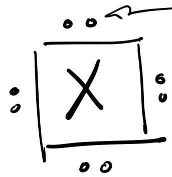


Valence are used for Chemical Reactions & for Chemical bonding

⟹ Determin all Chemical Reactivity

Lewis dot Structure

valence e⁻



1e⁻	2e⁻	3e⁻	4e⁻	5e⁻	6e⁻	7e⁻	8e⁻

Al $\quad 1s^2 2s^2 2p^6 \mid 3s^2 3p^1$

3 valence

Nobel Gasses have 8 valence e⁻
= filled energy level
& very stable
Nobility of PT

Main Group Elements
have between 1-8 valence e⁻

$s+p = 8$
$2+6 = 8$

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.30											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.84	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
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87 Fr Francium (223)	88 Ra Radium (226)		104 Rf Rutherfordium (261)	105 Db Dubnium (262)	106 Sg Seaborgium (266)	107 Bh Bohrium (264)	108 Hs Hassium (277)	109 Mt Meitnerium (268)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (272)	112 Cn Copernicium (285)	113 Nh Nihonium (284)	114 Fl Flerovium (289)	115 Mc Moscovium (288)	116 Lv Livermorium (289)	117 Ts Tennessine	118 Og Oganesson

2e⁻ 6e⁻ =8e

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Values rounded to four significant figures

1e⁻   2e⁻   # of valence e⁻   3e⁻ 4e⁻ 5e⁻ 6e⁻ 7e⁻   8e⁻

Group A values = # valence e⁻ for the group

	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
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Ion Formation

Ion is a atom or molecule that has unequal number of protons & electrons

Cations more protons = + Charge (p^+)

Anions More electrons = − Charge (e^-)

Catagory = Cation

Neutral Atom	name = Sodium		name = Sodium Ion		
Na	Na	$\longrightarrow$	Na^+	+	$1e^-$
$11p^+$	$11p^+$		$11p^+$		
$11e^-$	$11e^-$		$10e^-$		
net ∅ Charge	net = ∅		net = +1		

$1s^2 2s^2 2p^6 | 3s^1$

1A   1 valence e^-

Sodium Ion = $[1s^2 2s^2 2p^6 3s^0]^+$

$\underbrace{\text{Neon} = 1s^2 2s^2 2p^6}$

Isoelectronic
(Same electronic Config)

Chlorine	Chlorine		Chloride	
Cl	Cl + e^-	$\longrightarrow$	Cl^-	
$17p^+$	$17p^+$		$17p^+$	
$17e^-$	$17e^-$		$18e^-$	
∅	net = ∅		net = −1	

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

-ide Suffix = anion

$[1s^2 2s^2 2p^6 3s^2 3p^6]^-$

Na^+

$$Na\cdot \xrightarrow{-e^-} Na^{\oplus}$$

$$\underbrace{1s^2}_{2}\,\underbrace{2s^2 2p^6}_{8}\,\underbrace{3s^1}_{1} \qquad [\underbrace{1s^2}_{2}\,\underbrace{2s^2 2p^6}_{8}\,\underbrace{3s^0}_{0}]^+$$

2 - 8 - 1 → 2 - 8 - 0

Cl^-

$$:\ddot{\underset{\cdot\cdot}{Cl}}\cdot \xrightarrow{+e^-} :\ddot{\underset{\cdot\cdot}{Cl}}:^{\ominus}$$

$$\underbrace{1s^2}_{2}\,\underbrace{2s^2 2p^6}_{8}\,\underbrace{3s^2 3p^5}_{7} \qquad [\underbrace{1s^2}_{2}\,\underbrace{2s^2 2p^6}_{8}\,\underbrace{3s^2 3p^6}_{8}]^-$$

2 - 8 - 7 → 2 - 8 - 8

#e^- in each shell or principle energy level

$$\underset{2\text{-}8\text{-}1}{Na\cdot} + \underset{2\text{-}8\text{-}7}{:\ddot{\underset{\cdot\cdot}{Cl}}\cdot} \longrightarrow \underset{2\text{-}8}{Na^{\oplus}} + \underset{2\text{-}8\text{-}8}{:\ddot{\underset{\cdot\cdot}{Cl}}:^{\ominus}}$$

Trade e^-

$$Mg \longrightarrow Mg^{2+} + 2e^-$$

$1s^2 2s^2 2p^6 | 3s^2$ → $[1s^2 2s^2 2p^6]^{2+}$

2 - 8 - 2 → 2 - 8 - 0

Magnesium Ion

$$O + 2e^- \longrightarrow O^{2-}$$

$1s^2 | 2s^2 2p^4$ → $[\underbrace{1s^2}_{2}\,\underbrace{2s^2 2p^6}_{8}]^{2-}$

2 - 6 → 2 - 8

Oxyide

+1 +2 +3 share e^- -3 -2 -1

All Cations but multiple Charge States ?

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
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Nomenclature - A system of naming

Cation - Name of medal followed by ion

Ex

Na	Sodium	Na^{+}	Sodium ion
Mg	Magnesium	Mg^{2+}	Magnesium ion
Al	Alumimum	Al^{3+}	Alumimum ion

Anion - Elemental anion named by root-ide

Ex

		Group		
F	Fluorine	7A	F^{-}	Fluoride
O	Oxygen	6A	O^{2-}	Oxide
N	Nitrogen	5A	N^{3-}	Nitride
S	Sulfur	6A	S^{2-}	Sulfide
P	Phosphorus	5A	P^{3-}	Phosphide
Br	Bromine	7A	Br^{-}	Bromide

Activity 4 - Writing Formulas and Names

Goals

- Write the electron dot structure for an atom.
- Predict the charge of an ion from its electron dot structure.
- Use the periodic table to determine the ionic charge of a metal or nonmetal ion.
- Write the correct formula and name of an ionic or covalent compound.
- Write the correct formula and name of a compound containing polyatomic ions.

Pre-lab Questions *(answer these on a separate sheet using complete sentences)*

1. Where are the valence electrons in an atom located?
2. How is a positive ion formed from an atom? Why is the charge positive?
3. How is a negative ion formed from an atom? Why is the charge negative?
4. How are the group numbers on the periodic table related to the number of valence electrons? To ionic charge?
5. How do subscripts represent the charge balance in polyatomic ions?
6. According to what rubric are electrons shared in covalent compounds (i.e. what does electron sharing accomplish?
7. How do the names of covalent compounds differ from the names of ionic compounds?
8. What are polyatomic ions? How are they named?

Concepts to Review

Electronic structure (energy levels)
Formation of positive and negative ions
Balancing ionic charge
Ionic and covalent compounds
Writing formulas of ionic and covalent compounds
Naming ionic and covalent compound

Introduction

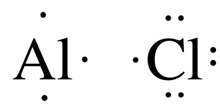
Most of the chemical reactivity of an element is determined by the **valence electrons**, which are the electrons in the highest energy level (or outermost electron shell). Usually in a compound, each atom has an **octet** of electrons (i.e. eight of these) in each of the valence shells. An octet of valence electrons provides atoms with the stable electron configuration found among the noble gases, a group of elements that are particularly stable and inert (unreactive). The first noble gas ($_2$He) does not have an octet since the second electron fills the first (n=1) valence shell, which can accommodate only two electrons.

Required Materials

A Periodic Table of the Elements.

A. Electron Dot Structures

When atoms of metals in groups 1A, 2A or 3A react with atoms of nonmetals in groups 5A, 6A and 7A, the metals lose valence shell electrons and the nonmetals gain valence shell electrons. We can predict the number of electrons lost or gained by analyzing the electron dot structures of the atoms. In an electron dot structure, the valence electrons are represented as dots around the symbol of the atom. For example, aluminum has 13 electrons, 2 in the first energy level, 8 in the second energy level and 3 in the third energy level. To describe this electronic structure we write the **electron arrangement** as 2-8-3. The last number represents the valence electrons so aluminum has three valence electrons and thus an electron dot structure with three dots. Chlorine (electron arrangement 2-8-7) has seven valence electrons and an electron dot structure with seven dots.



Main group metals (group A elements) with 1, 2 or 3 valence electrons *lose* their valence electrons to reach a stable electron configuration with a filled outer shell. For example, an aluminum atom loses its three valence electrons to reach stability and thus acquires an ionic charge of 3+. It is now an aluminum ion with a new electron arrangement of 2-8 (note the complete octet in the outer shell). Positive ions keep the same name as the element.

	Aluminum atom	Aluminum ion	
Symbol	Al	Al^{3+}	
Electron arrangement	2-8-3	2-8-0	(3 electons lost)
Number of protons	13 p^+	13 p^+	(same)
Number of electrons	13 e^-	10 e^-	(3 fewer electrons)
Net ionic charge	0	3+	

When nonmetals with 5, 6 or 7 valence electrons combine with metals, they *gain* electrons to complete their outer shells, and form stable (negatively charged) ions. For example, a chlorine atom gains one valence electron to become stable with an electron arrangement of 2-8-8. With the addition of one electron, chlorine becomes a chloride ion with an ionic charge of 1-. (When two elements combine to form a binary compound called a salt, the name of the negative ion ends in ***-ide***.)

	Chlorine atom	Chloride ion	
Symbol	Cl	Cl^-	
Electron arrangement	2-8-7	2-8-8	(electon added)
Number of protons	17 p^+	17 p^+	(same)
Number of electrons	17 e^-	18 e^-	(1 more electron)
Net ionic charge	0	1-	

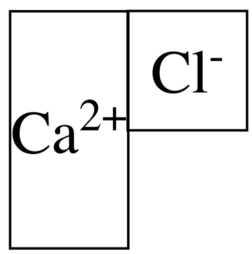
In the worksheet, *write the electron arrangements for atoms and their ions. Write the symbol, ionic charge, and name of each ion.*

B. Writing Ionic Formulas

The group number on the periodic table can be used to determine the ionic charges of elements in each family. **Nonmetals form ions only if they combine with a metal; if they combine with another non-metal, they form covalent (non-ionic) compounds.**

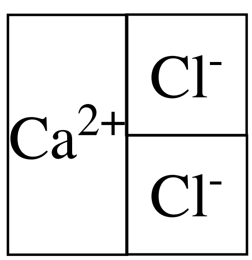
Group number	1A	2A	3A	4A	5A	6A	7A
Valence electrons	$1e^-$	$2e^-$	$3e^-$	$4e^-$	$5e^-$	$6e^-$	$7e^-$
Change	lose	lose	lose	none	gain	gain	gain
Ionic charge	1+	2+	3+	none	3-	2-	1-

In an ionic formula, the **total loss of electrons and total gain of electrons are equal.** The overall net charge is *zero*. This means that the total amount of positive charge must be made equal to the total amount of negative charge. To balance charge, we determine the smallest number of positive and negative ions that give an overall charge of *zero*. We can illustrate the process by showing the ions Ca^{2+} and Cl^- as geometric shapes that represent the amount of ionic charge:



The charge is balanced by using two Cl^- ions to match the shape of the Ca^{2+}. Charge balance occurs with one calcium ion and two chloride ions. This is shown by the subscripts in the formula $CaCl_2$. The subscript "1" for Ca is understood; it is never written. Note that *only the symbols are written in the formula, not the ionic charges*.

Balancing amount of ionic charge — Resulting formula



$$= \quad CaCl_2$$

In the following worksheet, *write the symbols of the positive and negative ions using the periodic table to determine the charge. Determine the number of each ion that will give a charge balance. Write the correct formula using subscripts to indicate that two or more ions were needed. Write the names of the ionic compounds by placing the metal name first, then the nonmetal name ending in* **-ide**.

C. Ionic Charges for Transition Metals

Most of the transition metals form more than one type of positively charged ion (or cation). We will illustrate this variable valence (combining capacity) with *iron*. Iron forms two cations, one (Fe^{2+}) with a 2+ charge, and another (Fe^{3+}) with a 3+ charge. To distinguish between the two ions, a Roman numeral that gives the ionic charge of that particular ion follows the element name (see below). The Roman numeral is always included in the names of compounds with positive ions that can have variable charge (or oxidation) states. (In an older naming system, the ending *-ous* indicates the lower valence and the ending *-ic* indicates the higher one. Compound names using this system still appear on old reagent bottles and in old chemistry texts.)

Ions	Names	Formula of Compound	Name of Compound
Fe^{2+}	iron (II) ion or ferrous ion	$FeCl_2$	Iron (II) chloride or ferrous chloride
Fe^{3+}	iron (III) ion or ferric ion	$FeCl_3$	Iron (III) chloride or ferric chloride

Among the transition metals a few elements (zinc, silver and cadmium) form only a single type of ion; these have fixed ionic charges and are *not* variable, hence these *do not* use a Roman numeral in their names. Examples are: **Zn^{2+}, zinc ion; Ag^{+}, silver ion; Cd^{2+}, cadmium ion.**

D. Polyatomic Ions

When an ionic compound consists of three or more kinds of atoms, there is a generally a central atom (usually a metal), and a group of attached nonmetal atoms. Such ions are called *polyatomic ions*. A polyatomic ion is a group of covalently bonded atoms with an overall charge that is usually negative. The most common polyatomic ions consist of the nonmetals C, N, S, P, Cl or Br, combined with two to four oxygen atoms. Some examples are given below. The ions are named by replacing the ending of the nonmetal with -**ate** or -**ite**. The form for each central element's most common oxidation state takes the -*ate* ending; the -*ite* ending is for the ion with one less oxygen atom than the -*ate* ion has.
Ammonium ion, NH_4^+, is unusual because it has a positive charge and a metal-like name.

Common polyatomic ions		One oxygen less than common ion	
HO^-	hydroxide ion		
NO_3^-	nitrate ion	NO_2^-	nitrite ion
CO_3^{2-}	carbonate ion		
HCO_3^-	hydrogen carbonate ion or bicarbonate ion		
SO_4^{2-}	sulfate ion	SO_3^{2-}	sulfite ion
HSO_4^-	hydrogen sulfate ion or bisulfate ion	HSO_3^-	hydrogen sulfite ion or bisulfite ion
PO_4^{3-}	phosphate ion	PO_3^{3-}	phosphite ion

Some Common polyatomic list

Note that there are polyatomic ions that consist of only two different types of atoms (e.g. hydroxide ion; see above) as well as those that have multiple copies of the same atom (e.g. azide ion: N_3^-).

In the worksheet, *write the formulas of compounds that contain ions of transition metals with variable valences. Write a correct name for each compound listed. Be sure to indicate the ionic charge if the transition metal has a variable valence by using a Roman numeral.*
To write the correct formula of a compound with a polyatomic ion, determine the ions required to achieve charge balance just as earlier. When two or more polyatomic ions are needed, enclose the formula of the ion in parentheses, and write the subscript **outside** the parentheses. *No change is made in the formula of the polyatomic ion itself.*

Consider the formula of the compound formed by Ca^{2+} and NO_3^- ions. The ions are $\mathbf{Ca^{2+}}$ and $\mathbf{NO_3^-}$. Since 2 nitrate ions will be needed to balance the charge on the calcium ion, we will need to indicate two NO_3^- ions using parentheses: $\mathbf{Ca(NO_3)_2}$. Note that the formula of the nitrate ion is not changed.

In the worksheet, *determine the positive ions and negative polyatomic ions needed for charge balance. Write the formula using parentheses if necessary. Name the compounds listed using the correct names of the polyatomic ions.*

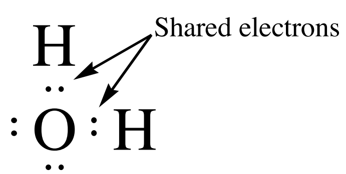
E. Covalent (Molecular) Compounds

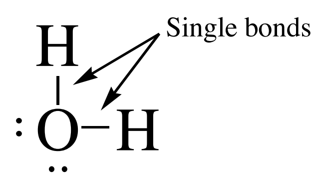
Covalent bonds form between nonmetal atoms located in Groups 4A, 5A, 6A or 7A. In a **covalent compound,** octets are achieved by sharing electrons between atoms. The sharing of one pair of electrons is referred to as a single bond. A double bond is the sharing of two pairs of electrons between atoms. In a triple bond, three pairs of electrons are shared. To write the formula of a covalent compound, determine the number of electrons needed to complete an octet. For example, nitrogen in Group 5 has five valence electrons so that it needs 3 more electrons for an octet; each nitrogen atom shares 3 electrons in covalent compounds.

Electron Dot Structures

The formulas of covalent compounds are determined by sharing valence electrons until each atom has an octet. For example, in water, oxygen shares two electrons with two hydrogen atoms. Oxygen has an octet and the hydrogen atoms are stable because they have two electrons in their outer (valence) shells. (Note that shared electron pairs are often represented as lines connecting the atoms that share them.)

Dot Structure for H_2O

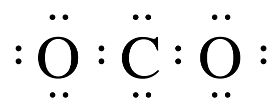




In another example, consider CO_2, a molecule that has two double bonds. In electron dot structures, carbon has 4 valence electrons and each oxygen atom has 6. Thus for CO_2, a total of 16 (4 + 6 + 6) electrons can be used in forming the octets by sharing electrons. We can use the following steps to determine the electron dot structure for CO_2:

1. Connect the atoms with pairs of electrons, which uses 4 electrons:

2. Add electrons to complete octets around all atoms:



3. Count the number of electrons used. We used 20. But only 16 are available. Therefore, we must economize by removing 2 pairs of electrons (4 electrons), and move 2 other pairs in between the C and the O's in order to maintain octets. This will form two double bonds:

Names of Covalent Compounds

Binary (two-element) covalent compounds are named by using *prefixes* that give the number of atoms of each element in the compound. The first nonmetal is named by the element name; the second ends in *-ide*. The prefixes are derived from Greek names: mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, octa-, nona-and deca-. (Higher ones exist, but are rarely used.) Usually the prefix mono- is not shown for the first element.

Formula of Covalent Compound	Name
CO	carbon **mon**oxide
CO_2	carbon **di**oxide
PCl_3	phosphorus **tri**chloride
N_2O_4	**di**nitrogen **tetr**oxide ("a" dropped before the vowel "o" in "oxide")
SCl_6	sulfur **hexa**chloride

In the worksheet, *write the electron dot structure for each nonmetal. Then write electron dot structures for the covalent compounds. Name each of the covalent compounds, using the numerical prefixes when appropriate.*

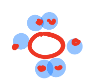
Activity 4 - Writing Formulas and Names Worksheet

Name______________________________

Section____________ **Date**_______

Exercise A. Electron Dot Structures

1. Using the example given, complete this table.

Element	Atomic Number	Electron arrange-ment of atom	Electron dot structure of atom	Loss/gain of electrons by atom	Electron arrange-ment of ion	Ionic charge	Symbol of ion	Name of ion
Sodium	11	2-8-1	Na•	Lose 1 e^-	2-8	1+	Na^+	Sodium ion
Oxygen	8	$1s^2 2s^2 2p^4$ 2-6		Gain $2e^-$	$[1s^2 2s^2 2p^6]^{2-}$ 2-8	2-	O^{2-}	Oxide
Aluminum								
Potassium								
Chlorine								
Calcium								
Nitrogen								
Sulfur								

2. Review the "Name of ion" column above. What distinguishes the naming of the metal cations from the naming of nonmetals anions?

Ionic Formulas of Ionic Compounds

M^{+} NM^{-} = Combine in ratios to make a net zero compound.

$\#$ of + = $\#$ of −

Sodium ion	Fluoride
Na^{+}	F^{-}
+1	−1

= NaF

no charges

no space

Sodium fluoride

Cation 1st Anion 2nd

In both formula & name it is Cation 1st & anion 2nd

Mg^{2+}	F^{-} F^{-}
2+	2−

$\left.\begin{matrix}F^{-}\\F^{-}\end{matrix}\right\}2$

MgF_2

no charges written

charges are implied

Magnesium fluoride

nothing in name that tells how many.

Al^{3+}	O^{2-}
Al^{3+}	O^{2-}
	O^{2-}
$\cancel{3+}$	$\cancel{2-}$
$6+$ ✓	$\cancel{4-}$
	$6-$ ✓

2   3

Al_2O_3 ← Always lowest possible

Aluminum Oxide

Mg^{2+}	O^{2-}
$2+$	$2-$

~~Mg_2O_2~~

MgO

magnesium Oxide

Exercise B. Writing Ionic Formulas:

1. Use the periodic table to help complete the table below.

Name	Positive ion	Negative ion	Formula
Sodium oxide	Na^+	O^{2-}	Na_2O
Mg Cl Magnesium chloride	Mg^{2+}	Cl^-	$MgCl_2$
Potassium chloride			
Calcium oxide			
Aluminum bromide			
Lithium phosphide			
Aluminum sulfide			
Aluminum nitride			
Calcium nitride			

2. Name the following ionic compounds:

Na_2S Sodium Sulfide

MgF_2 ____________________

MgS ____________________

K_3N ____________________

Ca_3P_2 ____________________

$AlCl_3$ ____________________

3. Review the answers in problems 1 and 2 of exercise B above. What do the subscripts represent?

Transition Metal Ions

Transition metal ions come in multiple charge states.

Why? ⇒ d-orbitals

Sometimes they lose s-orbital
they lose d-orbital
they lose both s & d

Most Common

$+1/+2$	$+2/+3$	$+2/+4$
Cu^{+}/Cu^{2+}	Fe^{2+}/Fe^{+3}	Pb^{+2}/Pb^{+4}
Hg_2^{+2}/Hg^{2+}	Co^{+2}/Co^{+3}	Sn^{+2}/Sn^{+4}
	Mn^{+2}/Mn^{+3}	
	Ni^{+2}/Ni^{+3}	
	Cr^{+2}/Cr^{+3}	

FeO
White

Fe^{2+}	O^{2-}
$2+$	$2-$

Iron (II) Oxide

Fe_2O_3
Red Colored
(Rust)

Fe^{3+}	O^{2-}
Fe^{3+}	O^{2-}
	O^{2-}
$6+$	$6-$

Iron (III) Oxide

Roman numeral tells Charge State on metal, and used on all transition metals with more than 1 Charge State.

Zn^{2+}
Ag^{+}

only have 1 charge state $\Rightarrow$ like main group metal
Roman numeral <u>not</u> used for these two

Exercise C. Ionic Charges for Transition Metals

1. Complete the table below.

Name	Positive ion	Negative ion	Formula
Iron (II) bromide	Fe^{2+}	Br^-	$FeBr_2$
Iron (II) chloride			
Iron (III) sulfide	Fe^{3+}	S^{2-}	Fe_2S_3
Copper (II) chloride			
Copper (II) sulfide			
Copper (II) nitride			
Zinc^{2+} oxide			
Silver^{+} sulfide			

Fe^{3+} Fe^{3+} | S^{2-} S^{2-} S^{2-}

$6+$ | $6-$

2. Name the following ionic compounds:

Cu_3P ______________________________

Fe_2O_3 ______________________________

FeI_3 ______________________________

$CuCl$ ______________________________

$ZnBr_2$ ______________________________

3. Consider your answers in problems 1 and 2 of exercise C above. What do the roman numerals in parentheses represent?

How to go from formula back to name?

Cu_3P =

$\frac{3+}{3Cu}$

Cu^{+}	P^{3-} (fixed charge state)
Cu^{+}	
Cu^{+}	
3+	3−

Copper (I) Phosphide

Always work from the anion back to Cation Charge State.

Cr_2S_3

Cr^{3+}	S^{2-}
Cr^{3+}	S^{2-}
	S^{2-}
6+	6−

$\frac{6^+}{2} = 3+$

Chromium (III) Sulfide

Exercise D. Polyatomic Ions

1. Complete the table below.

Name	Positive ion	Negative ion	Formula
Sodium nitrate	Na^{+}	NO_3^{-}	$NaNO_3$
Lithium carbonate	Li^{+}	CO_3^{2-}	Li_2CO_3
Potassium sulfate			
Calcium bicarbonate			
Aluminum hydroxide			
Lithium sulfite			
Sodium phosphate			
Iron (II) phosphate			

Li^{+} | CO_3^{2-}
Li^{+}
2+ | 2−

2. Name the following ionic compounds:

$CaSO_4$ ____________________________

Cu_3PO_4 ____________________________

$Al(NO_3)_3$ ____________________________

Na_2CO_3 ____________________________

$MgSO_3$ ____________________________

$Fe(HCO_3)_2$ ____________________________

Careful of transition metals needing roman numerals

3. Consider *all* of the nomenclature exercises in exercises B, C and D. What are the rules for the correct placement of parentheses in the naming and writing chemical formulas of ionic compounds?

Polyatomic Ion

$\underbrace{\text{Poly}}_{\text{many}}\underbrace{\text{atomic}}_{\text{atom}}$ Ion

$SO_4^{2-} \leftarrow 50e^-$

S $16p^+$ $\quad 16$

O $8p^+ \times 4 = 32$

$48p^+$

Have more or less e^- than the total $\#p^+$

Monatomic Ion

$\underbrace{\text{Mon}}_{\text{one}}\underbrace{\text{atomic}}_{\text{atom}}$ Ion

O^{2-}

N^{3-}

Mg^{2+}

Al^{3+}

Fe^{3+}

Have more or less e^- than the $\#p^+$

Polyatomic Family

Cl Chlorine Family w/ Oxygen

Chloride (Cl^-)

Prefix Root Suffix

Same Charge as Monatomic ion

ClO_4^- perChlorate per = more

ClO_3^- Chlorate "ate" more oxygen

Root Chlor

ClO_2^- Chlorite "ite" less oxygen

ClO^- hypo Chlorite hypo = less

Sulfur family S^{2-}

S^{2-} Sulfide

SO_4^{2-} Sulfate

SO_3^{2-} Sulfite

Phosphorus family P^{3-}

P^{3-} Phosphide

PO_4^{3-} Phosphate

PO_3^{3-} Phosphite

* Nitrogen family N^{3-}

N^{3-} Nitride

NO_3^{-} Nitrate

NO_2^{-} Nitrite

Polyatomic Compounds

Na^+	PO_4^{3-}
Na^+	
Na^+	

$\left.\begin{matrix} Na^+ \\ Na^+ \\ Na^+ \end{matrix}\right\} 3 \qquad \left.PO_4^{3-}\right\} 1$

Na_3PO_4

Sodium Phosphate

Fe^{3+}	SO_4^{2-}
Fe^{3+}	SO_4^{2-}
	SO_4^{2-}
6+	6-

$\left.\begin{matrix} Fe^{3+} \\ Fe^{3+} \end{matrix}\right\} 2 \qquad \left.\begin{matrix} SO_4^{2-} \\ SO_4^{2-} \\ SO_4^{2-} \end{matrix}\right\} 3$

$Fe_2(SO_4)_3$

Iron (III) Sulfate

Exercise E. Lewis Dot Structures of Atoms and Molecules

1. ***Electron dot formulas of elements:*** Atoms are represented by symbol with valence e^-'s represented by dots. Complete the following table. Distribute dots on all four sides before pairing.

Hydrogen	Carbon	Nitrogen	Oxygen	Sulfur	Chlorine
H·					

2. ***Electron dot formulas of covalent compounds:*** Lewis dot structures must have the correct number of valence electrons displayed in bonded or nonbonded pairs along with the octet rule being obeyed (duet rule for H). Complete the following table for the given binary covalent compounds.

Compound	Electron dot structure	Name
HCl	$\mathrm{H:\ddot{\underset{..}{Cl}}:}$	Hydrogen chloride
SBr_2		Sulfur dibromide
PCl_3	skip	
OF_2		
SO_3		

Questions and Problems

1. Write the correct formulas for the following ions:

sodium ion ___________	oxide ion ___________	calcium ion ___________
chloride ion __________	sulfate ion __________	iron (II) ion ___________

2. Write the correct name of the following compounds.

CuO	____________________
$\cancel{N_2O_4}$	____________________
$Al(NO_3)_3$	____________________
$\cancel{PCl_3}$	____________________
$FeCO_3$	____________________
$Cu(OH)_2$	____________________
Ag_2O	____________________

3. Identify the following compounds as ionic or covalent. (circle I or C) and write the corresponding molecular formula.

Sodium oxide	I C	__________________________
Iron (III) bromide	I C	__________________________
Sodium carbonate	I C	__________________________
Aluminum sulfite	I C	__________________________
Carbon tetrachloride	I C	__________________________
Nitrogen tribromide	I C	__________________________

Skip

4. Your friend wants to know what the formula $FeSO_4$ on her vitamin bottle means and what the name of this ingredient is. Help her understand the meaning of the symbols and the correct name associated with this formula (i.e write a brief answer to her question).