Naming Compounds BASICS

1. Periodic table layout and the properties of metals, non-metals, and metalloids
2. Mono-atomic ions
3. Polyatomic ions (table E)
4. Ionic bonding
5. Naming ionic compounds, writing ionic compound formulas, writing formula from names
6. Transitional metal with multiple cations: the Roman Numeral or STOCK naming
7. Molecular compounds
8. Naming molecular compounds using the prefix methods, DIATOMIC ELEMENTS
9. Oxidation numbers, using to make sensible molecular compound ratios
10. Naming of acids and bases (tables K and L).

Periodic Table Layout (revisited)

The periodic table was designed to list the elements in a way to group them by similarity, and to also order them with regular "periodic" repetitions of chemical properties. It has a oddly shaped outer border, but to keep the elements in groups that make sense, this is the shape it must have.

The Periodic Law states: when the atoms are arranged in order of increasing atomic number (in this weird shape) that there is a periodic repetition of chemical properties, that arise in the groups.

The vertical columns are the GROUPS. Elements in GROUPS have similar chemical properties.

Group 1 are the ALKALI metals. Group 1 elements have just one electron in its outermost orbital.
Outermost orbitals are called VALENCE ORBITALS.
Group 1 elements all have 1 VALENCE ELECTRON in their VALENCE ORBITALS.

Group 2 are the ALKALINE EARTH metals.
Group 2 elements have two valence electrons in their valence orbitals.
This gives every atom in group 2 similar bonding abilities and ion formation.

Group 18 are the NOBLE gases. These have eight valence electrons in its valence orbital, except for He.
Helium is too small and has just that baby sized orbital which fits up to 2 electrons.
All group 18 elements have only complete outer orbitals, the PERFECT ARRANGEMENT of ELECTRONS. Because of this they do not make any compounds, they do not bond with other elements.

Group 17 are the HALOGEN gasses. They all have 7 valence electrons in its valence orbital.

Going left to right, the horizontal rows on the periodic table are called PERIODS. The periods number matches the number of orbitals in any atom. Period 2 elements all have 2 orbitals. Period 5 elements all have 5 orbitals. (check this now, make sure)

All elements on the left side of the DARK STAIRCASE LINE are the metals. (except hydrogen)

The nonmetals are all on the right side of this STAIRCASE line (except hydrogen).

There are nine elements TOUCHING THE STAIRCASE, and seven of them are called the Metalloids, or the semi-metals. They are metals that have some non-metallic properties, or non-metals with some metallic properties.

Two elements that touch the line, aluminum & polonium (the Al-Po exception) are not metalloids. They are both metals and have only metallic characteristics.

Sometimes the nice order of the table is slightly less than perfect, but it's the best anyone has yet devised to keep properties ordered, atomic numbers ordered, and periodicity
intact. A few asterisks are good for the soul.

MONO-ATOMIC IONS

All atoms are neutral because all atoms have the same number of positive protons as they do negative electrons. When the positives equal the negatives, you have a neutral. **Atoms like to be neutral.**

In order to bond, a metal atom atom would have to give up this neutrality and form into an ion. To do so, it loses exactly enough electrons to match a configuration of a noble gas. Metals, because of their electron configurations will only lose electrons rather than gain them, to get this noble gas configuration come about.

When an atom gains that perfect NOBLE GAS electron configuration it is said to be ISOELECTRIC to that noble gas.

Isoelectric means... having the same electron configuration as...

The sodium atom has a 2-8-1 electron configuration. If it changes to just 2-8, it has “lost” it’s outmost electron. Since is now has 11 protons but only 10 electrons, it now has a +1 overall charge.

It has become isoelectric to neon. It does not become neon, or does it become a gas.
It’s sodium, but now it is a sodium positive one ion. It’s able to bond.

The sulfur atom has a 2-8-6 electron configuration and is neutral. 16 protons= 16 electrons.
The positives equal the negatives, which is normal for all atoms.
To bond with a metal it must gain enough electrons to become isoelectric to a noble gas.

If it gains 2 electrons, its configuration becomes 2-8-8, matching argon—becoming isoelectric to argon. Sulfur does no become argon, and doesn’t become a gas either. It now has the same orbital situation as argon. It can bond, but is has a –2 overall charge. It still has 16 protons, but now 18 electrons.

Group 1 metals all have one valence electron, which each would give up to obtain a noble gas electron configuration. Na at 2-8-1 would become the sodium ion (Na^{+1})with a 2-8 configuration. Since the number of protons is constant, the ion becomes a net +1 charge.

Group 1 metals, the ALKALI METALS, all “lose” one electron to become isoelectric to a noble gas, so they all form into +1 cations.

A positive ion is called a CATION.

Group 2, the ALKALINE EARTH METALS, all have two valence electrons. Each of these metals will lose these two valence electrons when they form into cations, leaving them then with full outer orbitals, and all end up +2 cations. For example, Mg becomes the Mg^{+2} cation when it loses its 2 outer, valence electrons.

Aluminum atoms have a 2-8-3 electron configuration. In order to bond, the aluminum atom gives up three electrons to become isoelectric to neon, making the aluminum cation +3! The Al^{+3} cation has a 2-8 electron configuration.

METALS MAKE CATIONS, all metals ONLY lose electrons in order to become isoelectric to noble gases, to get that "bonding orbital system".

On the other side of the table are the nonmetals. Due to their electron arrangements, they all have more valence electrons than the metals do, it makes more sense for them to gain electrons to get a complete outer orbital. Nonmetals ONLY gain electrons to become isoelectric to noble gases.

Group 17 atoms (the HALOGENS) all have 7 valence electrons. It makes sense that they'd each gain one electron to get a complete outer orbital. Hence they form negative one (-1) ANIONS.
The gain of one negative electron by the neutral atoms yields a -1 charge extra.

Group 16 atoms all have outer, valence electron orbitals that end with 6 electrons. In order to gain the perfectly full orbital like the noble gases they need to gain 2 electrons (two extra negative electrons) They only form negative 2 ions like O^{-2}, and S^{-2}.

Group 15 atoms all have 5 valence electrons. In order to get a perfectly full orbital like a noble gas, they need to gain 3 electrons (that's 3 negative electrons extra) they form negative 3 ions like N^{-3}, and P^{-3}.

A negative ion is called an ANION.

Noble gases do not make any ions at all.

POLYATOMIC IONS

Positive cations and negative anions form for metals and nonmetals. When one atom becomes an ion these are the monoatomic ions (one atom becomes one ion).

In chemistry it is possible for small groups of atoms to unite as a single ion even though they might be 2 to 9 atoms in a small group, and they act as a single cation or anion. These are called polyatomic ions. The only poly-atomic ions that you need to be familiar with are listed in Table E.

The polyatomic ions come in cation form and in anion form.

Their names stick, we never change these names (like we change anions names to the –ide name).

There are many similar but different polyatomic ions (sulfite and sulfate, carbonate and hydrogen carbonate, chromate and dichromate, for example). Put your finger into the right box, little differences mean different things completely. USE THE CORRECT IONS, with the correct ionic charge.

Forming Ionic Compounds

Neutral metal atoms "lose" electrons, and become positive cations. They don't really lose electrons, that is not allowed. Rather, the electrons are transferred to the nonmetal atoms, which become negative anions at the same time. The nonmetals don't quite "gain" electrons. The electrons are transferred from metal to nonmetal.

The classic starter compound is sodium chloride, NaCl. It forms when one atom of sodium transfers an electron away, to a chlorine atom. The sodium atom "loses" one electron, the chlorine atom gains this electron. A perfect transfer of electrons, forming a +1 cation and a –1 anion.

The smallest part of an ionic compound is not a molecule. These ionic compounds exist only in the crystal form as solids, millions of ions connected. You could "imagine" just a single unit of an ionic compound formula, for instance "a NaCl" unit. This imaginary particle is called a single FORMULA UNIT. You can't have one, but you can imagine that it might be the smallest possible unit of an ionic compound.

Starting point	Electron transfers	Forming	→	A neutral ionic compound
Na atom 2-8-1	"loses" 1 electron	Na^{+1} cation		
	↓		Ions attract due to opposite charges	→ NaCl
Cl atom 2-8-7	"gains" 1 electron	Cl^{-1} anion		

Starting point	Electron transfers	Forming	→	A neutral ionic compound
Mg atom 2-8-2	"loses" 2 electrons	Mg^{+2} cation		
	↓		Ions attract due to opposite charges	→ MgO
O atom 2-7	"gains" 2 electrons	O^{-2} anion		2:2 ratio changes to 1:1 simple whole number ratios only

Starting point	Electron transfers	Forming	→	A neutral ionic compound
Al atom 2-8-3	"loses" 3 electrons	Al^{+3} cation		
	↓		Ions attract due to opposite charges	→ Al_2S_3
S atom 2-8-6	"gains" 2 electrons	S^{-2} anion		The smallest, simple whole number ratio that sums to neutral.
				2 Al^{+3} = +6 3 S^{-2} = – 6

John Dalton said that compounds form in simple whole number ratios. There are a limited number of possible ions from the periodic table, cations and anions. They are +1, +2, +3, and -1, -2, and -3. A neutral ionic compound can only be made from a balanced ratio of positive cations and negative anions. The only possible ratios of ions from these are as follows:

	cation charge	anion charge	ratio of cations : anions in a neutral ionic compound	Example compounds
	+1	-1	1:1	NaCl sodium chloride
	+2	-1	1:2	MgF_2 magnesium fluoride
	+3	-1	1:3	$AlBr_3$ aluminum bromide
	+1	-2	2:1	Li_2O lithium oxide
☼	+2	-2	(2:2 is not simple →) 1:1	MgO magnesium oxide
	+3	-2	2:3	Al_2S_3 aluminum sulfide
	+1	-3	3:1	K_3P potassium phosphide
	+2	-3	3:2	Ba_3N_2 barium nitride
☼	+3	-3	(3:3 is not simple →) 1:1	AlP aluminum phosphide

The "easy" way to figure out IONIC COMPOUND FORMULAS

Five Examples

Na^{+1} and Cl^{-1} → criss-cross the charges and you end up with Na_1Cl_1, but we don't write ones: NaCl

Al^{+3} and Br^{-1} → criss-cross the charges and you end up with Al_1Br_3, but we don't write ones: $AlBr_3$

Li^{+1} and O^{-2} → criss-cross the charges and you end up with Li_2O_1, but we don't write ones: Li_2O

Al^{+3} and S^{-2} → criss-cross the charges and you end up with Al_2S_3

Ca^{+2} and S^{-2} → criss-cross the charges and you end up with Ca_2S_2 but 2:2 changes to 1:1, or CaS

MOLECULAR COMPOUNDS

Are formed from 2 or more non-metals combining together. There are NO METALS, there are NO IONS, and the compounds formed are still neutral, but because they form from neutral ATOMS.

The electrons are NOT transferred from one atom/ion to another. Here the electrons are shared, they remain with their original atoms.

The smallest part of a molecular compound is called the molecule. There are specific ratios of atoms that can make real molecules. To determine how these atoms fit together (or do not fit together), we use the selected oxidation states listed in our periodic tables.

Selected oxidation states ARE NOT IONS, nor ionic charges that need to be neutral.
They are a set of numbers that allow you to determine which atom ratios are possible or impossible for molecules to form into.

The numbers are listed and come in positive or negative values. These numbers ARE NOT ionic forms, and many atoms have many different oxidation states. Many more selected oxidation states exist, not all of them are listed.

We use these numbers to see how the atoms can combine into particular ratios and make compounds that can exist. The atoms need to combine so that these oxidation numbers end up to a net zero, similar of course to ionic compounds needing to balance the + cation charges against equal but opposite - anion charges.

The oxidation numbers are a set of numbers, used to make ratios that sum to zero, but are not ions! There are NO CHARGES for molecular compounds.

EXAMPLE: Nitrogen has many oxidation numbers, oxygen just one. (see periodic table)

Nitrogen	Oxygen	What are ALL of the possible nitrogen-oxygen compounds in high school?
~~-3~~	-2	Nothing here. The negative oxidation states of nitrogen can never "balance" or sum to zero with a –2 oxidation state of oxygen, so those are not used to form N-O compounds.
~~-2~~		They're like your other shoes in your closet. You don't need them today, but you might wear them with a different outfit.
~~-1~~		They are real, but not used when bonding to oxygen, they are for bonding to some atom with a positive selected oxidation state, like chlorine, or many other atoms.
+1		N^{+1} and O^{-2} combine in a 2:1 ratio: N_2O dinitrogen monoxide
+2		N^{+2} and O^{-2} combine in a 1:1 ratio: NO nitrogen monoxide
+3		N^{+3} and O^{-2} combine in a 2:3 ratio: N_2O_3 dinitrogen trioxide
+4		N^{+4} and O^{-2} combine in a 1:2 ratio: NO_2 nitrogen dioxide
+5		N^{+5} and O^{-2} combine in a 2:5 ratio: N_2O_5 dinitrogen pentoxide

The oxidation states also tell us what is NOT possible.

Carbon trioxide CO_3 is NOT POSSIBLE. (it does look so good though, doesn't it?) ☺

The Periodic Table show us that carbon can be C^{-4}, C^{+2} or C^{+4} only, while oxygen can only be O^{-2}.

These are possible: C^{+2} and O^{-2} $\rightarrow$ CO C^{+4} and O^{-2} $\rightarrow$ CO_2

For CO_3 to exist, that would mean that C^{+6} would need to be available to allow this to happen:
C^{+6} and O^{-2} $\rightarrow$ CO_3
This cannot ever form. carbon trioxide is not a real molecular compound.

The CARBONATE ANION on Table E is CO_3^{-2}, but it's NOT a neutral molecular compound, it is a polyatomic anion made up of 4 atoms plus 2 extra electrons. It's not the same thing.

Molecular compounds ARE compounds, not ions.

When the oxidation numbers are the same sign (both negative, or both positive) we disregard them since they cannot combine in any way that would give us a net zero. They're like extra shoes, we need them for other compounds, but not now.

The DIATOMIC ELEMENTS

There are 7 elements that do not exist in the pure form as individual atoms.

They exist only in pairs, or twins. They are H, O, N, Cl, Br, I, and F (the HONClBrIF twins).

These seven elements exist in the pure form as H_2, O_2, N_2, Cl_2, Br_2, I_2, and F_2 only.

They are diatomic—two of the same atoms bonded into a molecule.

They are technically molecules, as they are made up of 2 or more non-metals bonded together by sharing electrons in a molecular bond.

Transitional Metal Cations

The transitional metals include groups 3 to 12, plus we include the metals in the "triangle" shape that are under the staircase on the Periodic Table (aluminum to thallium and to polonium). Because of their complex sub-orbitals that they have, that you do not have to learn about, they do not always follow the simple become isoelectric to a noble gas. They are able to lose any number of electrons and become stable cations. The transitional metals live in funkytown.

The oxidation numbers shown on the periodic table are PRIMARILY for making molecular compounds, where nonmetals and nonmetals bond together, the selected oxidation numbers in the transitional metal part of the table show the cations that can form.

All of group 1 metals have a +1 selected oxidation state. They also only make +1 cations.

All of group 2 metals have a +2 selected oxidation state. They also only make +2 cations.

Many transitional metals, such as scandium, zinc, yttrium, zirconium, silver, cadmium, indium, tantalum, and tungsten, etc. make ONLY ONE type of cation. That would be the same charged cation as the selected oxidation state on our periodic table.

Some transitional metal cations: Sc^{+3}, Zn^{+2}, Y^{+3}, Zr^{+4}, Ag^{+1}, Cd^{+2}, In^{+3}, Ta^{+5}, W^{+6}, etc.

Why this is is too difficult for high school chemistry, but the cation charges that are possible are in the boxes of the elements for you to point at. You can know them, but not why they can form.
These cations bond in the same way as the other "simple" cations in groups 1 and 2 do.

Many transitional metals make more than one stable type of cation, indicated by MORE THAN ONE POSITIVE oxidation state.

Titanium is # 22. It can make a +2, +3 or a +4 cation. Copper, number 29, can make a +1 or +2 cation. The cations that are possible in high school are the positive oxidation states listed on the periodic table.

Lead, element number 82 can make a Pb^{+2} or Pb^{+4} cation.
Vanadium, element number 23 forms V^{+2}, V^{+3}, V^{+4}, and V^{+5} cations.

Different ions of copper will form different ionic compounds, with different ratios cations to anions. These have different formulas, different properties, they are different compounds.

<table>
<tr><td>Copper ion</td><td>Combines with</td><td>Oxygen anion</td><td>And makes</td><td rowspan="3">We can't use the simple 1st name rule and 2nd name rule or else both of these compounds would have the same name!

They can't both be copper oxide!</td></tr>
<tr><td>Cu^{+1}</td><td>and</td><td>O^{-2}</td><td>Cu_2O</td></tr>
<tr><td>Cu^{+2}</td><td>and</td><td>O^{-2}</td><td>CuO</td></tr>
</table>

Copper ion	Combines with	Oxygen anion	And makes	names
Since we are using the Cu^{+1} cation, we insert that roman numeral into the name				
Cu^{+1}	and	O^{-2}	Cu_2O	Copper (I) oxide
Cu^{+2}	and	O^{-2}	CuO	Copper (II) oxide
Since we are using the Cu^{+2} cation, we insert that roman numeral into the name				

In this next example, chromium makes 3 different cations, and they all combine to chlorine in different ratios, making three different chromium chloride compounds, that all have different properties, different formulas, and different names.			
cation	anion	formula	name
Cr^{+2}	Cl^{-1}	$CrCl_2$	chromium (II) chloride SAY: chromium roman numeral 2 chloride
Cr^{+3}	Cl^{-1}	$CrCl_3$	chromium (III) chloride SAY: chromium roman numeral 3 chloride
Cr^{+6}	Cl^{-1}	$CrCl_6$	chromium (VI) chloride SAY: chromium roman numeral 6 chloride

In this last example, silver, element 47, only makes one cation, Ag^{+1}. It can only bond to a nonmetal in one way. We'll use fluorine for example. Since there is just one cation possible, no roman numeral is needed. It's got to be the Ag^{+1} cation, there are no other silver ions possible.			
cation	anion	formula	name
Ag^{+1}	F^{-1}	AgF	silver fluoride no roman numeral necessary

IONIC COMPOUNDS

Cations + anions form neutral compounds

Only metals form cations.
Only non-metals form anions

Ionic compounds are all neutral. The charges of the cations to anions must balance. Simple whole number ratios only. Use the ion charges to “criss-cross” the formulas.

Ions transfer electrons from metal to nonmetals, These ions bond with IONIC BONDS.

Metals make cations that have lost 1, 2, or 3 electrons. This makes them +1, +2, or +3

Nonmetals make anions that have gained 12, 2, or 3 electrons. This makes them –1, –2, or –3

The smallest part of an ionic compound is called a FORMULA UNIT.
Ionic compounds only exist as solids in a giant crystal form. A single grain of NaCl has millions of ions. You can’t really have just one (like Lay’s Potato Chips!) You can have a single unit of an ionic compound, but only theoretically.

NAMING IONIC COMPOUNDS

1st name rule: say the name of the metal cation.

2nd name rule: say the name of the anion*

*If the anion is MONO-ATOMIC, change the anion atomic name to end in -ide.

*If the anion is POLY-ATOMIC, don't change the polyatomic ion name from Table E

*If the cation is polyatomic, don’t change the name of that ion from Table E.

*If transitional metal is involved and it makes more than one cation use a Roman Numeral to distinguish which cation you are using.

MOLECULAR COMPOUNDS

Have no ions. Atoms combine as atoms.
NO METALS in any molecular compounds.

Bonds between 2 or more Non-Metals form into molecules.

Atoms combine together to form neutral molecular compounds

They are neutral because they form from neutral atoms.

The ratios are also very particular, and we use the selected oxidation numbers to determine what these ratios are. We make “T” charts to figure out which compounds are possible.

The electrons are shared by the atoms that bond into molecules. There is no transfer of electrons from one atom to another.

The smallest part of a molecular compound is called a molecule.

NAMING MOLECULAR COMPOUNDS

1st name rule: Single atoms get called by their name, multiple atoms get a prefix.

2nd name rule: 2nd atom always gets a prefix.

These rules are clear if you remember the names & formulas of these compounds:
CO, CO_2, H_2O. Carbon monoxide, carbon dioxide, and dihydrogen monoxide.

The prefixes from 1—10: mono, di, tri, tetra, penta, hexa, hepta, octa, nona, and deca.

Properties of Metals, non-metals, metalloids, ionic & molecular compounds

property	metals	nonmetals	metalloids	ionic compounds	molecular compounds
luster	yes	no	some have luster, like silicon	no	not usually
malleable, ductile	yes, yes	no, mostly brittle	few	no, mostly brittle	no, mostly brittle
conduct heat	yes	no	no	no	no
conduct electricity	yes	no	some can, such as silicon	no	no
form cations or anions?	cations only	anions only	Only these are possible in our class As^{-3} Ge^{+2} Ge^{+4} Sb^{+3} Sb^{+5} Te^{-2}	formed from cations & anions, but are neutral	none
smallest particle	atoms	atoms or molecules	atoms	formula units	molecules
density	fairly high	fairly low	varies, see table S	medium high	varies
melting point	relatively high	relatively lower	varies	extremely high	low compared to metals or ionic compounds
phase at room temp	solid	solid, liquid, or gas	solid	solid	solid