NS G4 (pg 1 of 2) Ionic Compounds – Using Roman #'s

Name_____Per___

Hopefully you did your homework and practice by reading NS G3 and doing P G2 – Binary Ionic Compounds
You would have worked with writing and naming ionic compounds involving elements in Groups 1, 2, & 3 combining with elements in Groups 5, 6, & 7

So what about the rest of the metals?

Lets consider naming ionic compounds involving the transition metals and the "other six" metals under the staircase ($_{31}$Ge, $_{50}$Sn, $_{51}$Sb, $_{82}$Pb, $_{83}$Bi, $_{84}$Po).

Because these elements are metals, their electrons are lose and can be stolen, but because they have "d" electrons (those 10 electrons stuffed deep inside) these elements can lose varying amounts of electrons producing a variety of ions with different charges. The good news is that you do NOT have to figure out what ions these elements are going to form. You just need to be able to write formulas and name compounds with enough information.

What to do with an atom such as lead? This element is able to lose either 2 or 4 valence electrons and thus it can form, 2+ or 4+ ions. Since you do not need to predict which of these two options occurs, the name of a lead chloride compound must give more information so you can tell which lead ion has formed. That information comes in the form of a Roman Numeral indicating the charge on the lead ion.

The atoms	The atoms	The rational for the compound	The formula	The name
Pb & Cl	$Pb^{+?}$ Cl^{-}	Without being told the charge on the lead, you would have no way of knowing what charge is formed.		Lead(??) chloride

You do not know what charge lead will form, so you need more information. This comes in the form of a Roman numeral which tells you the + charge on a single lead ion lead.

The name	The atoms	The ions	The rational for the compound	The formula
lead(II) chloride	Pb & Cl	Pb^{+2} Cl^{-}	Two chlorines (1− each) needed to make a total of 2− to balance the one (2+) lead ion	$PbCl_2$
lead(IV) chloride	Pb & Cl	Pb^{4+} Cl^{-}	Since the lead is 4+ in this compound, four chlorines (1− each) are needed balance the charge	$PbCl_4$
chromium(III) oxide	Cr & O	Cr^{3+} O^{2-}	The charge on the chromium is 3+, thus a combination of 2 chromiums and 3 oxides each provide a balanced total of 6+/6− charge. You can use the criss cross method.	Cr_2O_3
tin (IV) oxide	Sn & O	Sn^{4+} O^{-2-}	The charge on the tin is 4+, thus with two oxides (2− each) will produce a balanced total of 4+/4− charge. If you criss-cross this time the tin and oxygen would combine as Sn_2O_4 however you must write the chemical formula in the *lowest* whole number ratio, thus reduce.	SnO_2

Working the process "backwards"

This can get tricky when given the formula and asked to determine the correct name. You'll need to work backwards to determine the Roman # by starting with the charge of the nonmetal, knowing that the total negative charge must be equal to the total positive change.

The formula	The charge on nonmetal	The total negative charge	The process for determining the Roman numeral for the name	The name
NiF_3	F^-	three F^- = 3− total	Since the fluoride carries a −1 charge, three of them equals −3, and the nickel ions total charge must be opposite in sign, but equal in magnitude, thus the Ni ion must be 3+, resulting in nickel(III) oxide	nickel(III) fluoride
CrP_2	P^{3-}	two P^{3-} = 6−	Since the phosphide carries a 3− charge, two of them equal 6−, and the chromium ions total charge must be opposite in sign, but equal in magnitude, thus Cr must be 6+, resulting in chromium(VI) oxide	Chromium (VI) phosphide
Cu_2S	S^{2-}	one S^{2-} is of course a charge of 2−	Since the sulfide carries a 2− charge, and the copper ions total charge must be opposite in sign, 2+, but there are two copper ions causing 2+ total, thus the each copper must be 1+, resulting in copper(I) sulfide	Copper (I) sulfide

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A few reminders:

- You will need to know your Roman numerals 1 through 7: I, II, III, IV, V, VI, VII
- You will need to memorize that zinc and cadmium are always 2+, and that silver is always 1+
- The darkened Noble gases have a complete energy level and do not have the ability to gain or lose electrons thus do not naturally form compounds.
- The grayed out hydrogen, boron, carbon and silicon do not generally engage in formation of ions, but instead share their valence electrons. We will study this covalent bonding later in the course.