

## Chapter 3 Compounds and Bonding

**Nomenclature and formulas of ionic compounds.** Notice that some periodic tables have a staircase on the right side of the chart. This artificial boundary represents a very approximate division of the metallic like elements on the left and the non-metals on the right. Elements close to the staircase have properties somewhere in between metals and non-metals and are called metalloids. For example, metals usually are good conductors of electricity, non-metals are poor conductors and are often used as insulators and metalloids are often semiconductors and are crucial for use in today's high tech world. Refer to the table below for properties of metals and non-metals.

### Metals

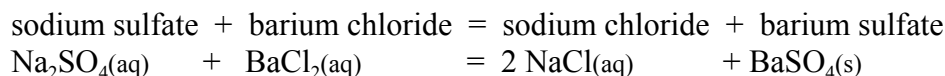
usually form ions with positive charges  
 conductor of electricity  
 conductor of energy  
 lustrous  
 ductile (can be hammered)  
 malleable (pulled into wires)  
 opaque as thin sheet  
 solids (except Hg)  
 Periodic table of formulas (see last page of this chapter)

### Non-metals

usually form ions with negative charges  
 insulator  
 insulator  
 non-lustrous  
 non-ductile  
 brittle  
 transparent as thin sheet  
 solids, liquids, gases

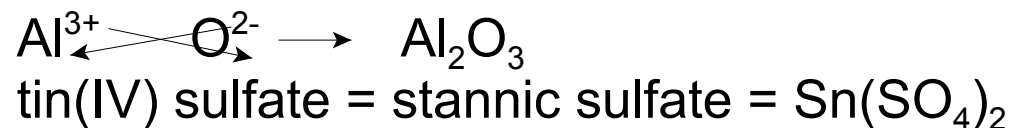
To be able to communicate in chemistry, we need to be familiar with the nomenclature guidelines for compounds. When metals react by donating electrons (oxidation) to form positively charged ions (called cations), the ion retains the name of the element. However, some metals form ions of different charges. For example, iron can lose 2 or 3 electrons and the name needs to distinguish between these options. As a result there are two compounds of iron with chlorine, iron(II) chloride or  $\text{FeCl}_2$  and iron(III) chloride or  $\text{FeCl}_3$ . Unfortunately another naming system developed historically that still is commonly used today. For iron, the names of the ions are derived from its Latin name ferrum. The names of the two compounds with chlorine for  $\text{FeCl}_2$  and  $\text{FeCl}_3$  are ferrous chloride and ferric chloride respectively. Ferrum is also the reason that the symbol for iron is Fe.

For the non-metals which tend to gain electrons (reduction) when reacting, the ending of the element is usually changed to ide as with chloride. However, for pronunciation reasons, the spelling is sometimes modified. From the preceding paragraphs, you should have been able to gain a sense of the way the formulas of metal non-metal compounds are named and how the formulas are determined. The table that follows contains a list of the most commonly encountered positive ions (cations) and negative ions (anions) and the charges on the ion. You will quickly notice that many are not simply the ion of an element but contain two or more elements. These ions are called polyatomic ions. Within the ion, the elements are bonded together with covalent bonds. These ions deserve special attention because for many reactions, they stay intact throughout the reaction. In other words, when sodium sulfate reacts with barium chloride, the positive and negative ions switch partners in what is called a double replacement reaction but the sulfate ion is not changed in the reaction.



The equation above provides additional information that will be discussed further when we learn more about how to write balanced reactions (equations). The (aq) indicates the chemical is dissolved in water and (s) indicates that because the compound is not soluble in water, a solid precipitates out of the solution. This is actually the incentive for the reaction to occur and drives the reaction from left to the right.

To determine the formula of a compound composed a positive and negative ion, the basic goal is to make sure the result has no net charge or is neutral. In addition, the ratio of the ions as indicated by the subscripts must be reduced to the lowest whole number ratio. For sodium and chloride, this is easy as the ions have equal but opposite charges (NaCl). The same is true for sodium and nitrate (NaNO<sub>3</sub>). With sodium and oxide, two sodium ions are required to provide a neutral compound (Na<sub>2</sub>O). For aluminum oxide, coupling the Al<sup>3+</sup> ion with the O<sup>2-</sup> ion requires a little more thought. For the 3 and 2, find the lowest common multiplier or 6 and then divide by the charge. 6/3 yields 2 aluminum ions and 6/2 yields 3 oxygen ions resulting in a formula of Al<sub>2</sub>O<sub>3</sub>. It is possible to circumvent the thinking process by using a criss cross method as shown below. Remember to reduce the resulting subscripts to the lowest whole number ratio. Also, if the formula includes more than one polyatomic ion, the polyatomic ion must be enclosed in parentheses.



2-1. Determine the formulas of the following compounds: a. potassium sulfide b. silver oxide  
c. iron(III) chloride (or ferric chloride) d. ammonium nitrate e. calcium phosphate f. zinc oxide  
g. magnesium hydroxide h. copper(II) sulfate (or cupric sulfate) i. aluminum sulfate

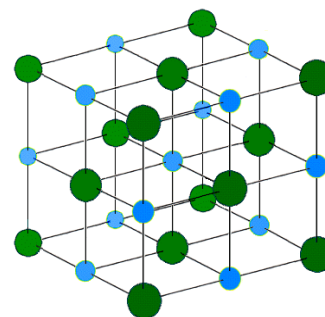
**A. Positive Ions**

<u>+ 1</u>	Ammonium	$\text{NH}_4^+$	Mercury(I) (mercurous)	$\text{Hg}_2^{2+}$
	Copper(I) (cuprous)	$\text{Cu}^+$	Potassium	$\text{K}^+$
	Hydrogen	$\text{H}^+$	Silver	$\text{Ag}^+$
	Hydronium	$\text{H}_3\text{O}^+$	Sodium	$\text{Na}^+$
	Lithium	$\text{Li}^+$		
<u>+ 2</u>	Barium	$\text{Ba}^{2+}$	Magnesium	$\text{Mg}^{2+}$
	Cadmium	$\text{Cd}^{2+}$	Manganese(II) (manganous)	$\text{Mn}^{2+}$
	Calcium	$\text{Ca}^{2+}$	Mercury(II) (mercuric)	$\text{Hg}^{2+}$
	Cobalt(II) (cobaltous)	$\text{Co}^{2+}$	Nickel(II) (nickelous)	$\text{Ni}^{2+}$
	Copper(II) (cupric)	$\text{Cu}^{2+}$	Strontium	$\text{Sr}^{2+}$
	Iron(II) (ferrous)	$\text{Fe}^{2+}$	Tin(II) (stannous)	$\text{Sn}^{2+}$
	Lead(II) (plumbous)	$\text{Pb}^{2+}$	Zinc	$\text{Zn}^{2+}$
<u>+ 3</u>	Aluminum	$\text{Al}^{3+}$	Cerium(III) (cerous)	$\text{Ce}^{3+}$
	Antimony(III)	$\text{Sb}^{3+}$	Chromium(III) (chromic)	$\text{Cr}^{3+}$
	Arsenic(III)	$\text{As}^{3+}$	Iron(III) (ferric)	$\text{Fe}^{3+}$
	Bismuth(III)	$\text{Bi}^{3+}$		
<u>+ 4</u>	Lead(IV) (plumbic)	$\text{Pb}^{4+}$	Tin(IV) (stannic)	$\text{Sn}^{4+}$
<u>+ 5</u>	Antimony(V)	$\text{Sb}^{5+}$	Bismuth(V)	$\text{Bi}^{5+}$
	Arsenic(V)	$\text{As}^{5+}$		

**B. Negative ions**

<u>- 1</u>	Acetate	$\text{C}_2\text{H}_3\text{O}_2^-$	Hydrogen sulfate (bisulfate)	$\text{HSO}_4^-$
	Bromate	$\text{BrO}_3^-$	Hydrogen sulfite (bisulfite)	$\text{HSO}_3^-$
	Bromide	$\text{Br}^-$	Hydroxide	$\text{OH}^-$
	Chlorate	$\text{ClO}_3^-$	Hypochlorite	$\text{ClO}^-$
	Chloride	$\text{Cl}^-$	Iodate	$\text{IO}_3^-$
	Chlorite	$\text{ClO}_2^-$	Iodide	$\text{I}^-$
	Cyanate	$\text{NCO}^-$	Nitrate	$\text{NO}_3^-$
	Cyanide	$\text{CN}^-$	Nitrite	$\text{NO}_2^-$
	Fluoride	$\text{F}^-$	Perchlorate	$\text{ClO}_4^-$
	Hydride	$\text{H}^-$	Permanganate	$\text{MnO}_4^-$
	Hydrogen carbonate (bicarbonate)	$\text{HCO}_3^-$	Thiocyanate	$\text{SCN}^-$
<u>- 2</u>	Carbonate	$\text{CO}_3^{2-}$	Sulfate	$\text{SO}_4^{2-}$
	Chromate	$\text{CrO}_4^{2-}$	Sulfide	$\text{S}^{2-}$
	Dichromate	$\text{Cr}_2\text{O}_7^{2-}$	Sulfite	$\text{SO}_3^{2-}$
	Oxalate	$\text{C}_2\text{O}_4^{2-}$	Tetrathionate	$\text{S}_4\text{O}_6^{2-}$
	Oxide	$\text{O}^{2-}$	Thiosulfate	$\text{S}_2\text{O}_3^{2-}$
	Persulfate	$\text{S}_2\text{O}_8^{2-}$		
<u>- 3</u>	Ferricyanide	$\text{Fe}(\text{CN})_6^{3-}$	Phosphate	$\text{PO}_4^{3-}$
<u>- 4</u>	Ferrocyanide	$\text{Fe}(\text{CN})_6^{4-}$		

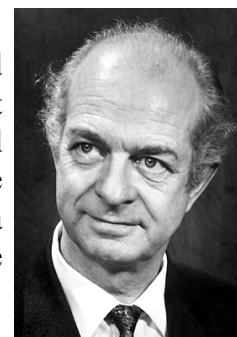
The reverse process or the naming of a compound given the formula involves naming the positive and negative ions. The number of ions present in the formula is not included in the name but as seen above can be determined by following the guidelines for the determination of the formula. The complication comes in for positive ions that have more than one possible oxidation state (charge). The oxidation state must be specified in the name or the reader will not know which compound was meant. There are two naming systems, the stock method and the common name. Using iron as an example [e.g., iron(II) or ferrous and iron(III) or ferric], either the oxidation state (II) or (III) must be included in the name or the common name must be used. Iron sulfate would be an inadequate name for  $\text{Fe}_2(\text{SO}_4)_3$ . Either iron(III) sulfate or ferric sulfate can be used. The oxidation state can be determined from the formula by using logic and your knowledge of the charge of one of the partners. It should be known because of very common usage that sulfate has a negative 2 charge. By examining the formula, you should be able to deduce that iron has to be +3 for the formula to be a neutral compound. An alternative is to use the reverse criss cross method but remember that on rare occasions the subscripts have been reduced. This can lead you to inadequate formulas with the reverse criss cross method.



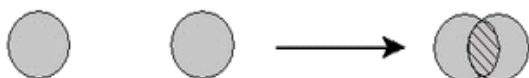
**Molecules.** While some chemists refer to  $\text{NaCl}$  as a molecule, the word molecule does not appropriately describe  $\text{NaCl}$  or other ionic compounds. Like the atom is the smallest unit that retains the properties of an element, a molecule is the smallest unit that retains the properties of a compound. However, sodium chloride exists predominately as a crystal and each sodium is surrounded by 6 chlorides and each chloride is surrounded by 6 sodium ions. The array pictured to the right repeats itself trillions upon trillions of times in a typical salt crystal. The properties of  $\text{NaCl}$  are a consequence of this array. A simple  $\text{NaCl}$  unit would not have the same properties and should not be called a molecule and but could be called a formula unit. The same holds for other ionic compounds. This is the reason that the formulas for ionic compounds are reduced to the lowest whole number ratios. The formula gives simply the ratio and not the number of atoms of each element as molecules of ionic compounds do not exist.

2-1. Determine the name(s) for the following formulas: a.  $\text{NaBr}$  b.  $\text{Ca}(\text{NO}_3)_2$  c.  $(\text{NH}_4)_2\text{S}$   
d.  $\text{K}_2\text{CO}_3$  e.  $\text{Cu}(\text{OH})_2$  f.  $\text{Mg}(\text{C}_2\text{H}_3\text{O}_2)_2$  g.  $\text{HgCl}_2$  h.  $\text{Ag}_2\text{SO}_4$  i.  $\text{Li}_3\text{PO}_4$  j.  $\text{SnO}_2$  k.  $\text{ZnI}_2$  l.  $\text{CrCl}_3$

**Covalent bonds.** The discussion above focused on nomenclature and formulas of metal to non-metal compounds involving ionic bonds. By far, most compounds that have been characterized involve bonds between carbons and between carbon and other non-metals. Bonding in these cases does not involve transfer of electrons and instead electrons are shared by the two partners. As a result, both partners have orbitals that are filled and energetically more stable than if the atoms existed without bonding.



Linus Pauling



Formation of covalent bond by overlap of two s orbitals.

**Electronegativity.** All physical and chemical process involve energy changes. There are two energy terms, enthalpy and an entropy related portion. The energetics of processes will receive a little more attention later but suffice it to say that the energy issues for an element to gain electrons become more favorable going across and up the periodic table with fluorine the most favorable. The two time Nobel prize winner, Linus Pauling, developed an electronegativity scale as a measure of the tendency for elements to gain electrons. Fluorine has the highest value of about 4 and alkali metals (group 1) have the lowest values near 1.

<i>H</i> = 2.20						
<i>Li</i> = 0.98	<i>Be</i> = 1.57	<i>B</i> = 2.04	<i>C</i> = 2.55	<i>N</i> = 3.04	<i>O</i> = 3.44	<i>F</i> = 3.98
<i>Na</i> = 0.93	<i>Mg</i> = 1.31	<i>Al</i> = 1.61	<i>Si</i> = 1.90	<i>P</i> = 2.19	<i>S</i> = 2.58	<i>Cl</i> = 3.16
<i>K</i> = 0.82	<i>Ca</i> = 1.00	<i>Ga</i> = 1.81	<i>Ge</i> = 2.01	<i>As</i> = 2.18	<i>Se</i> = 2.55	<i>Br</i> = 2.96
<i>Rb</i> = 0.82	<i>Sr</i> = 0.95	<i>In</i> = 1.78	<i>Sn</i> = 1.96	<i>Sb</i> = 2.05	<i>Te</i> = 2.1	<i>I</i> = 2.66

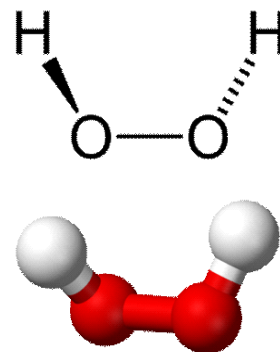
It stands to reason that big electronegativity differences could lead to a transfer of an electron and as a result the formation of an ionic bond. As a very rough guideline, when the electronegativities of two bonding partners differ by more than 1.7, the resulting bond should be considered ionic. If the electronegativity difference is less than 0.4, the electrons are shared about equally and the bond is covalent. Between 0.4 and 1.7, the bond is still covalent. However, the electrons are shared unevenly and the electron density is greater near the more electronegative of the two elements resulting in a polar covalent bond. For our use, since electronegativities are usually not memorized or even handy, it is sufficient to remember that bonds between the same element (carbon to carbon) and carbon to hydrogen bonds are covalent. Bonds between metals (and the polyatomic ammonium) and non-metals (and polyatomic ions) are ionic and the remaining bonds (between different non-metals) are polar covalent. The polarity of bonds is an extremely important concept as the degree of polarity plays an important role in determining chemical reactivity and physical properties including solubility. Bonds between metals are called metallic bonds and will not be discussed in this course.

**Nomenclature of compounds containing non-metals other than those with multiple carbon carbon bonds.** The element closest to fluorine (the more electronegative elements) is named last and the ending is again ide. The more electropositive element is named first. Except for cases where there is only one atom of the first element, prefixes are use to indicate how many atoms of each element are present.

CO	carbon monoxide
CO <sub>2</sub>	carbon dioxide
N <sub>2</sub> O	dinitrogen monoxide
NO <sub>2</sub>	nitrogen dioxide
N <sub>2</sub> O <sub>4</sub>	dintrogen tetroxide (Note that the subscripts for non-metal to non-metals need to reflect the number of each atom in the molecular formula and should not be reduced. Ionic compounds do not form molecules and therefore the formula indicates merely the ratio.)

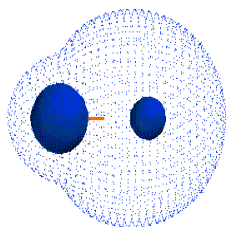
Determination of the formula from the name should be straightforward as the name gives the number of atoms of each element present in the formula. Unlike ionic compounds where the subscripts have been reduced to the lowest whole number ration, the subscripts give not only the ratio but the actual

number of the atoms of the atoms of each element in the formula. Nitrogen dioxide and dinitrogen tetroxide in the table above have the same ratio (1:2) of nitrogen to oxygen but because of their molecular nature are two different compounds with different properties. In the figures for dihydrogen dioxide (commonly called hydrogen peroxide), note that the formula is  $H_2O_2$  and the subscripts cannot be reduced because of the molecular nature of the substance.

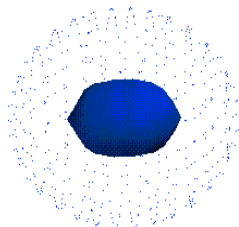
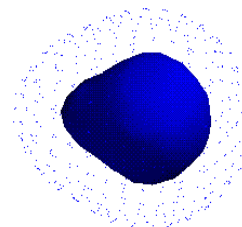


**Diatomic molecules.** You might have noticed that when a symbol for nitrogen, oxygen and the halogens (group 7A elements) is used for these elements in their elemental state (in nature uncombined with another element), the formulas are written  $H_2$ ,  $N_2$ ,  $O_2$ ,  $F_2$ ,  $Cl_2$ ,  $Br_2$  and  $I_2$ . These elements are found to be diatomic molecules in nature. The reason for this is fairly simple. All of these elements and especially fluorine and oxygen are very electronegative. They each can achieve greater stability by acquiring enough electrons to achieve an inert gas structure of electrons. For fluorine, this means gaining one electron. If fluorine forms a covalent bond to another fluorine, both fluorine atoms has filled orbitals and the result is more stable than the isolated atoms. However, it is important to recognize that these elements can achieve even more stability by bonding with a less electronegative partner thus acquiring greater electron density. As a result, all of these elements are very reactive.

**Polar covalent bonds and polar molecules.** Other than carbon - carbon, carbon - hydrogen bonds and bonds between identical elements (such as the diatomic molecules discussed in the previous paragraph) almost all non-metal to non-metal bonds due to different electronegativities involve an unequal sharing of electrons. This gives rise to more electron density on one partner than the other. For hydrogen fluoride for instance, the electrons in the covalent bond that holds the molecule together spend more time in the vicinity of the fluorine than near the hydrogen. This gives rise to a dipole (separation of charge) in the molecule with a partial positive charge on the hydrogen and a partial negative charge on the fluorine. The resulting polar covalent bond strongly influences the properties of HF. The molecules  $LiH$ ,  $H_2$  and  $HF$  are shown below. For  $LiH$ ,  $Li$  is bigger than  $H$  but the greater electronegativity of  $H$  results in more electron density around the hydrogen. For  $H_2$ , the electrons are shared equally and a symmetrical electron cloud results. For  $HF$ ,  $H$  is smaller than  $F$  and  $F$  is more electronegative than  $H$  resulting in a greater electron density around the fluorine.



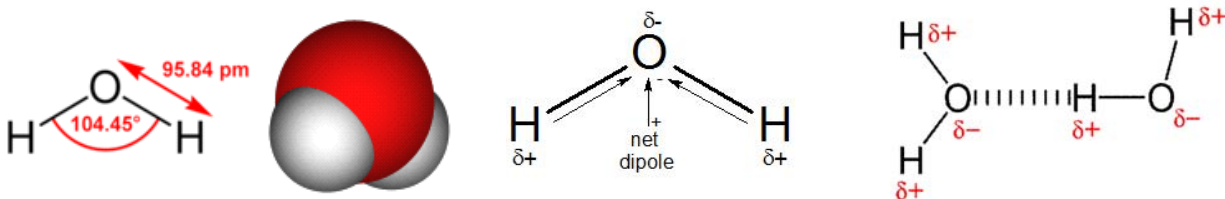
LiH

 $H_2$ 

HF

For life on earth, water is clearly of utmost importance. For  $H_2O$ , the oxygen is more electronegative than the hydrogen. This results in two dipoles from each of the hydrogens to the oxygen. The net result of the two dipoles depends on the shape of water which could be linear or bent. If water was linear, the two dipoles would cancel each other, water would not be polar and life

as we know it would probably not exist. However, we will find that the shape of water is predictable from basic concepts. The shape is predicted to be bent with a bond angle a little over  $100^\circ$ . Experimental evidence is completely consistent with a bent shape for water with a bond angle of  $104.45^\circ$ . The figures on the right show the partial charges (the Greek letter  $\delta$  is used to represent partial) that result from the polarity of the O-H bond. As shown in the third figure, the two dipoles from hydrogen to oxygen add up to a significant net dipole from the middle of the hydrogens towards the oxygen. Since positive and negative charges attract each other, there is a strong intermolecular attraction between adjacent water molecules that extends throughout the liquid.



The properties of water and the consequences of these strong attractions will be discussed in the next chapter.

**Organic chemistry.** Although there are 118 known elements, only one element, carbon has its own branch of chemistry. Compounds containing carbon comprise in addition to water almost all of the substances in your body. There are many reasons for this. Although some science fiction authors have based their stories on silicon based life forms, it is very doubtful that such life forms could exist anywhere in the Universe. Carbon has unique and ideal properties that make it ideally suited for life. Carbon is a group 4 element and bonds to many other elements. More importantly, many of the bonds to other elements are strong. Having 4 electrons in its outer orbitals, carbon achieves stability by forming 4 bonds. This produces a mind boggling number of possible compounds. Perhaps most importantly, carbon forms strong bonds to itself. This results in molecules with varying numbers of carbons linked together. A very common result is that carbon forms extremely long molecules with hundreds or even thousands of carbons in a chain (sometimes linked with other non-metals). These molecules are called polymers and usually contain many simpler molecules attached together one after another. Examples are proteins, RNA, DNA, carbohydrates, nylon, rubber, polyester. Organic chemistry has its own nomenclature system that will not be covered now.

# PERIODIC TABLE OF COMPOUNDS WITH H\* AND Cl

1 H HCl																2 He	
3 Li LiH LiCl	4 Be BeH <sub>2</sub> BeCl <sub>2</sub>											5 B B <sub>2</sub> H <sub>6</sub> * BCl <sub>3</sub> *	6 C CH <sub>4</sub> CCl <sub>4</sub>	7 N NH <sub>3</sub> NCl <sub>3</sub>	8 O H <sub>2</sub> O	9 F HF	10 Ne
11 Na NaH NaCl	12 Mg MgH <sub>2</sub> MgCl <sub>2</sub>											13 Al AlH <sub>3</sub> AlCl <sub>3</sub>	14 Si SiH <sub>4</sub> * SiCl <sub>4</sub>	15 P PH <sub>3</sub> * PCl <sub>3</sub> *	16 S H <sub>2</sub> S SCl <sub>2</sub>	17 Cl HCl	18 Ar
19 K KH KCl	20 Ca CaH <sub>2</sub> CaCl <sub>2</sub>	21 Sc ScH <sub>3</sub> * ScCl <sub>3</sub>	22 Ti TiH <sub>2</sub> TiCl <sub>4</sub> *	23 V VH* VCl <sub>4</sub> *	24 Cr CrCl <sub>3</sub> *	25 Mn MnCl <sub>2</sub> *	26 Fe FeCl <sub>3</sub> *	27 Co CoCl <sub>2</sub> *	28 Ni NiCl <sub>2</sub>	29 Cu ZnH <sub>2</sub> CuCl <sub>2</sub> *	30 Zn [GaH <sub>3</sub> ] <sub>2</sub> ZnCl <sub>2</sub>	31 Ga GeH <sub>4</sub> * Ga <sub>2</sub> Cl <sub>6</sub> *	32 Ge AsH <sub>3</sub> * GeCl <sub>4</sub> *	33 As H <sub>2</sub> S AsCl <sub>3</sub> *	34 Se HBr SeCl <sub>2</sub> *	35 Br	36 Kr
37 Rb RbH RbCl	38 Sr SrH <sub>2</sub> SrCl <sub>2</sub>	39 Y YH <sub>3</sub> * YCl <sub>3</sub>	40 Zr ZrH <sub>2</sub> ZrCl <sub>4</sub> *	41 Nb NbCl <sub>5</sub> *	42 Mo MoCl <sub>6</sub> *	43 Tc TcCl <sub>4</sub> *	44 Ru RuCl <sub>3</sub> *	45 Rh RhCl <sub>3</sub>	46 Pd PdCl <sub>2</sub>	47 Ag AgCl	48 Cd CdH <sub>2</sub> CdCl <sub>2</sub>	49 In InH InCl <sub>3</sub> *	50 Sn SnH <sub>4</sub> SnCl <sub>4</sub> *	51 Sb SbH <sub>3</sub> SbCl <sub>3</sub> *	52 Te TeCl <sub>2</sub>	53 I HI ICl	54 Xe
55 Cs CsH CsCl	56 Ba BaH <sub>2</sub> BaCl <sub>2</sub>	57 La LaH <sub>3</sub> * LaCl <sub>3</sub>	72 Hf HfH <sub>2</sub> HfCl <sub>4</sub>	73 Ta Ta <sub>2</sub> H TaCl <sub>5</sub> *	74 W WCl <sub>6</sub> *	75 Re ReCl <sub>4</sub> *	76 Os OsCl <sub>3</sub> *	77 Ir IrCl <sub>3</sub> *	78 Pt PtCl <sub>4</sub> *	79 Au AuCl*	80 Hg HgH <sub>2</sub> HgCl <sub>2</sub> *	81 Tl TlCl <sub>3</sub> *	82 Pb PbH <sub>4</sub> PbCl <sub>4</sub> *	83 Bi BiH <sub>3</sub> BiCl <sub>3</sub>	84 Po PoH <sub>2</sub> PoCl <sub>2</sub> *	85 At	86 Rn
87 Fr	88 Ra	89 Ac AcH <sub>2</sub> AcCl <sub>3</sub>	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 2003	114 Fl	115 2003	116 Lv	117 2010	118 2006
			58 Ce CeH <sub>2</sub> CeCl <sub>3</sub>	59 Pr PrCl <sub>3</sub>	60 Nd NdCl <sub>3</sub> *	61 Pm PmCl <sub>3</sub>	62 Sm SmCl <sub>3</sub> *	63 Eu EuCl <sub>3</sub> *	64 Gd GdCl <sub>3</sub>	65 Tb TbCl <sub>3</sub>	66 Dy DyCl <sub>3</sub> *	67 Ho HoCl <sub>3</sub>	68 Er ErH <sub>3</sub> ErCl <sub>3</sub>	69 Tm TmCl <sub>3</sub> *	70 Yb YbCl <sub>3</sub>	71 Lu LuCl <sub>3</sub>	
			90 Th ThH <sub>2</sub> ThCl <sub>4</sub>	91 Pa PaCl <sub>5</sub> *	92 U UH UCl <sub>6</sub> *	93 Np NpCl <sub>4</sub> *	94 Pu PuCl <sub>3</sub>	95 Am AmCl <sub>3</sub> *	96 Cm CmCl <sub>3</sub>	97 Bk BkCl <sub>3</sub>	98 Cf CfCl <sub>3</sub> *	99 Es EsCl <sub>3</sub> *	100 Fm	101 Md	102 No	103 Lr	

\*In some cases the hydrogen is more electronegative (hydrides) than its partner and in other cases it is more electropositive. An asterisk after a compound indicates that other combinations of the element with H or Cl exist. Formulas obtained from: <http://www.webelements.com/>