Ch. 7 Notes ~ Formation of Compounds

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics.*

7.1 notes

I. Valence Electron Review

A. **Valence electrons**—the *electrons in the highest energy level* of at atom
1) tells *how many are available for bonding*
2) tells *how many e- will be gained or lost in forming ions*

B. Review electron dot diagrams (Lewis structures)
"right, left, up, down, top all the way around – counterclockwise"

GROUP #:	VALENCE # out of 8:	USUALLY, WHEN FORMING IONS:
Group IA (1)	1	loses 1
Group IIA (2)	2	loses 2
Group IIIA (13)	3	loses 3
Group IVA (14)	4	can lose, gain, or share
Group VA (15)	5	gains 3
Group VIA (16)	6	gains 2
Group VIIA (17)	7	gains 1
Group VIIIA (18)	8	does not form ions

II. Ion Formation

A. *atomic neutrality: atoms are neutral (net charge of zero) # protons = # electrons*

B. **ion**—*a charged atom(s) or a charged group of atoms*
1) formed by gain or loss of electrons
2) **cation**— *a positive ion formed by losing electrons*
3) **anion**— *a negative ion formed by gaining electrons*
   a) *common ending is —IDE* (chloride, bromide, iodide...)
   b) *ions of Group VIIA*, the halogens, are called *halide ions*

CP AN **Cations Positive, Anions Negative** "Cat People Are Nice"
YOU CAN'T LOSE OR GAIN PROTONS TO FORM AN ION !

4) an ion has different properties than its element (Na atom vs. Na^+ ion)
5) **isoelectronic**—having the same number of electrons
   . the reason why an atom loses or gains electrons is to be isoelectronic with the closest Noble Gas.

C. ionic charges (oxidation numbers) of the groups

Charge Chant Song:
" +1, +2, +3, mixed, -3, -2, -1, 0 …
+2 in the middle, unless they tell you otherwise"

(Note – you can write charges with the sign before or after the number. The norm is to use the sign after the number, but the song sounded better with the order reversed. In textbooks, you will not see the number 1 used with a charge. Only + or - will be shown for 1+ or 1-.)

Group number:	IA	IIA	IIIA	IVA	VA*	VIA*	VIIA*	VIIIA
	1	2	13	14	15	16	17	18
Main ionic charge:	1+	2+	3+	M	3-	2-	1-	none

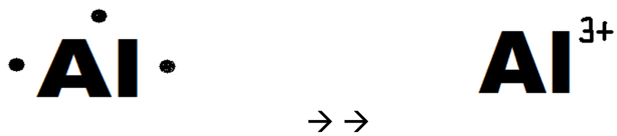
M most of the Group IVA don't usually form ions; when they do, there are mixed charges possible
* when applicable

D. **Octet Rule**

1) *reactivity of atoms is based on achieving a complete octet (8/8)=* s^2p^6
2) *atoms in compounds tend to have a stable* **Noble Gas configuration**
3) exception – **pseudo-Noble gas configuration**: $s^2p^6d^{10}$ (see next section)

E. examples of ions

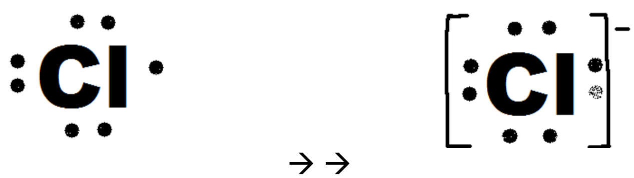
E1) Aluminum loses three electrons when its ion is formed. Show what happens with "before and after" electron dot diagrams. The aluminum ion is isoelectronic with which Noble gas?



13 p^+
13 e^- (3 valence)

13 p^+
10 e^- (loses three to empty the original valence)
isoelectronic with Ne

E2) Chlorine gains one electron when its ion is formed. Show what happens with "before and after" electron dot diagrams. The chloride ion is isoelectronic with which Noble gas?



17 p+
17 e- (7 valence)

17 p+
18 e- (gains one to fill the valence)
isoelectronic with Ar

III. Electron Configurations for Ions

A. Ions become **isoelectronic** with the Noble Gases.

1) **cation** examples (s^2p^6)

** *Cations lose electrons to become isoelectronic with the previous Noble Gas.* **

E3) Na has one valence e^-. It forms a 1+ ion, exposing its filled n = 2 shell.

	Na	→	Na^+	+	e^-
Complete config.:	$1s^2\,2s^2\,2p^6\,\underline{3s^1}$		$1s^2\,\underline{2s^2\,2p^6}\,3s^0$		the 3s electron
Valence config.:	$3s^1$		$2s^2 2p^6$		

The sodium ion has the electron configuration of neon (10 e-), but it still has a nucleus with 11 protons, keeping it sodium.

E4) Al has three valence e^-. It forms a 3+ ion, exposing its filled n = 2 shell.

	Al	→	Al^{3+}	+	$3e^-$
Complete config.:	$1s^2\,2s^2\,2p^6\,\underline{3s^2\,3p^1}$		$1s^2\,\underline{2s^2\,2p^6}\,3s^0\,3p^0$		3s and 3p electrons
Valence config.:	$3s^2\,3p^1$		$2s^2\,2p^6$		

The aluminum ion has the electron configuration of neon (10 e-), but it still has a nucleus with 13 protons, keeping it aluminum.

2) **anion** examples (s^2p^6)

*** Anions become isoelectronic with the Noble Gas at the end of their row. ***

E5) Cl has seven valence electrons. It forms a 1- ion and has the configuration of argon.

	Cl	+	e^-	→	Cl^-
Complete config.:	$1s^2 2s^2 2p^6 \underline{3s^2 3p^5}$		shared or stolen e^-		$1s^2 2s^2 2p^6 \underline{3s^2 3p^6}$
Valence config.:	$3s^2 3p^5$				$3s^2 3p^6$

The chloride ion has the electron configuration of argon (18 e-), but it still has a nucleus with 17 protons, keeping it an ion of Cl.

E6) S has six valence electrons. It forms a 2- ion and has the electron configuration of argon.

	S	+	2 e^-	→	S^{2-}
Complete config.:	$1s^2 2s^2 2p^6 \underline{3s^2 3p^4}$		shared or stolen e^-		$1s^2 2s^2 2p^6 \underline{3s^2 3p^6}$
Valence config.:	$3s^2 3p^4$				$3s^2 3p^6$

The sulfide ion has the electron configuration of argon (18 e-), but it still has a nucleus with 16 protons, keeping it an ion of S.

B. **Pseudo-Noble gas configuration** = $s^2p^6d^{10}$

1) found in ions such as Zn^{2+} and Ga^{3+}
2) a metal that already has transition (d) electrons forms a cation
3) this is an exception to normal s^2p^6 Noble Gas configurations
4) example **E7) :**

	Zn	→	Zn^{2+}	+	2 e^-
Complete:	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^2}\ 3d^{10}$		$1s^2\ 2s^2\ 2p^6\ \underline{3s^2\ 3p^6}\ 4s^0\ \underline{3d^{10}}$		the 4s electrons
Valence config.:	$4s^2$		$3s^2\ 3p^6\ 3d^{10}$		

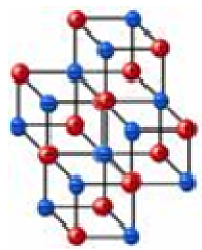
C. Ways to achieve a stable outer energy level

1) *transferring of electrons* = **ionic bond**
2) *sharing of electrons* = **covalent bond**

7.2 notes

IV. Formation of **Ionic Compounds** *(electron transfer)*

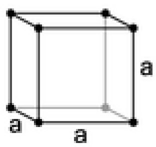
A. **ionic bond**—*a chemical attraction between positive ions (cation) and a negative ion (anion)*
   1) cation is the metal
   2) anion is the nonmetal
B. *total of positive charges = total of negative charges*
C. **Noble Gas configurations** of both ions are achieved
D. *electronegativity differences are large*
   1) the cation is not electronegative ("not greedy" or electropositive) and will let its electron(s) be taken
   2) the anion is more electronegative ("greedy") and will take the electron(s)
E. **formula unit**— the *lowest whole-number ratio of ions* in an ionic compound
   1) do not use the term "molecule" to describe an ionic compound
   2) ionic compounds occur in *repeating units in their crystals*
F. *crystalline solids*—a structure containing *3-D repeating patterns* of formula units



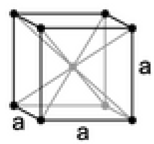
1) crystals are arranged in a *positive-negative alternating setup*:

Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-
Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+
Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-
Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+
Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-
Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+	Cl-	Na+

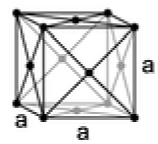
2) examples of crystal shapes: SC, BCC, FCC, HCP



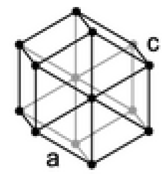
simple cubic



body-centered cubic (BCC)



face-centered cubic (FCC)



hexagonal close-packed (HCP)

3) setup is called a **crystal lattice** which is hard and brittle
4) **lattice energy**
   a) the strength needed to separate 1 mole of ions from the crystal lattice
   b) smaller ions have greater lattice energy
   c) common unit = kJ/mol

G. **salt**—*any ionic compound,* not just sodium chloride

V. Cations and Anions in ionic compounds
   A. formation of **cations**
      1) the **octet rule** applies
      2) example

E8) Francium has one valence electron (87 total e^-). It is easier for it to lose the valence electron than it is to gain seven more to complete the "shell." It forms a 1+ ion, exposing a previously filled shell.

$$Fr \rightarrow Fr^+ + e^-$$

The francium ion is isoelectronic with radon (86 e^-), but it still has a nucleus with 87 protons, keeping it an ion of Fr instead of changing it into an atom of Rn.

   B. formation of **anions**
      1) the **octet rule** applies
      2) *common ending is —IDE* (chloride, sulfide, oxide...)
      3) example

E9) Selenium has six valence electrons (34 total e^-). It only needs two more to achieve a complete octet. It forms a 2- ion and has the configuration of krypton (36 e^-).

$$Se + 2\,e^- \rightarrow Se^{2-}$$

The selenide ion is isoelectronic with krypton (36 e-), but it still has a nucleus with 34 protons, keeping it an ion of Se instead of changing it into an atom of Kr.

VI. Electron Dot Diagrams for Ionic Compounds
   A. make sure the compound is composed of metals and nonmetals
   B. draw the dot diagrams of the cations and anions, using different colors if possible or open circles and closed circles for electrons from the different ions

(Remember all electrons are alike, but show which ones came from which ions in the transfer.)

C. all ions should be stable
   1) all cation dot diagrams show no dots
   2) all anion dot diagrams show a complete octet
   3) all charges will balance out to zero

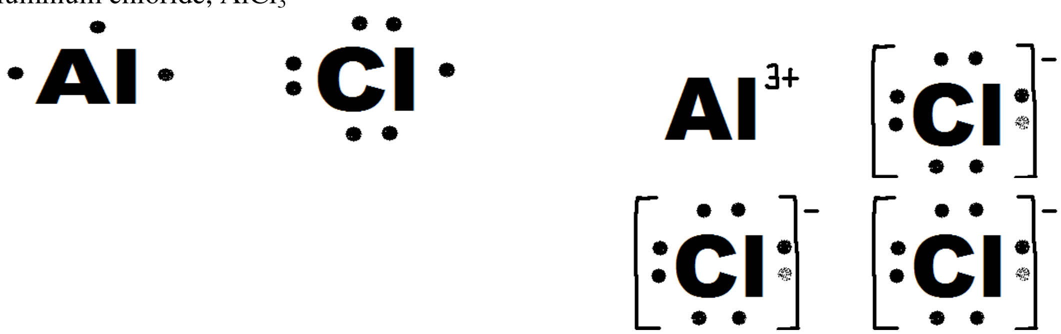
D. examples

E10) sodium chloride, NaCl

sodium atom chlorine atom → → sodium ion chloride ion

CATION CHARGE TOTAL = 1+
ANION CHARGE TOTAL = 1-
NET CHARGE = 0

E11) aluminum chloride, $AlCl_3$



aluminum atom chlorine atom → → aluminum ion chloride ions

(three Cl^- are needed to balance the charges)

CATION CHARGE TOTAL = 3+
ANION CHARGE TOTAL = 3 x (1-) = 3-
NET CHARGE = 0

VII. Comparison of properties of ionic and molecular compounds

	Molecular compounds	**Ionic compounds**
Particles	Molecules	Formula units (salts) made of ions
Elements	Nonmetals	Metals & nonmetals
Conductivity	Low "nonelectrolytes"	High (when molten or in aqueous solution) "electrolytes"
State at room temp.	Solid, liquid, or gas	Solid
Type of Bond	Polar or nonpolar covalent	Ionic

7.3 notes

VIII. Compounds in General

A. **compound** ("cmpd.")—*a substance formed from more than one element*

B. **molecule**—*a group of atoms with no net charge*

C. two general types of compounds

1) **molecular compound**—*composed of atoms*
   - usually liquids or gases at room temp.
   - usually have a low melting point (m.p.) and boiling point (b.p.)
   - *composed of nonmetals…* like CO_2

2) **ionic compounds**—*composed of positive/negative ions*
   - usually crystalline solids at room temp.
   - usually have a high melting point (m.p.) and boiling point (b.p.)
   - *composed of metals and nonmetals…* like NaCl

D. **chemical formula**

1) symbols representing the composition of the smallest unit of a substance

2) *shows which elements are present and how many there are*

   H_2SO_4 = hydrogen, sulfur, oxygen = 7 atoms total

3) **molecular formula**—symbols representing the composition of a *molecular compound*

4) **formula unit**— the lowest whole-number ratio of ions in an *ionic compound*

   a. It is improper to use the term "molecule" to describe an ionic compound.

   b. Ionic compounds occur in repeating units in crystal lattices

5) number codes

   a. **subscript**—a number written slightly *below* the symbol Br_2

   b. **superscript**—a number written slightly *above* the symbol Sr^{2+}

(In chem., a superscript is not called an exponent. Nothing is being multiplied.)

IX. Ionic Compounds

A. *metals form cations (+)* *nonmetals form anions (--)*

B. charge is also called **oxidation number**

C. multiple charges (transition metals and others) – *if an atom forms more than one charge, you must say which one it is:*

1) Stock system uses number clues – this is the most common way to do it

2) Classical (Latin root) system

LOWER CHARGED ION: suffix "–OUS" Cu^{+} *= cuprous ion*
HIGHER CHARGED ION: suffix "–IC" Cu^{2+} *= cupric ion*

IMPORTANT REFERENCE TABLE FOR IONS WITH MULTIPLE CHARGES:

ION FORMULA	STOCK NAME	(CLASSICAL NAME)
Cu^{+}	**copper(I) ion**	cuprous ion
Cu^{2+}	**copper(II) ion**	cupric ion
Fe^{2+}	**iron(II) ion**	ferrous ion
Fe^{3+}	**iron(III) ion**	ferric ion
Pb^{2+}	**lead(II) ion**	plumbous ion
Pb^{4+}	**lead(IV) ion**	plumbic ion
Sn^{2+}	**tin(II) ion**	stannous ion
Sn^{4+}	**tin(IV) ion**	stannic ion

MORE →

Cr^{2+}	**chromium(II) ion**	chromous ion
Cr^{3+}	**chromium(III) ion**	chromic ion
Mn^{2+}	**manganese(II) ion**	manganous ion
Mn^{3+}	**manganese(III) ion**	manganic ion
Co^{2+}	**cobalt(II) ion**	cobaltous ion
Co^{3+}	**cobalt(III) ion**	cobaltic ion
Hg^{2+}	mercury(I) ion	mercurous ion
$Hg_2{}^{2+}$	mercury(II) ion	mercuric ion

3) **one-charge transition metal ions: Ag^{+}, Cd^{2+}, Zn^{2+}**

X. Binary compounds: binary ionic (this chapter) and binary molecular (next chapter)

A. **<u>binary compound</u>**—*composed of two elements*

B. **<u>binary ionic compound</u>**—("BI")—*metal cation / nonmetal anion combination*

1) compound composed of **<u>monatomic</u>** (*one symbol*) ions
2) crisscross formula method is used
3) name the cation first, then the anion (-*IDE* ending)
4) use Roman numerals if/when needed for the cation
5) examples:

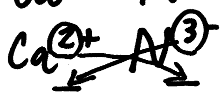
$CaBr_2$ = calcium bromide   KI = potassium iodide $FeCl_3$ = iron(III) chloride

C. *"crisscross" formula method for ionic compounds (charge balancing)*

1) *write the symbols of the two ions next to each other*
2) *write the charges as superscripts*
3) *balance the formula by crisscrossing the numbers   (net charge = 0)*
4) *the numbers are now subscripts, telling you how many of each symbol is in the formula*
5) *if the charges are the same, they cancel out*
6) *if the charges are multiples of each other, reduce them*

E12: calcium nitride
Ca^{2+} and N^{3-}

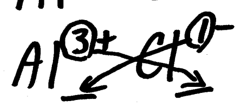
Ca^{2+} N^{3-}



Ca_3N_2

E13: aluminum chloride
Al^{3+} and Cl^{1-}

Al^{3+} Cl^{1-}



Al_1Cl_3

$AlCl_3$

XI. Polyatomic ions

A. **<u>polyatomic ion</u>**—*a group of charged atoms*

B. most end in –*ATE* or –*ITE*

C. To avoid confusion, you may keep the parentheses around the polyatomic ion if there is only one polyatomic ion in the formula, but be aware that textbooks do not do this.

Na(OH) = NaOH

POLYATOMIC IONS
Students must memorize these 20 ions.

<u>1^+ CHARGE:</u>
ammonium $(NH_4)^+$

<u>1^- CHARGE:</u>
acetate $(C_2H_3O_2)^-$ or $(CH_3COO)^-$
chlorate $(ClO_3)^-$
chlorite $(ClO_2)^-$
cyanide $(CN)^-$
hydrogen carbonate or bicarbonate $(HCO_3)^-$
hydroxide $(OH)^-$
hypochlorite $(ClO)^-$
nitrate $(NO_3)^-$
nitrite $(NO_2)^-$
perchlorate $(ClO_4)^-$
permanganate $(MnO_4)^-$

<u>2^- CHARGE:</u>
carbonate $(CO_3)^{2-}$
chromate $(CrO_4)^{2-}$
dichromate $(Cr_2O_7)^{2-}$
silicate $(SiO_3)^{2-}$
sulfate $(SO_4)^{2-}$
sulfite $(SO_3)^{2-}$

<u>3^- CHARGE:</u>
phosphate $(PO_4)^{3-}$
phosphite $(PO_3)^{3-}$

XII. Ternary Compounds

A. **<u>ternary compound</u>**—compound made up of *three different elements*

B. **<u>ternary ionic compound</u>**—("TI")— *metal cation / nonmetal anion combination of three symbols, involving polyatomic ions*

1) *crisscross* formula method is used
- *write the symbols of the two ions next to each other*
- *be careful to keep the parentheses around the polyatomic ion*
- *write the charges as superscripts*
- *balance the formula by crisscrossing the numbers (net charge = 0)*
- *the numbers are now subscripts, telling you how many of each symbol is in the formula*
- *if the charges are the same, they cancel out*
- *if the charges are multiples of each other, reduce them*

E14: strontium nitrate
Sr^{2+} and $(NO_3)^{1-}$

$Sr^{2+}(NO_3)^{1-}$
$Sr^{2+}(NO_3)^{1-}$
$Sr_1(NO_3)_2$
$Sr(NO_3)_2$

E15: aluminum sulfate
Al^{3+} and $(SO_4)^{2-}$

$Al^{3+}(SO_4)^{2-}$
$Al^{3+}(SO_4)^{2-}$
$Al_2(SO_4)_3$

2) name the cation first, then the anion

3) use Roman numerals if/when needed for the cation

C. examples

E15: sodium sulfate

Na_2SO_4

E16: potassium permanganate

$KMnO_4$

E17: iron(III) hydroxide

$Fe(OH)_3$

(from Ch. 10)

XIII. Hydrates

A. **water of hydration**—*water molecules chemically integrated into a crystalline structure*

B. **hydrate**—*a compound with water in its structure*

1) *general formula:* *[compound]* $\cdot$ H_2O

2) *naming: [compound name] (prefix)hydrate*

MONO-,	DI-,	TRI-,	TETRA-,	PENTA-,	HEXA-,	HEPTA-,	OCTA-,	NONA-,	DECA-
1	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**

$CuSO_4 \cdot 5H_2O$ copper(II) sulfate pentahydrate

C. **anhydrous**—*a hydrate without its water of hydration*

$Na_2SO_4 \cdot 2H_2O \rightarrow Na_2SO_4 + 2H_2O$

sodium sulfate dihydrate → anhydrous sodium sulfate + water

D. **efflorescence**—the release of water by a hydrate (heating not needed)

E. **hygroscopic**—*removing water from the atmosphere*

F. **deliquescence**

1) absorbing excess water from the atmosphere to form a liquid substance

2) this is an extreme hygroscopic condition

G. **desiccant**

1) a *drying agent* which is hygroscopic

2) examples: Damp Rid; packets of silica powder in shoe boxes

7.4 notes

XIV. Metals

A. electrons and electricity: metallic bonds

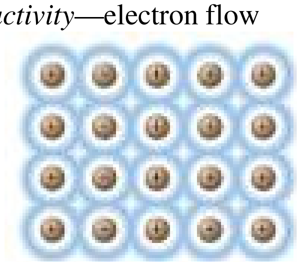
1) **metallic bonds**—*sharing of delocalized electrons between metal cations, allowing current to flow*

a) **delocalized electrons**— *shared electrons* not in one place; mobile

b) delocalized electrons are free to move throughout the piece of metal

c) **electron sea model**– metal cations are surrounded by "community" electrons

2) *conductivity*—electron flow



delocalized (shared) electrons

B. examples of crystal patterns in metals
1) BCC: Li, Na, K, Rb, Cs, Ba, Ra, V, Cr, Mn, Fe, Nb, Mo, Ta, W, Pd, Pt, Eu
2) FCC: Ca, Sr, Ni, Cu, Al, Au, Ag, Ge, Pb, Rn, Ir, Ce, Ac, Th, Yb
3) HCP: Be, Mg, Sc, Ti, Y, Zr, Zn, Lu, Hf, Co, Tc, Ru, Re, Os, Cd, Tl, La, Pr, Nd, Gd, Tb, Dy, Er, Tm, Am, Cm

C. characteristics of metals
1) *ductile—able to be pulled into thin wires*
2) *malleable—able to be hammered into sheets*
3) shiny and lustrous
4) conductors of electricity
5) durable—strong

D. alloys
1) **alloy**—*physical blending of metals*
2) examples
   a) brass = Cu, Zn
   b) bronze = Cu, Zn, Sn
   c) stainless steel = Fe, Cr, Ni
   d) 10K gold = Au, Ag, Cu