Compounds

Word	Definition
Binary compound	A compound that consists of two elements
Dipole attractions	The attraction of the δ- end of one polar molecule to the δ + end of another polar molecule.
Dipole moment	An arrow along the line of symmetry in a polar molecule that shows the net direction that electrons are being pulled towards the partially negative end of the molecule.
Electronegativity difference	The difference in electronegativity between two elements in a bond.
Electrical conductivity (metals)	The ability of a substance to allow electrons to pass from atom to atom through itself from a source or electricity to an electrical ground.
Electrolyte	A solution containing dissolved ions that can conduct electricity through it.
Empirical formula	The simplest whole-number mole ratio of elements in a compound, used to write the formulas of ionic compounds.
Formula mass	The sum of the atomic masses of an element or compound, measured in grams/mole. Reported to the nearest tenth of a gram per mole.
Hydrogen bonds	The strong attraction of the H ($\delta+$) end of one polar molecule to the N, O or F ($\delta-$) end of another polar molecule. The two ends form temporary covalent bonds.
Intermolecular Attractive Force	The forces that hold molecules together in the solid and liquid phase. These are the forces that must be overcome to melt or boil a substance. Also called "van der Waal's forces".
Ionic compound	Compounds consisting of a metal and a nonmetal that are ionically bonded in a whole-number ratio.
London dispersion force	The weak attractive force caused by temporary dipoles in nonpolar molecules.
Metallic bond	A bond formed between metal atoms as they collectively share their conducting electrons evenly between metal kernals.
Molecular formula	The actual number of nonmetal atoms in a molecule, a whole-number multiple of the empirical formula.
Molecule	A particle made of nonmetal atoms covalently bonded together to form a distinct particle.
Network solid	A crystal lattice formed from covalently bonded nonmetal atoms with no distinct molecules.
Nonpolar molecule	A molecule with symmetrical electron distribution resulting in any polar bonds canceling each other out to yield no partially charged ends.
Percent composition	The formula mass of an element divided by the formula mass of the compound that contains the element multiplied by 100.
Polar molecule	A molecule with asymmetrical electron distribution resulting in partially charged ends.
Polyatomic Ion	An ion formed by atoms bonding together in such a way that a net charge is formed.
Ternary compound	A compound that consists of three or more elements, usually containing a polyatomic ion.

1) Types of Compounds (HW: p. 24, 25)

Essential Question: How are the properties of a substance dependent on its composition and structure?

Compound: substance formed by the chemical bonding of atoms. The type of compound is determined by the type of bonding involved.

Ionic Compounds: formed by ionic bonding. These are found in crystal form of alternating + and – charged ions.

- Since a full + and – charge is involved, the attraction between ions is strong. High melting and boiling points, tend not to evaporate.
- When dissolved in water or melted, the ions separate and allow for electrical conduction. The solution so formed is called an **electrolyte**.
- Attractions between ions is called the **ionic attraction**. Duh. ☺
- Melting or dissolving ionic compounds breaks the ionic bond. Ionic reactions where the bonds have already been broken are extremely fast.

Molecular Compounds: formed by covalent bonding, either polar or nonpolar. These form individual particles called molecules which can attract to each other to form the solid or liquid phase. Molecules can have oppositely charged ends, which allow them to attract to one another. These are called **intermolecular attractive forces**, and are weaker than ionic attractions. Therefore, molecular compounds are more easily melted and boiled, so their melting and boiling points are low compared to ionic compounds. They also tend to evaporate more quickly. Their solids are soft (think wax or water ice, compared to something like steel!)

HOWEVER

Dissolving in water and melting **does not break the covalent bond**. There are no ions to carry electrical current, so molecules **do not conduct electricity**. Acids are the exception to this, but for now we will turn a blind eye to them…

Network Solids: also formed by covalent bonding (usually nonpolar), but it doesn't form separate molecules. Instead, it forms one single crystal made of nonmetal atoms connected with a continuous network of covalent bonds with no areas of weakness that can break apart. Molecules can be separated from each other, but network solids have no such weakness. They are among the hardest substances known to science. They also occupy the top of the Mohs hardness scale (talc at 1, diamond at 10). Quartz (SiO_2) has a hardness of 7, corundum (Al_2O_3, rubies and sapphires) have a hardness of 9, and diamond (pure crystalline carbon) has a hardness of 10, the top of the scale. Being made only of nonmetals (or metalloids), network solids are nonconductors of electricity and poor conductors of heat. They are also quite brittle.

Metallic Compounds: Metals don't technically form compounds with other metals. Metal atoms share electrons by losing them. In a way it is like a game of "hot potato" with their valence electrons**.** If one could see metal atoms in action they would appear as positively charged kernels surrounded by a sea of moving (conducting) electrons. These electrons are evenly distributed throughout the metal and yet able to move freely. This allows metals to conduct electricity in <u>all</u> phases.

Physical Properties of Types of Compounds

Type of Compound	Attractive Force Strength	Melting Point/Boiling Point	Vapor Pressure	Electrical Conductivity	Malleable/ Brittle	Type of Bond	Examples
Ionic	Strong	High	Low	Only in (l) and (aq)	Brittle	Ionic	NaCl,. $ZnSO_4$
Molecular	Weak	Low	High	Never	Brittle	Covalent	H_2O, NH_3, $C_6H_{12}O_6$
Network	Very Strong	Very High	Low	Never	Brittle	Covalent	C (diamond), SiO_2 (quartz)
Metallic	Varies	Varies	Varies	Always	Malleable	Metallic	Zn, Cu, Ca, Mg, Na, Fe

MOLECULAR POLARITY

How can you tell if a molecule is polar or nonpolar?

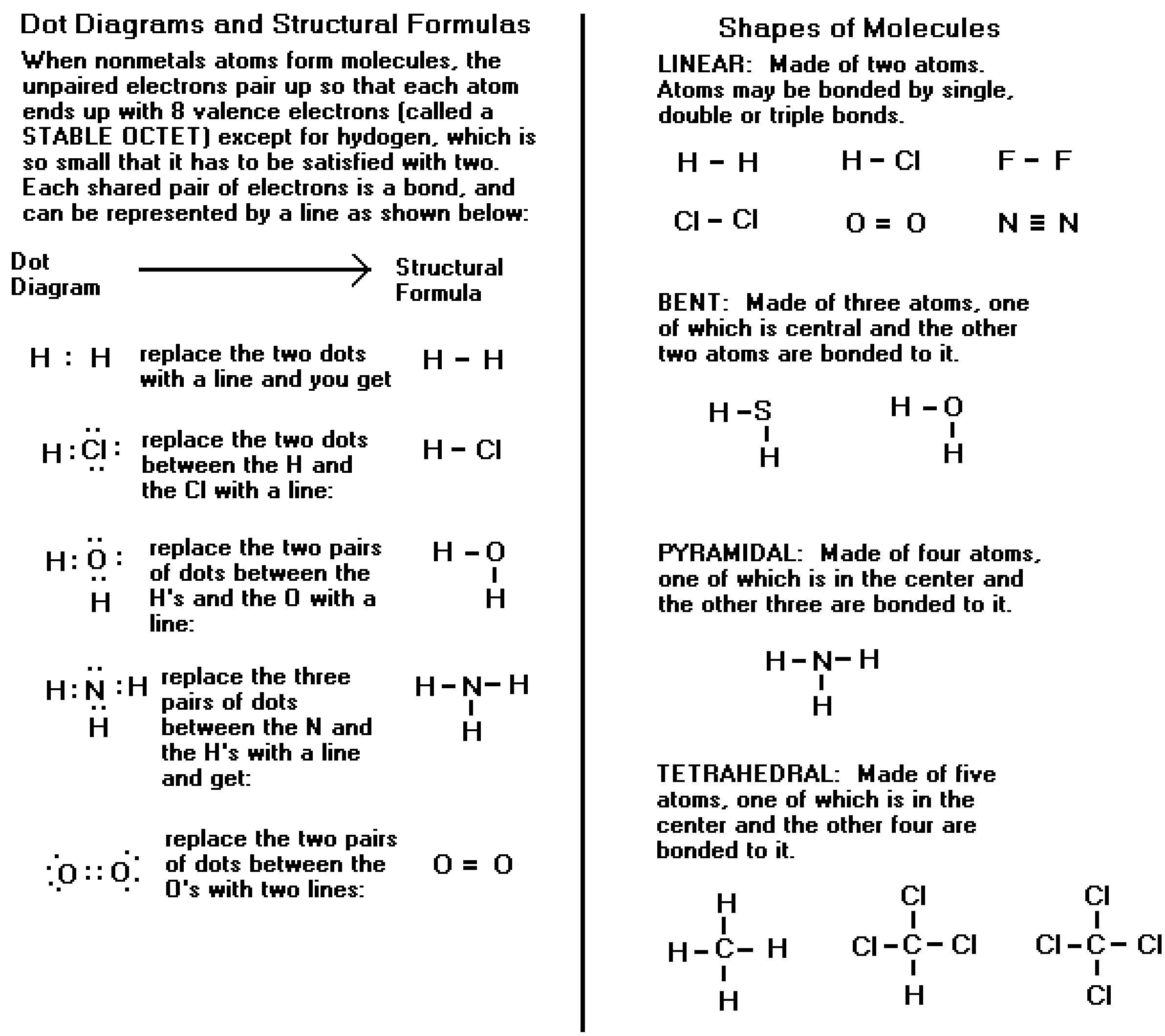
The polarity of the molecule is different than the polarity of the bond. A bond is polar if the END between the bonding atoms is 0.5 or higher. A molecule can have polar bonds and still be a nonpolar molecule. **The polarity of the molecule is determined the polarity and positioning of all the bonds in the molecule**.

Why is it important to know the polarity of a molecule? The polarity can tell you many things:

1) How high or low the melting and boiling point of the substance is
2) How easily the liquid form of the substance evaporates
3) Whether the substance will dissolve in water, or in some other solvent

Molecular polarity causes intermolecular attractive forces, which is what allows certain adhesives (tapes, glues) to hold things together. Attractive forces also hold the two twisted strands of RNA together to form DNA. Attractive forces allow insects and geckos to walk up walls, even glass walls. Think of attractive forces as molecular Velcro…you can attach and detach the molecules from each other without doing any damage to the molecules themselves. Some molecules have stronger Velcro than others. How do you know which ones these are? Read on!

<u>First thing you need to know is what is the shape of the molecule? Its shape determines its properties.</u>

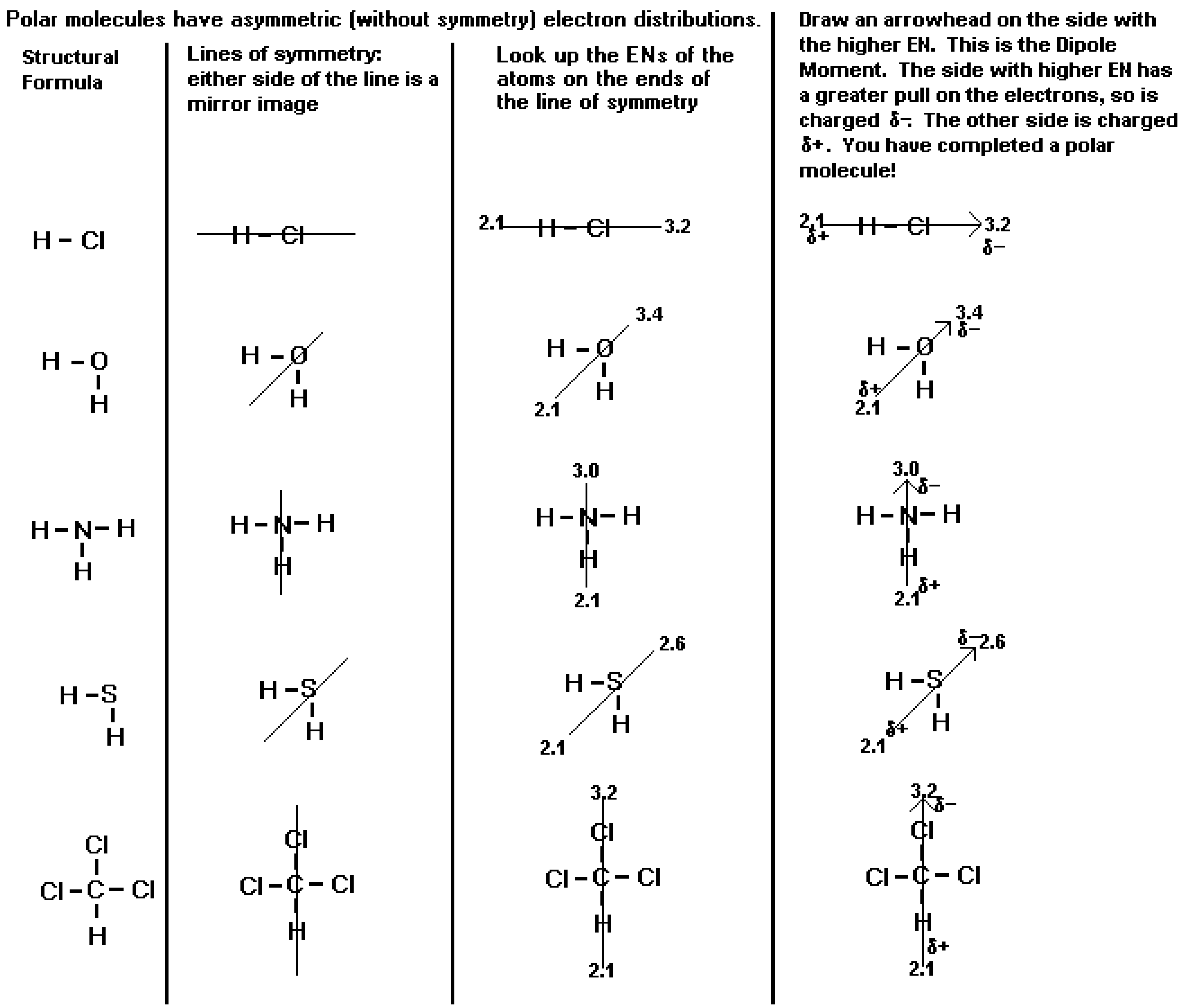


Determining Molecular Polarity

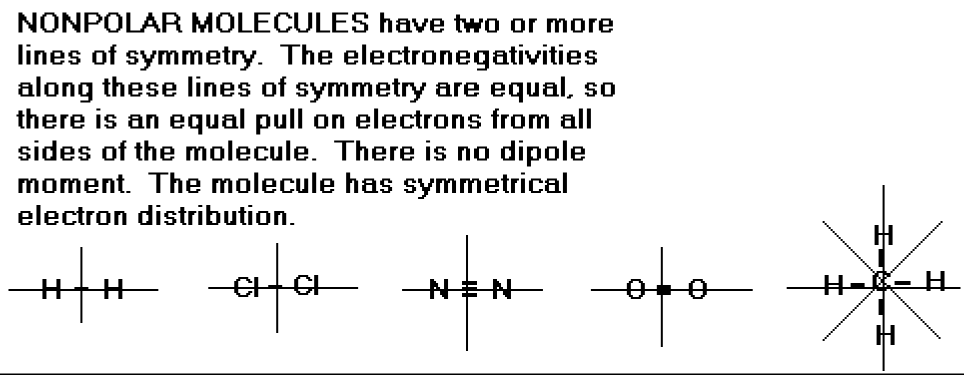
POLAR MOLECULES: Nonmetals in a molecule share electrons, sometimes equally (nonpolar bond) or unequally (polar covalent bond) when forming molecules. If the molecule does not have a symmetrical shape (asymmetrical), then there is a greater concentration of electrons on one side of the molecule compared to the other side. This makes one side charged partially negative and the other side partially positive. The oppositely charged ends of these molecules are "poles", making the molecule polar. Polar molecules can attract each other, δ+ end of one molecule attracting to the δ- end of the other molecule. These are the intermolecular attractive forces mentioned on page 5!

If a molecule has polar bonds in an asymmetrical arrangement, **the molecule as a whole will have a net polarity towards the more electronegative atom in the molecule**. The net direction of electron pull towards the more electronegative atom can be diagrammed by determining the sole axis of symmetry and **drawing an arrow towards the more electronegative atom**.

POLAR MOLECULE. One line of symmetry. The line of symmetry forms the Dipole Moment. Evaluate the electronegativities of the atoms on ether end of the line of symmetry. The more electronegative end will be charged partially negative, the other end partially positive. This is weaker than an ionic full charge. The line of symmetry has an arrowhead put on the side with more electronegativity, this is called the dipole moment, or direction in which the electrons are pulled.

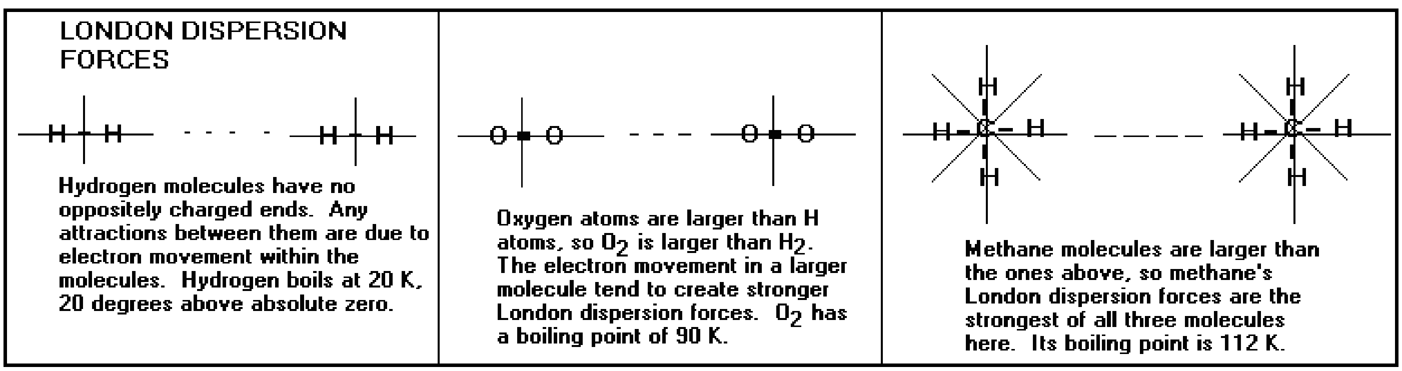


NONPOLAR MOLECULES). If the molecule has a symmetrical shape, then the electrons are distributed evenly through the molecule, and the whole molecule is nonpolar (even if it contains polar bonds). Nonpolar molecules have equal pull of electrons on all sides of the molecule, so no side develops a pole. Since the molecule lacks oppositely charged ends, any attractive forces will be extremely weak. Small nonpolar molecules are usually found in the gaseous state at room temperature (CH_4, methane, also known as natural gas; C_3H_8, propane, also known as bottled gas and C_4H_8, butane, also known as lighter fuel), larger molecules can be liquids (C_8H_{18}, octane, also known as gasoline and C_6H_6, benzene, which is a liquid capable of dissolving plastic) and huge nonpolar molecules can be found in the solid phase ($C_6H_4Cl_2$, paradichlorobenzene is one molecule that mothballs have been made of).

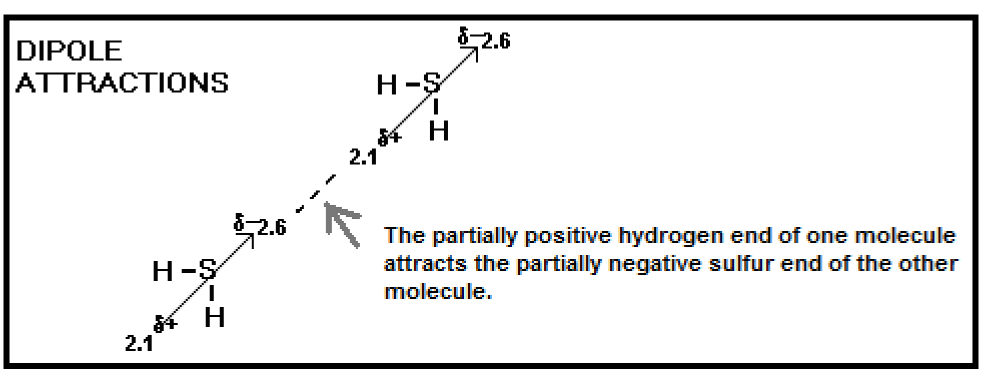


Attractive Forces – what holds molecules together in the solid and liquid phase

Nonpolar molecules form LONDON DISPERSION FORCE attractions. Since there are no permanent positive or negative ends, these attractions are extremely weak. The attractions are a combination of temporary poles due to electron movement around the molecule or, in the case of huge molecules, they actually get tangled up with each other like sticky strands of spaghetti or yarn. London dispersion forces generally get stronger as the size of the molecule increases.



Polar molecules form DIPOLE attraction: the simple attraction of the oppositely charged ends of two molecules. The partially positive end of one molecule attracts to the partially negative end of a different molecule. This attraction allows these substances to exist as solids and liquids at higher temperatures than are possible for nonpolar molecules of equivalent size.



Some polar molecules form HYDROGEN BONDS between them. Most notable among the molecules with this special type of attractive force is water. In water, the hydrogen atom from the partially positive end of one water molecule is attracted to the oxygen atom from the partially negative end of another water molecule. Yes, there is the electrostatic attraction, which is very strong…the END between H and O is 1.3, which is strongly polar…but there is more going on than just electrostatic attraction. The hydrogen of one water molecule actually forms temporary covalent bonds with the oxygen from the other water molecule! This is because the oxygen is so electronegative (and has such a small radius) that it actually delocalizes the hydrogen's shared electron with the oxygen atom it is already bonded to. This causes the hydrogen atom to be simultaneously bonded to both the oxygen it shares a molecule with and the oxygen atom in the neighboring water molecule. The bond between the hydrogen and its own oxygen atom is much stronger, but the weak, temporary covalent bond that it forms with the other molecule's oxygen atom makes hydrogen bonds responsible for most of water's amazing properties. These properties include:

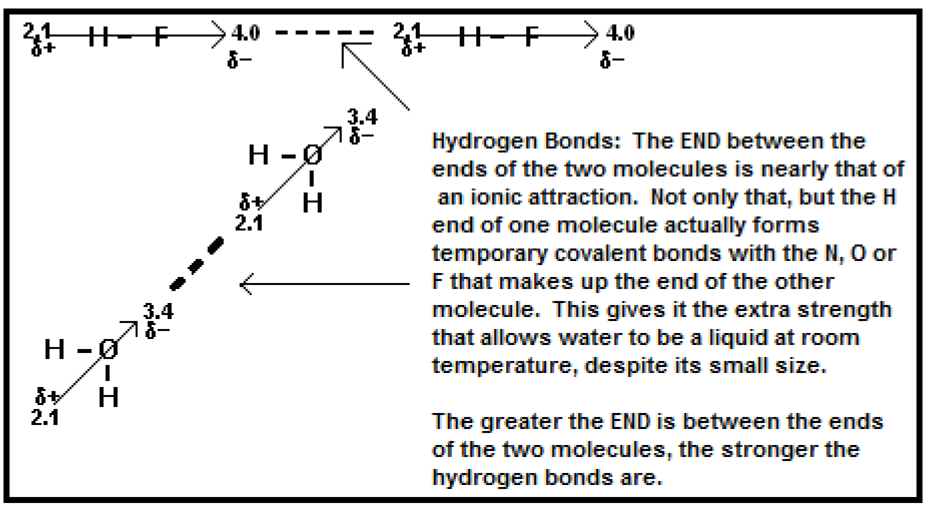
- An extremely high melting point, boiling point, heat of fusion and heat of vaporization for a molecule its size

- The ability of water to form a "skin" at the surface called surface tension that allows insects to walk across it and for a meniscus to form

- The ability of water molecules to climb up a narrow space, called "capillary action". This is how water is transported from the roots of a plant to the leaves. Water molecules that climb up the xylem tubes of plants pull the water molecules behind them up as well

- Water molecules can be deflected by an electric field

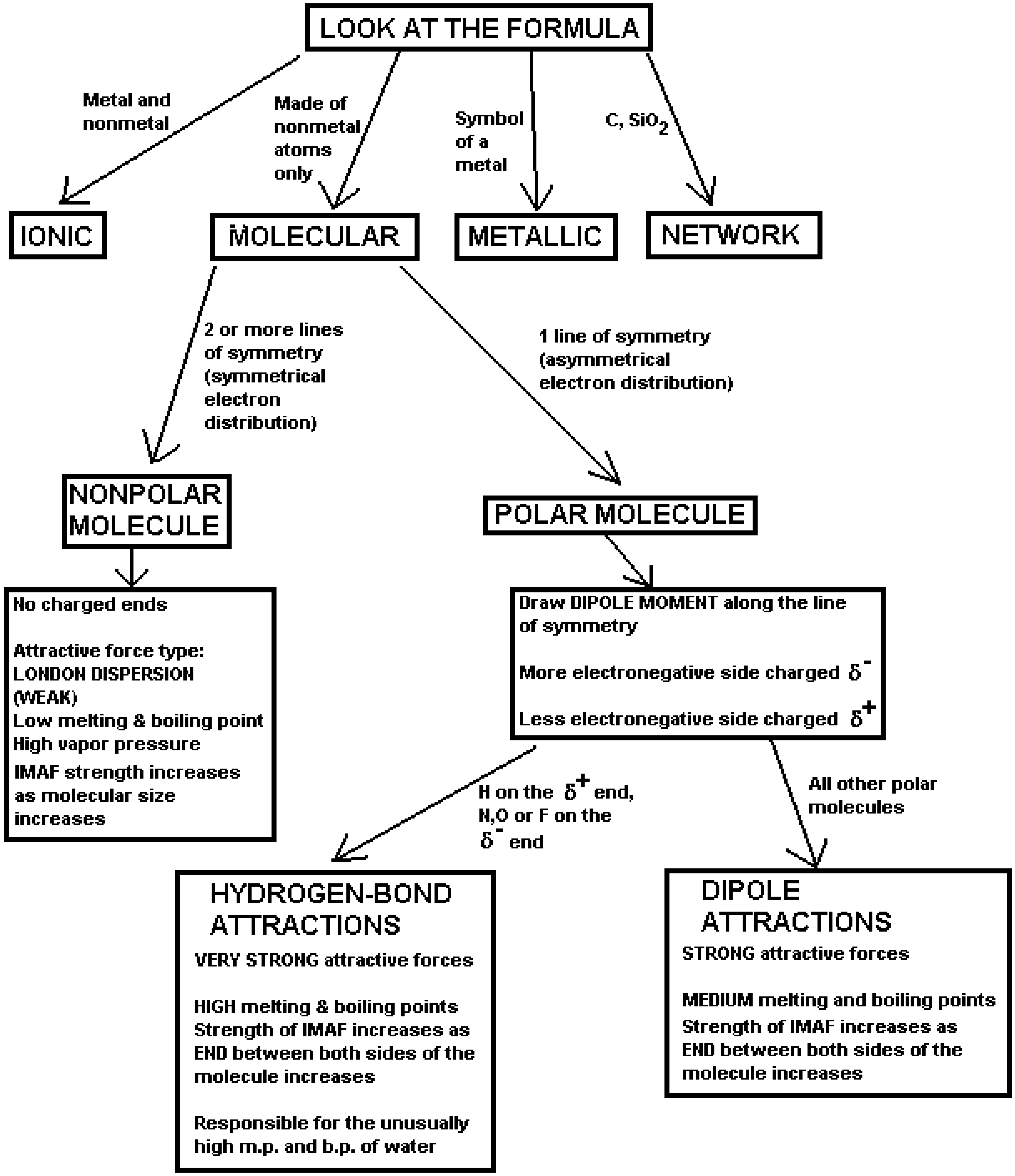
Other atoms that hydrogen can form a hydrogen bond with are the other nonmetal atoms that also have a high electronegativity and very small atomic radius; nitrogen (3.0) and fluorine (4.0). Even though chlorine has a higher electronegativity (3.2) than nitrogen does, chlorine does not form hydrogen bonds. This is because chlorine has one more principal energy level than N, O or F and that gets in the way of the temporary covalent bond forming.



Attractive Force Type, Strength And Resulting Molecular Properties

Lines Of Symmetry	Molecule Polarity	Attractive Force Type	Attractive Force Strength	Melting & Boiling Points	Vapor Pressure (ability to evaporate)
0 or 1	Polar	(if H end of one molecule is attracted to N,O or F end of the other molecule) HYDROGEN BOND	Strong	High	Poor
0 or 1	Polar	(any other polar molecule) DIPOLE	Moderate	Moderate	Fair
2 or more	nonpolar	LONDON DISPERSION FORCES	Weak	Low	Good

SUMMING IT ALL UP: A FLOWCHART



So, what should you be able to do now?

1) Identify whether a compound is molecular, ionic, metallic or network based on its properties
2) Draw dot diagrams of simple molecules
3) Draw structural formulas of simple molecules
4) Determine the shape of simple molecules
5) Determine if simple molecules are polar or nonpolar
6) If polar, draw the dipole moment and identify the partially charged ends
7) Determine the attractive force type that attracts specific simple molecules to each other.

2) Naming and Writing Binary and Ternary Formulas (HW: p. 26-29)

Essential Question: Just how do you write chemical formulas, anyway?

Binary Compounds - Compounds made of TWO elements chemically combined in a simple, whole-number ratio.

The ratio of the elements in the compounds is dependent on the oxidation numbers (charges) of the elements that are combining. Atoms combine in a ratio such that the oxidation numbers add up to ZERO.

In a chemical formula, the CATION is written first, and the ANION is written second. The number 1 is not written in a formula. NaCl has 1 Na and 1 Cl, Li_2O has 2 Li's and 1 O.

WRITING BINARY FORMULAS: X will designate a metal ion and Y will designate a nonmetal ion.

Ions to be Bonded	How to determine formula (to have the + and the – charges add up to ZERO)	Formula	Example
X^{+1} and Y^{-1}	+1 and –1 add up to zero already.	XY	K^{+1} and F^{-1} → **KF**
X^{+2} and Y^{-2}	+2 and –2 add up to zero already.	XY	Zn^{+2} and S^{-2} → **ZnS**
X^{+3} and Y^{-3}	+3 and –3 add up to zero already.	XY	Al^{+3} and N^{-3} → **AlN**
X^{+2} and Y^{-1}	It takes two –1 charges to cancel out the +2 charge.	XY_2	Ca^{+2} and Cl^{-1} → **$CaCl_2$**
X^{+3} and Y^{-1}	It takes three –1 charges to cancel out the +3 charge	XY_3	Fe^{+3} and Br^{-1} → **$FeBr_3$**
X^{+1} and Y^{-2}	It takes two +1 charges to cancel out the -2 charge.	X_2Y	Li^{+1} and O^{-2} → **Li_2O**
X^{+1} and Y^{-3}	It takes three +1 charges to cancel out the -3 charge	X_3Y	Na^{+1} and P^{-3} → **Na_3P**
X^{+2} and Y^{-3}	Common denominator: 2 X 3 = 6. It takes three +2's to make +6 and two –3's to make –6	X_3Y_2	Cu^{+2} and N^{-3} → **Cu_3N_2**
X^{+3} and Y^{-2}	Common denominator: 2 X 3 = 6. It takes three -2's to make -6 and two +3's to make +6	X_2Y_3	Cr^{+3} and S^{-2} → **Cr_2S_3**
X^{+4} and Y^{-2}	It takes two –2 charges to cancel out one +4 charge.	XY_2	Pb^{+4} and O^{-2} → **PbO_2**

NAMING BINARY FORMULAS

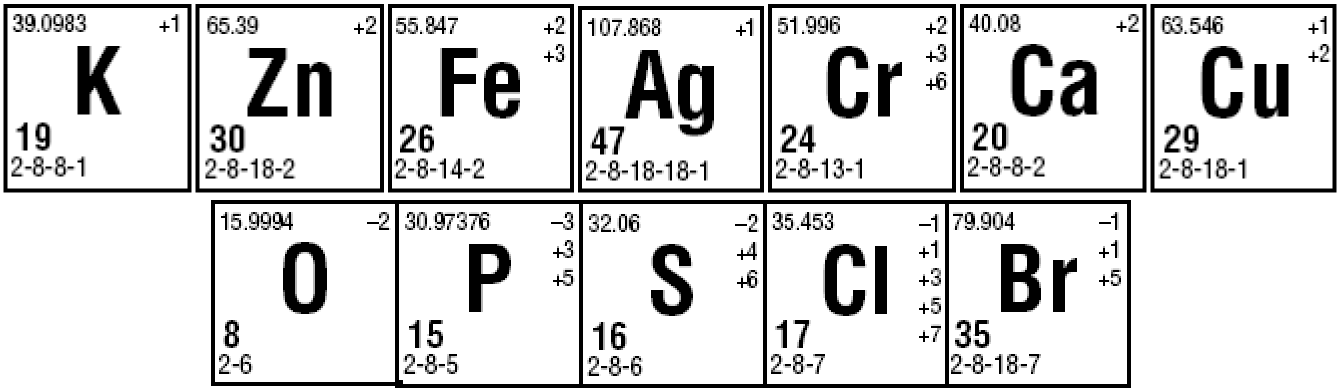
Use the name of the positive ion (if it has more than one listed charge, use the Stock system (where a Roman numeral is used after the ion name to indicate the charge), use the periodic table to determine if necessary). Use the name of the negative ion. Only the first listed charge is used when nonmetals form ionic compounds.

Put them together, metal first, nonmetal second. If you have lead (IV) (Pb^{+4})and oxide (O^{-2}), the name is lead (IV) oxide. (PbO_2)

Roman Numerals:

If the formula contains a metal that has more than one charge listed and the charge of the metal ion in that formula is…	…then use this Roman numeral!
+1	(I) one
+2	(II) two
+3	(III) three
+4	(IV) one before five = four
+5	(V) five
+6	(VI) one after five = six
+7	(VII) two after five = seven

Refer to the ion charges here when looking at the sample problems below:



Formula	How to determine name	Name
KCl	K only has one charge, so K^{+1} is called **potassium**. Cl^{-1} is called **chloride**	Potassium chloride
ZnO	Zn only has one charge, so Zn^{+2} is called **zinc**. O^{-2} is called **oxide**.	Zinc oxide
$FeBr_2$	Fe has two charges listed, so use Br to determine Fe's charge. Each Br is –1, so two of them would be –2. To add up to zero, Fe has to be charged +2, called **iron (II)**. Br^{-1} is called **bromide**.	Iron (II) bromide
$FeBr_3$	Fe has two charges listed, so use Br to determine Fe's charge. Each Br is –1, so three of them would be –3. To add up to zero, Fe has to be charged +3, called **iron (III)**. Br^{-1} is called **bromide**.	Iron (III) bromide
Ag_2S	Ag only has one charge, so Ag^{+1} is called **silver**. S^{-2} is called **sulfide**.	Silver sulfide
Cr_2O_3	Cr has three charges listed, so use O to determine Cr's charge. Each O is –2, so three of them would be –6. To add up to zero, Cr has to be charged +6 total. Since there are two Cr's, each one is charged +3, called **chromium (III)**. O^{-2} is called **oxide**.	Chromium (III) oxide
CrO_3	Cr has three charges listed, so use O to determine Cr's charge. Each O is –2, so three of them would be –6. To add up to zero, Cr has to be charged +6, called **chromium (VI)**. O^{-2} is called **oxide**.	Chromium (VI) oxide
Ca_3P_2	Calcium only has one charge, so Ca^{+2} is called **calcium**. P^{-3} is called **phosphide**.	Calcium phosphide
CuS	Cu has two charges listed, so use S to determine Cu's charge. S is –2, so to add up to zero, Cu has to be charged +2, called **copper (II)**. S^{-2} is called **sulfide**	Copper (II) sulfide
Cu_2S	Cu has two charges listed, so use S to determine Cu's charge. S is –2, so to add up to zero, Cu has to be charged +2 total. Since there are two copper ions, each one must be charged +1, called **copper (I)**. S^{-2} is called **sulfide**	Copper (I) sulfide

WRITING FORMULAS GIVEN THE NAME

Name	How to write the formula (verify charges using periodic table)	Formula
Potassium sulfide	Potassium is K^{+1}, sulfide is S^{-2}	K_2S
Cobalt (II) oxide	Cobalt (II) is Co^{+2}, oxide is O^{-2}	CoO
Cobalt (III) oxide	Cobalt (III) is Co^{+3}, oxide is O^{-2}	Co_2O_3
Tin (II) sulfide	Tin (II) is Sn^{+2}, sulfide is S^{-2}	SnS
Tin (IV) sulfide	Tin (IV) is Sn^{+4}, sulfide is S^{-2}	SnS_2
Calcium phosphide	Calcium is Ca^{+2}, phosphide is P^{-3}	Ca_3P_2
Aluminum bromide	Aluminum is Al^{+3}, bromide is Br^{-1}	$AlBr_3$
Silver nitride	Silver is Ag^{+1}, nitride is N^{-3}	Ag_3N

Ternary compound - an ionic compound containing at least one POLYATOMIC ion.

Polyatomic Ion…ion made from two or more atoms (Reference Table E). Look them up and notice four things:

Table E
Selected Polyatomic Ions

H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	dimercury (I)	$Cr_2O_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$C_2H_3O_2^-$ / CH_3COO^-	acetate	NO_2^-	nitrite
		NO_3^-	nitrate
CN^-	cyanide	O_2^{2-}	peroxide
CO_3^{2-}	carbonate	OH^-	hydroxide
HCO_3^-	hydrogen carbonate	PO_4^{3-}	phosphate
$C_2O_4^{2-}$	oxalate	SCN^-	thiocyanate
ClO^-	hypochlorite	SO_3^{2-}	sulfite
ClO_2^-	chlorite	SO_4^{2-}	sulfate
ClO_3^-	chlorate	HSO_4^-	hydrogen sulfate
ClO_4^-	perchlorate	$S_2O_3^{2-}$	thiosulfate

1) They are all made of **more than one atom**

2) Almost all of them have a **NEGATIVE** charge

3) Almost all of them contain **OXYGEN**

4) The positively charged ones end in **–ium**, the negatively charged ones end in **–ide, -ite or –ate:**

TWO atoms: -ide (CN^-, OH^-, O_2^{2-})

LESS O's: -ite (NO_2^- = nitrite)
(hypo- means one less oxygen than –ite)

MORE O's: -ate (NO_3^- = nitrate)
(per- means one more oxygen than –ate)

The rules for writing ternary compounds are the same as for binary formulas, with the following modifications:

1) Parentheses surrounding the polyatomic ion and a subscript will be used when there is more than one polyatomic ion in the formula.

2) In naming, the + ion still retains its name (with the Stock System used where appropriate). The only + polyatomic ion of note is the ammonium (NH_4^+) ion, this retains its name.

3) The NEGATIVE ion will most likely be the polyatomic ion, which retains its name.

WRITING TERNARY FORMULAS GIVEN THE NAME

Name	How to write the formula (verify charges using periodic table)	Formula
Ammoniium sulfide	Ammonium is K^{+1}, sulfide is S^{-2}	$(NH_4)_2S$
Cobalt (II) cyanide	Cobalt (II) is Co^{+2}, cyanide is CN^{-1}	$Co(CN)_2$
Cobalt (III) sulfate	Cobalt (III) is Co^{+3}, sulfate is SO_4^{-2}	$Co_2(SO_4)_3$
Tin (II) nitrate	Tin (II) is Sn^{+2}, nitrate is NO_3^{-1}	$Sn(NO_3)_2$
Tin (IV) nitrate	Tin (IV) is Sn^{+4}, nitrate is NO_3^{-1}	$Sn(NO_3)_4$
Calcium phosphate	Calcium is Ca^{+2}, phosphate is PO_4^{-3}	$Ca_3(PO_4)_2$
Aluminum dichromate	Aluminum is Al^{+3}, dichromate is $Cr_2O_7^{-2}$	$Al_2(Cr_2O_7)_3$
Silver nitrite	Silver is Ag^{+1}, nitrite is NO_2^{-3}	$AgNO_2$

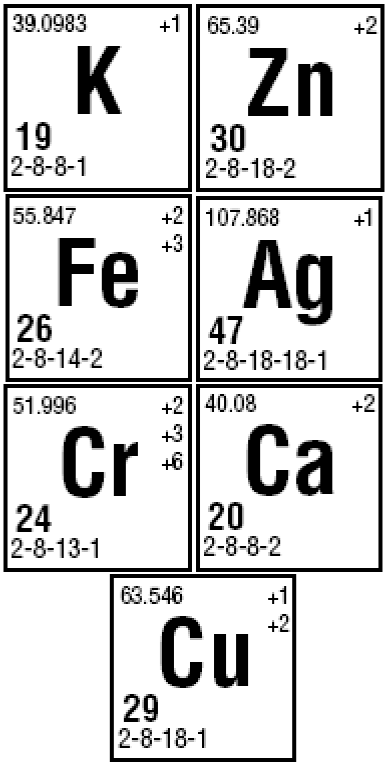


Table E
Selected Polyatomic Ions

H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	dimercury (I)	$Cr_2O_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$C_2H_3O_2^-$ / CH_3COO^-	acetate	NO_2^-	nitrite
		NO_3^-	nitrate
CN^-	cyanide	O_2^{2-}	peroxide
CO_3^{2-}	carbonate	OH^-	hydroxide
HCO_3^-	hydrogen carbonate	PO_4^{3-}	phosphate
		SCN^-	thiocyanate
$C_2O_4^{2-}$	oxalate	SO_3^{2-}	sulfite
ClO^-	hypochlorite	SO_4^{2-}	sulfate
ClO_2^-	chlorite	HSO_4^-	hydrogen sulfate
ClO_3^-	chlorate	$S_2O_3^{2-}$	thiosulfate
ClO_4^-	perchlorate		

NAMING TERNARY COMPOUNDS GIVEN THE FORMULAS

Formula	**How to determine name**	**Name**
KNO_3	K only has one charge, so K^{+1} is called **potassium**. NO_3^{-1} is called **nitrate.**	Potassium nitrate
$ZnSO_4$	Zn only has one charge, so Zn^{+2} is called **zinc**. SO_4^{-2} is called **oxide**.	Zinc sulfate
$Fe(NO_2)_2$	Fe has two charges listed, so use NO_2 to determine Fe's charge. Each NO_2 is –1, so two of them would be –2. To add up to zero, Fe has to be charged +2, called **iron (II)**. NO_2^{-1} is called **nitrite**.	Iron (II) nitrite
$Fe(NO_2)_3$	Fe has two charges listed, so use NO_2 to determine Fe's charge. Each NO_2 is –1, so three of them would be –3. To add up to zero, Fe has to be charged +3, called **iron (III)**. NO_2^{-1} is called **bromide**.	Iron (III) nitrite
Ag_2CO_3	Ag only has one charge, so Ag^{+1} is called **silver**. CO_3^{-2} is called **carbonate.**	Silver carbonate
$Cr_2(SO_3)_3$	Cr has three charges listed, so use SO_3 to determine Cr's charge. Each SO_3 is –2, so three of them would be –6. To add up to zero, Cr has to be charged +6 total. Since there are two Cr's, each one is charged +3, called **chromium (III)**. SO_3^{-2} is called **sulfite**.	Chromium (III) sulfite
$Cr(SO_3)_3$	Cr has three charges listed, so use SO_3 to determine Cr's charge. Each SO_3 is –2, so three of them would be –6. To add up to zero, Cr has to be charged +6, called **chromium (VI)**. SO_3^{-2} is called **sulfite**.	Chromium (VI) sulfite
$Ca_3(PO_4)_2$	Calcium only has one charge, so Ca^{+2} is called **calcium**. PO_4^{-3} is called **phosphate**.	Calcium phosphate
$CuCrO_4$	Cu has two charges listed, so use CrO_4 to determine Cu's charge. CrO_4 is –2, so to add up to zero, Cu has to be charged +2, called **copper (II)**. CrO_4^{-2} is called **chromate.**	Copper (II) chromate
Cu_2CrO_4	Cu has two charges listed, so use CrO_4 to determine Cu's charge. CrO_4 is –2, so to add up to zero, Cu has to be charged +2 total. Since there are two copper ions, each one must be charged +1, called **copper (I)**. CrO_4^{-2} is called **chromate**	Copper (I) chromate

NAMING AND FORMULA WRITING FOR MOLECULAR COMPOUNDS

Molecular formulas can be named in two ways:

1) The Stock System

This works the same as with ionic compounds. Use a Roman numeral to represent the charge of the first atom written in the formula (the one with lower electronegativity).

Molecular Formula	Name (Stock System)	Name (Stock System)	Molecular Formula
CO_2	Carbon (IV) oxide	Sulfur (II) oxide	SO
CO	Carbon (II) oxide	Sulfur (IV) oxide	SO_2
NO_2	Nitrogen (IV) oxide	Carbon (IV) chloride	CCl_4
NO_3	Nitrogen (VI) oxide	Nitrogen (III) chloride	NCl_3
N_2O_5	Nitrogen (V) oxide	Phosphorous (III) oxide	P_2O_3

2) The Prefix System

This system uses prefixes to describe how many atoms of each element are found in the molecule:

1 atom	2 atoms	3 atoms	4 atoms	5 atoms	6 atoms
mono- (or none)	di-	tri-	tetra- (or tetr-)	penta (or pent-)	hexa- (or hex-)

Molecular Formula	Name (Prefix System)	Name (Prefix System)	Molecular Formula
CO_2	Carbon dioxide	Sulfur trioxide	SO_3
CO	Carbon monoxide	Sulfur dioxide	SO_2
NO_2	Nitrogen dioxide	Carbon tetrachloride	CCl_4
NO_3	Nitrogen trioxide	dinitrogen tetroxide	N_2O_4
N_2O_5	Dinitrogen pentoxide	diphosphorous trioxide	P_2O_3

WARNING: THE PREFIX SYSTEM CAN NOT BE USED TO NAME IONIC COMPOUNDS. ONLY THE STOCK SYSTEM CAN BE USED FOR IONIC COMPOUNDS.

3) Formula Mass (HW: p. 30, 31)

Essential Question: What is the most important physical property any compound or element has when it comes to manufacturing?

Interpreting Chemical Formulas

Empirical Formulas tell you the simplest, whole-number, mole ratio of the elements in an ionic compound. The subscript tells you how many moles of that particular ion you have in one mole of the compound. If there is no subscript, then the number of moles of that ion is 1.

A Mole is to chemistry what a "dozen" is to donuts or eggs. Eggs and donuts are sold by the dozen, chemical formulas and reactions are written by the mole. **One mole = 6.02×10^{23} of anything**. That is the number of atoms that weigh their atomic mass in grams. For example, hydrogen has an atomic mass of 1.0, so one mole of hydrogen atoms will have a mass of 1.0 gram.

When Interpreting Formulas, if there are (parentheses) around the element, any subscript outside the parentheses multiplies all of the elements inside the parentheses by that amount.

$Ca(NO_3)_2$ has a 2 outside the parentheses, so double the number of atoms inside to get the total number of atoms of that element (2 N's and 6 O's). Since Ca is not inside the parentheses, it is not doubled. There is only 1 Ca in the formula.

Empirical Formula	# of Ions of Each Element In Compound	# of Moles of Each Element in a Mole of the Compound	Total # of Moles of Ions in a mole of the Compound
NaCl	1 Na, 1 Cl	1 Na, 1 Cl	1+1 = **2**
$CaCl_2$	1 Ca, 2 Cl	1 Ca, 2 Cl	1+2 = **3**
Al_2O_3	2 Al, 3 O	2 Al, 3 O	2+3 = **5**
KNO_3	1K, 1 N, 3 O	1K, 1 N, 3 O	1+1+3 = **5**
K_2SO_4	2 K, 1 S, 4 O	2 K, 1 S, 4 O	2+1+4 = **7**
$Al(NO_3)_3$	1 Al, 3 N, 9 O	1 Al, 3 N, 9 O	1+3+9 = **13**
$Al_2(SO_4)_3$	2 Al, 3 S, 12 O	2 Al, 3 S, 12 O	2+3+12 = **17**

Coefficients: tell you how many moles of the compound you have. To find the total number of ions you have, multiply the number of a particular ion by the coefficient. If there is no coefficient written, then the coefficient is 1.

Formula	Moles of Each Ion In One Mole of Compound	Formula With Coefficient	Moles of Each Ion in Given Number of Moles Of Compound
KBr	1 K, 1 Br	2 KBr	2 X (1 K, 1 Br) = **2 K, 2 Br**
MgF_2	1 Mg, 2 F	3 MgF_2	3 X (1 Mg, 2 F) = **3 Mg, 6 F**
$CaSO_4$	1 Ca, 1 S, 4 O	4 $CaSO_4$	4 X (1 Ca, 1 S, 4 O) = **4 Ca, 4 S, 16 O**
$Au(NO_2)_3$	1 Au, 3 N, 6 O	5 $Au(NO_2)_3$	5 X (1 Au, 3 N, 6 O) = **5 Au, 15 N, 30 O**
$Co(OH)_3$	1 Co, 3 O, 3 H	4 $Co(OH)_3$	4 X (1 Co, 3 O, 3 H) = **4 Co, 12 O, 12 H**
$Mg_3(PO_4)_2$	3 Mg, 2 P, 8 O	3 $Mg_3(PO_4)_2$	3 X (3 Mg, 2 P, 8 O) = **9 Mg, 6 P, 24 O**

How are the formulas arrived at experimentally?

Qualitative analysis - what elements are in the compound? Spectrometers are very handy tools for doing this.

Quantitative analysis - how much of each element is in the compound? Often, chemical reactions are used to determine this. Decomposition reactions break down compounds into their component elements.

Types of Chemical Formulas:

1) Molecular Formula - indicates the total number of atoms of each element needed to form the molecule.
MOLECULES ARE PARTICLES FORMED FROM THE COVALENT BONDING OF NONMETAL ATOMS.

A molecule of methane contains one atom of carbon and four atoms of hydrogen, so the molecular formula is CH_4.
A molecule of benzene contains six atoms of carbon and six atoms of hydrogen, so the molecular formula is C_6H_6.

2) Empirical Formulas
represents the simplest ratio in which the atoms combine to form a compound. Empirical formulas are used to represent **ionic compounds** which form crystals of alternating + and – charge instead of separate molecules. $CaCl_2$ is ionic, and the formula represents a ratio of 1 ion of Ca^{+2} to every 2 ions of Cl^{-1}. If the crystal contains 3000 Ca^{+2} ions, then it will contain 6000 Cl^{-1} ions. If the crystal contains one mole of Ca^{+2} ions, then it will contain two moles of Cl^{-1} ions.

All ionic compound formulas are empirical formulas. KCl has a 1:1 ratio, $CaCl_2$ has a 1:2 ratio and $Al_2(SO_4)_3$ has a 2:3:12 ratio, none of which can be simplified any further.

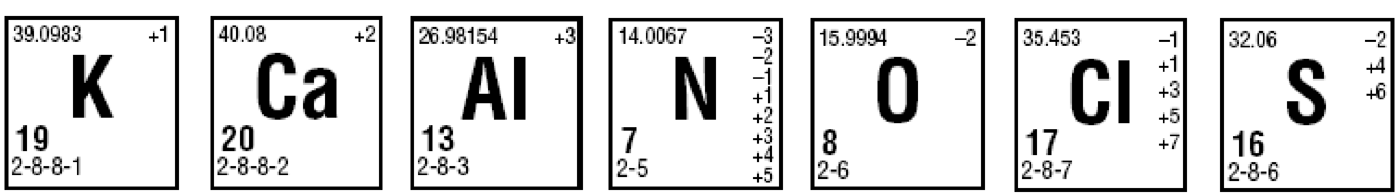
Molecular compounds like CH_4 (1:4), NO_2 (1:2) and H_2O (2:1) cannot be simplified any further. For them, the molecular formula is also the empirical formula. Others can be simplified. C_6H_6 (6:6) can be simplified to a 1:1 ratio, giving the empirical formula CH. H_2O_2 (2:2) can be simplified to 1:1, for an empirical formula of HO. C_4H_8 (4:8) can be simplified to a ratio of 1:2, giving the empirical formula of CH_2.

Gram Formula Mass (Molar Mass)

STOICHIOMETRY - The study of quantitative relationships that can be derived from chemical formulas and equations. In working with chemical formulas, these involve ratios between the number of moles of atoms of different elements that comprise the compound.
The mass of one mole of atoms of an element is equal to its atomic mass in grams. This is called the gram atomic mass of the element, and can be read directly off of the Periodic Table, **rounding to the nearest tenth**.
The mass of one mole of molecules or formula units is equal to the sum of the gram atomic masses of the atoms that make up a mole of a particular molecule or formula unit. This number represents the **gram formula mass**, also called just formula mass.

Determining the Gram Formula Mass of a Compound

Here are the elements that are in the compounds in the examples of how to determine the gram formula mass. Let's round each of their masses off to the nearest **TENTH** before doing the problems:



39.1	**40.1**	**27.0**	**14.0**	**16.0**	**35.5**	**32.1**

KCl: (1 K @39.1 g/mol each = **39.1 g/mole**) **+** (1 Cl @ 35.5 g/mole each = **35.5 g/mole**) = **74.6 g/mol**

$Ca(NO_3)_2$: (1 Ca @ 40.1 g/mol = **40.1 g/mol**) **+** (2 N @ 14 g/mol = **28.0 g/mol**) **+** (6 O @ 16.0 g/mol = **96.0 g/mol**) = **164.1 g/mol**

$Al_2(SO_4)_3$: (2 Al @ 27.0 g/mol = **54.0 g/mol**) **+** (3 S @ 32.1 g/mol = **96.3 g/mol**) **+** (12 O @ 16.0 g/mol = **192.0 g/mol**) = **342.3 g/mol**

A) CONVERTING MOLES TO GRAMS AND GRAMS TO MOLES

When you make a recipe, you have to go out and buy certain ingredients, bring them home, measure out the proper amounts and then follow the directions as to how to mix them and cook them to make the dish you want. The same thing goes in chemistry! If your company makes a particular product, there is a recipe that is followed to make that product. From extracting aluminum from ore to use in cans, foil or conduit to making the perfect jelly bean, there is a recipe that must be followed. That recipe is based on a CHEMICAL REACTION. To make twice as much product, simply double the recipe, just as you would for a double batch of chocolate chip cookies. The recipe for chemical reactions, though is made in MOLES, and our balances don't measure things out in moles. They measure in GRAMS! Here, then is how to convert back and forth between moles and grams, so you can follow your recipe!

1) Determining how many moles of substance there are in a certain mass of a substance

To convert grams to moles, divide by the formula mass: (given ~~g~~) / (~~g~~ /mole) = mole

How many moles of KCl are there in 182.9 grams of KCl?

1) Find the formula mass of KCl: (1 K @39.1 g/mol each = **39.1 g/mole**) **+** (1 Cl @ 35.5 g/mole each = **35.5 g/mole**) = **74.6 g/mol**

2) moles = (grams/ formula mass): 182.9 g / 74.6 g/mole = **2.45 moles**

How many moles of $Ca(NO_3)_2$ are there in 45.5 grams of $Ca(NO_3)_2$?

1) Find the formula mass of $Ca(NO_3)_2$: (1 Ca @ 40.1 g/mol each = **40.1 g/mole**) + (2 N @ 14.0 g/mol each = **28.0 g/mole**) + (6 O @16.0 g/mol each = **96.0 g/mol**) = **164.1 g/mol**

2) moles = (grams/ formula mass): 45.5 g / 164.1 g/mole = **0.277 moles**

2) Determining the mass of a certain number of moles of substance.

To convert moles to grams, multiply by formula mass: (given ~~moles~~ X g/ ~~moles~~ = g)

A 2.45 mole sample of KCl will have what mass in grams?

1) Find the formula mass of KCl: (1 K @39.1 g/mol each = **39.1 g/mole**) **+** (1 Cl @ 35.5 g/mole each = **35.5 g/mole**) = **74.6 g/mol**

2) grams = (moles X formula mass): 2.45 mole X 74.6 g/mole = **183 grams**

A 0.344 mole sample of $Ca(NO_3)_2$ will have what mass in grams?

1) Find the formula mass of $Ca(NO_3)_2$: (1 Ca @ 40.1 g/mol each = **40.1 g/mole**) + (2 N @ 14.0 g/mol each = **28.0 g/mole**) + (6 O @16.0 g/mol each = **96.0 g/mol**) = **164.1 g/mol**

2) grams = (moles X formula mass): 0.344 mole X 164.1 g/mole = **56.5 grams**

4) Percent Composition, Empirical and Molecular Formulas (HW: p. 32-35

Essential Question: How do we know how much of each element is in a compound, anyway?

A) PERCENTAGE COMPOSITION: The proportion by mass of the elements in a compound.

DETERMINATION OF PERCENT COMPOSITION

Qualitative analysis tells you WHAT elements are in a compound, but it takes QUANTITATIVE analysis to determine how much of each element is in a compound. One of the ways this is accomplished is to weigh a sample of the compound, then take the compound and decompose it into its individual elements, weigh each element and then calculate what the percent by mass of each element in the compound is.

1) Experimental Determination:

(given mass of element in sample / given mass of total sample) X 100%

A compound containing nitrogen and oxygen has a mass of 80.0 grams. Experiments show that the 80.0 grams is made up of 56.0 grams of oxygen and 24.0 grams of nitrogen. What is the % composition, by mass, of each element in the compound?

% O = (56.0/80.0) X 100 **= 70.0%** **%N** = (24.0/80.0) X 100 **= 30.0%**

2) Determination by formula: You need the periodic table to do this.

(formula mass of element in compound / formula mass of compound) X 100%

What is the % composition, by mass, of each element in SiO_2?

%Si = (28.1/60.1) X 100 = **46.8%** **%O =** (2 X 16.0 = 32.0), (32.0/60.1) X 100 = **53.2%**

What is the % composition, by mass, of each element in H_2O?

%H = (2 X 1.0 = 2.0), (2.0/18.0) X100 =**11%** **%O =** (16.0/18.0) = **88.9%** (doesn't add up to 100% because of sig figs)

What is the % composition, by mass, of each element in H_2SO_4?

%H = (2 X 1.0 = 2.0), (2.0/98.1) X 100 = **2.0%** **%S =** (32.1/98.1) X 100 = **32.7%** **%O =** (4 X 16.0 = 64.0), (64.0/98.1) X 100 = **65.2%**

3) Hydrates: Ionic solids with water trapped in the crystal lattice. The ratio of water to each mole of ionic compound is determined like this:

Ionic Compound's Formula • n H_2O

Where **n** is a whole number, and the dot shows that the water is not bonded to the crystal, but just trapped inside of it.

Crystals have open spaces between the atoms that make up the crystal. Just like food can get caught in the spaces between your teeth, water molecules can get trapped in the spaces between the ions in the crystal. Water molecules have charged ends, and they are attracted to the charged ions in the crystal. Some crystals are so good at attracting water molecules (**hygroscopic**) that they actually absorb moisture out of the air! An example of this would be those white packets you find in the box with new shoes. It's called "SILICA GEL DO NOT EAT". Silica gel is made up of crystals that are very hygroscopic. If any moisture gets into the box, it gets sucked up by the silica gel, so that your new shoes don't get all mildewy. Hygroscopic crystals are often used as **desiccants**, or drying agents. Because they are so good at absorbing moisture, you have to keep the container they are in tightly shut, or they will lose their ability to do their job quickly. Desiccants have a very poor shelf life because of this.

When the crystal has absorbed its fill of water molecules, it still appears dry. This is because the water is not in liquid form, it is in the form of individual molecules of water trapped between the crystal's ions. This new crystal is called a **hydrate**, or a **hydrated crystal**.

To find out how much water is trapped in the crystal, all you have to do is heat the crystal up to drive the water off. Weigh the crystal every so often. When the crystal's mass doesn't change for two weighings, that means that the water is all driven off and the crystal has no water left in it. This crystal is now **anhydrous**, or is called an **anhydride**. Now it is back to being able to reabsorb moisture back into it from the air!

Here is a typical experiment that is used to determine the mass of water trapped inside a hydrate:

1) Weigh an empty evaporating dish that you have just heated to make sure it is completely dry.

2) Add your hydrate to the evaporating dish and weigh it again. The mass of the hydrate equals mass of (hydrate + dish) – (dish).

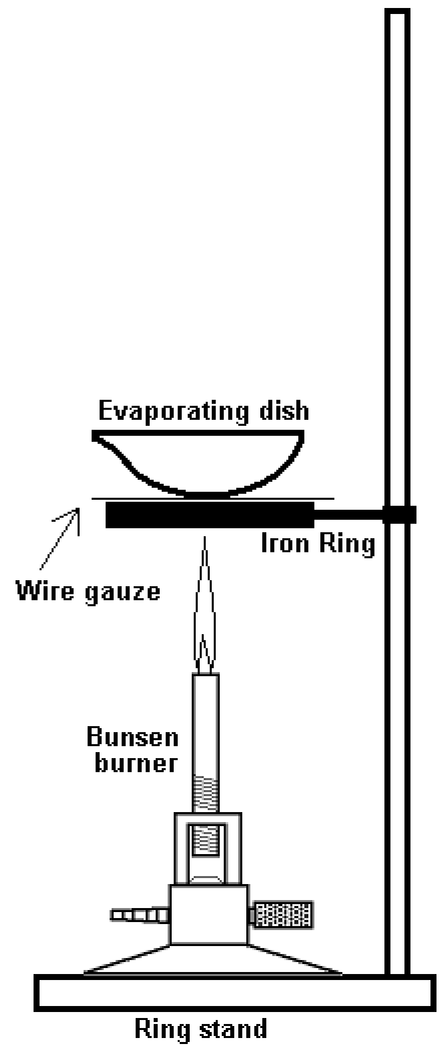
3) Place the evaporating dish with the hydrate in it above a Bunsen burner, on a ringstand. Once secure, light the Bunsen burner.

4) Heat the hydrate to drive the water out. Use a microspatula to stir the crystal so that the trapped water does not cause the crystal to clump up as it escapes, and to allow the crystal to get even heating.

5) Occasionally take the evaporating dish off the ringstand and weigh it. When it weighs the same two times in a row, all of the water has been driven out, and only the anhydride is left in the dish.

6) The mass of the water driven off = (mass of hydrate – mass of anhydride).

7) **The percent of water in the hydrate =**
(mass of water driven off / mass of hydrate) X 100.



Determining the % Of Water In A Hydrate Experimentally:

(given mass of water in hydrate / given mass of hydrate) X 100%

To determine the mass of the water that was driven out of the hydrate:

Take the mass of the original hydrate and subtract from it the mass of the anhydride.

To determine the percent of water in the hydrate:

Divide the mass of the water driven off from the hydrate by the total mass of the hydrate and multiply by 100.

2) A 10.40 gram sample of hydrated crystal is heated to a constant mass of 8.72 grams. This means all of the water has been driven out by the heat.

a) Calculate the mass of water that was driven out:

(mass of hydrate – mass of anhydride) = 10.40 grams – 8.72 grams = **1.68 grams of water driven off by the heat**

b) Calculate the percent, by mass, of water in this hydrate:

(given mass of water in hydrate / given mass of hydrate) X 100 = (1.68 grams / 10.40 grams) X 100 = 0.162 X 100 =

16.2% water by mass

3) An empty evaporating dish is found to have a mass of 29.993 grams. A sample of hydrated crystal is placed into the evaporating dish, the combined mass is 39.486 grams. The evaporating dish is placed on a ringstand and heated gently over a Bunsen burner flame, stirring constantly so that the water does not cause the drying crystal to clump together. The dish is removed from the flame, and weighed, then reheated, then weighed again, until the mass remains a constant 38.378 grams.

a) What was the mass of hydrated crystal that was put into the dish?

Mass of hydrated crystal = (mass of crystal + dish) – (mass of dish) = 39.486 g – 29.993 g = **9.493 grams of hydrate**

b) What was the mass of water that was driven out of the hydrate?

The total mass was 39.486 grams before heating, and 38.387 g after heating. The difference is the mass of water:

39.486 g – 38.387 g = **1.099 grams of water driven off by the heat**

c) Calculate the percent, by mass, of water in the hydrate:

(given mass of water in hydrate / given mass of hydrate) X 100 = (1.099 g / 9.493 g) X 100 = 0.116 X 100 =

11.6% water by mass

Determining the % Of Water In A Hydrate Given The Hydrate's Formula:

(formula mass of water in hydrate / formula mass of hydrate) X 100%

Water (H_2O) has a formula mass of (2 H @ 1.0 each = **2.0 g/mol**) + (1 O @ 16.0 g/mol each = **16.0 g/mol**) =

18.0 g/mol

If the coefficient of H_2O is 2, then the water weighs (2 X 18.0) = 36.0 grams in the hydrate.

If the coefficient of H_2O is 4, then the water weighs (4 X 18.0) = 72.0 grams in the hydrate.

To determine the formula mass of the hydrate:

1) Find the formula mass of the crystal.
2) Find the formula mass of the number of moles of water trapped in the hydrate (after the dot)
3) Add the two numbers together.

To determine the percent of water in the hydrate:

Divide the formula mass of the water in the hydrate by the total formula mass of the hydrate and multiply by 100.

What is the % composition, by mass, of each element in $CuCO_3 \bullet 2\ H_2O$?

Neat little trick here. There are 2 molecules of water trapped in each formula unit of $CuCO_3$. That's what the • means. This is a **hydrate**.
The formula mass = (1 Cu X 63.5 g/mol) + (1 C X 12.0 g/mole) + (5 O X 16.0 g/mol) + (4 H X 1.0 g/mol) = 159.5 g/mole.

Why 5 oxygens? Well, you have three of them in $CuC\mathbf{O_3}$ and two more in the **2** $H_2\mathbf{O}$.

Why 4 hydrogens? Well, you have four of them in **2** $\mathbf{H_2}O$.

%Cu = (63.5/159.5) X 100 = **39.8%** **%C =** (12.0/159.5) X 100 = **7.52%**

%O = (80.0/159.5) X 100 = **50.2%** **%H =** (4.0/159.5) X 100 = **2.5%**

What is the % composition, by mass, of H_2O in $CuCO_3 \bullet 2\ H_2O$?

The formula mass of the hydrate is 159.5 g/mole. There are 2 H_2O's, each one weighs
(2 H X 1.0 g/mole) + (1 O X 16.0 g/mole) = 18.0 g/mole. Therefore, % H_2O = [(2.0 X 18.0 = 36.0), (36.0) / (159.5) X 100 = **22.6%**

B) EMPIRICAL FORMULAS - The simplest, whole number, mole ratio of elements in a compound. This definition tells you how to get it!

1) Given the mass of each element in a sample of compound:

a) Convert mass to moles by dividing the mass of each element by the atomic mass of that element (mole ratio)
b) Divide all of the resulting mole numbers by the smallest mole value. This will give the smallest mole ratio.
c) Write the formula.

A 55.0 gram sample of a compound containing 0.9 g H, 12.2 g N, and 41.9 g O.

a) Convert mass to moles (divide by formula mass):

(0.9 g H) / (1.0 g/mol) = **0.9 moles H** (12.2 g N) / (14.0 g/mol) = **0.871 moles N** (41.9 g O) / (16.0 g/mol) = **2.62 moles O**

b) Divide all of the resulting mole numbers by the smallest mole value (0.871 is the smallest here), rounding your answer to the nearest whole number:

(0.9 moles H) / (0.871) = **1 mole H** (0.871 moles N) / (0.871) = **1 mole N** (2.62 mol O) / (0.871) = **3 moles O**

c) Write the formula, keeping the element symbols in the same order as presented in the problem:

$H\ N\ O_3$

2) Given the percentage composition of the compound:

a) Assume you have 100.0000000 (etc.) grams of compound. The percentages will convert themselves to grams.
b) Convert mass to moles by dividing the mass of each element by the atomic mass of that element (mole ratio)
c) Divide all of the resulting mole numbers by the smallest mole value
d) Write the formula.

A compound is found to have the following percent composition: 20.0 % Mg, 26.7 % S, 53.3 % O.

a) Assume you have 100 grams of compound. The percentages will convert themselves to grams.
If we have 100 grams of this compound, 20.0% of 100g is **20.0 g of Mg**, 26.7% of 100g is **26.7 g of S** and 53.3% of 10 g is **53.3 g of O**.

b) Convert mass to moles

(20.0 g Mg) / (24.0 g/mol)= **0.833 moles Mg** (26.7 g S) / (32.1 g/mol)= **0.831 mole S** (53.3 g O) / (16.0 g/mol)= **3.33 moles O**

c) Divide all of the resulting mole numbers by the smallest mole value (0.83 is the smallest value here):

(0.833 mole Mg) / (0.831) = **1 moles Mg** (0.831 mol S) / (0.831) = **1 mole S** (3.33 mol O) / (0.831) = **4 mol O**

d) Write the formula, keeping the element symbols in the same order as presented in the problem:

$Mg\ S\ O_4$

C) MOLECULAR FORMULAS - Molecular formulas are simply whole-number multiples of empirical formulas.

The empirical formula CH_2 has the following possible molecular formulas:

Empirical Formula, which has an…	Empirical Formula Mass of…	Multiply by…	…and get the Molecular formula…	…with a molecular mass of…
CH_2	14	1	CH_2 (not a possible molecular formula, C must form 4 bonds, not 2)	14
CH_2	14	2	C_2H_4	28
CH_2	14	3	C_3H_6	42
CH_2	14	4	C_4H_8	56
CH_2	14	5	C_5H_{10}	70

To determine the molecular formula of a compound given its molecular mass and empirical formula:

a) Determine the formula mass of the empirical formula.
b) Divide the molecular mass by the empirical mass. This will give you a whole-number multiple that tells you how many times larger the molecular formula is than the empirical formula
c) Multiply the whole number by the empirical formula. This will give the molecular formula.

The empirical formula of a compound is C_2H_3, and the molecular mass is 54.0 grams/mole. What is the molecular formula?

a) Determine the formula mass of the empirical formula.

C_2H_3 = (2 C X 12.0 g/mol) + (3 H X 1.0 g/mol) = **27.0 g/mole**

b) Divide the molecular mass by the empirical mass. This will give you a whole-number multiple that tells you how many times larger the molecular formula is than the empirical formula

(54.0 g/mol) / (27.0 g/mol) = **2.00**

c) Multiply the whole number by the empirical formula. This will give the molecular formula.

2.00 X C_2H_3 = **C_4H_6**

What is the molecular formula of a compound whose empirical formula is CH with a molecular mass of 52.0 grams/mole?

a) Determine the formula mass of the empirical formula.

CH = (1 C X 12.0 g/mol) + 1 H X 1.0 g/mol) = **13.0 g/mole**

b) Divide the molecular mass by the empirical mass. This will give you a whole-number multiple that tells you how many times larger the molecular formula is than the empirical formula

(52.0 g/mol) / (13.0 g/mol) = **4.00**

c) Multiply the whole number by the empirical formula. This will give the molecular formula.

4.00 X CH = **C_4H_4**

5) Stoichiometry of Formulas (HW: p. 36, 37)

Essential Question: How is the mole like the Grand Central Station of chemistry math?

Physical Quantities Directly Associated With the Mole

1) Mass - the mass of on one mole of a substance equals its formula mass (you did this yesterday)

a) mass --> mole (divide by formula mass)

How many moles do 22.0 grams of CO_2 weigh?
(22.0 g / 44.0 g/mole = **0.500 moles**)

b) mole --> mass (multiply by formula mass)

How many grams in 3.0 moles of $Ca(OH)_2$?
(3.0 moles X 64.1 g/mole = **190 grams**)

2) Molar Volume - the volume occupied by one mole of an ideal gas at STP is 22.4 liters, measured in liters/mole

a) volume --> mole (divide by 22.4 L/mole. You can use this number thanks to Avogadro's Hypothesis.)

How many moles of N_2 gas are occupied by 45.0 L of N_2 at STP?
(45.0 L) / (22.4 L/mol) = **2.01 moles of N_2 gas**

How many moles of C_3H_8 gas are occupied by 2.00 L of C_3H_8 at STP?
(2.00 L) / (22.4 L/mol) = **0.0893 moles of C_3H_8 gas**

44.8 L of CO_2 (g) equals how many moles at STP?
(44.8 L) / (22.4 L/mol) = **2.00 moles of CO_2 gas**

b) mole--> volume (multiply by 22.4 L/mole.)

0.5 moles of CO_2 (g) occupies what volume at STP?
(0.5 moles) X (22.4 L/mole) = **11.2 L**

How much volume will be occupied by 14.3 moles of He gas at STP?
(14.3 moles) X (22.4 L/mole) = **320. L**

How much volume will 0.25 moles of SO_2 gas occupy at STP?
(0.25 moles) X (22.4 L/mole) = **5.6 L**

3) Number of particles - the number of molecules, atoms or formula units in a mole equals 6.02 X 10^{23} particles/mole.

a) number --> mole (divide by 6.02 X 10^{23} particles/mole.)

How many moles of HCl contain 1.806 X 10^{24} molecules of HCl?
(1.806 X 10^{24} molecules) / (6.02 X 10^{23} molecules/mole) = **3.00 moles HCl**

How many moles of BaF_2 contain 3.01 X 10^{23} formula units of BaF_2?
(3.01 X 10^{23} formula units) / (6.02 X 10^{23} formula units/mole) = **0.500 moles BaF_2**

How many moles of Fe contain 9.03 X 10^{23} atoms of Fe?
(9.03 X 10^{23} atoms) / (6.02 X 10^{23} atoms/mole) = **1.50 moles Fe**

b) mole --> number (multiply by 6.02×10^{23} particles/mole.)

How many molecules of HCl are there in 0.500 moles of HCl?
(0.500 moles) X (6.02×10^{23} molecules/mole) = **3.01×10^{23} molecules HCl**

How many formula units of BaF_2 are there in 5.00 moles of BaF_2?
(5.00 moles) X (6.02×10^{23} formula units/mole) = **3.01×10^{24} formula units BaF_2**

How many atoms of Fe are there in 3.50 moles of Fe?
(3.50 moles) X (6.02×10^{23} atoms/mole) = **2.11×10^{24} atoms of Fe**

4) Number of atoms, given the number of molecules or formula units

Multiply the number of atoms in the formula by the number of molecules/formula units

How many atoms are there in 3.01×10^{23} molecules of HCl?

HCl contains two atoms in the formula, so (3.01×10^{23} molecules) X (2 atoms/molecule) = **6.02×10^{23} atoms**

How many atoms are there in 3.01×10^{24} formula units BaF_2?

BaF_2 contains three atoms in the formula, so (3.01×10^{24} formula units) X (3 atoms/formula unit) = **9.03×10^{23} atoms**

Two-Step Problems

What happens if you are given liters of gas and are asked to find how many grams it weighs? Don't panic! Convert to moles first!

How many grams of CO_2 (g) occupy 1.12 L?
First convert L → moles: (1.12 L) / (22.4 L/mole) = 0.0500 moles
Then convert moles → grams: (0.0500 moles) X (44.0 g/mole) = **2.20 grams**

How many molecules of CO_2 (g) occupy 44.8 L?
First convert L → moles: (44.8 L) / (22.4 L/mole) = 2.00 moles
Then convert moles → molecules: (2.00 moles) X (6.02×10^{23} molecules/mole) = **1.20×10^{24} molecules**

How much volume does 51.0 g of NH_3 (g) occupy at STP?
First convert grams → moles: (51.0 g) / (17.0 g/mole) = 3.00 moles
Then convert moles → L: (3.00 moles) X (22.4 L/mole) = **67.2 L**

How much volume will 3.01×10^{23} molecules of NH_3 (g) occupy at STP?
First convert molecules → moles: (3.01×10^{23} molecules) / (6.02×10^{23} molecules/mole) = 0.500 moles
Then convert moles → L: (0.500 moles) X (22.4 L/mole) = **11.2 L**

How many grams will 3.01×10^{23} molecules of NH_3 (g) weigh?
First convert molecules → moles: (3.01×10^{23} molecules) / (6.02×10^{23} molecules/mole) = 0.500 moles
Then convert moles → grams: (0.500 moles) X 17.0 g/mole) = **8.50 grams**

How many formula units are found in a 80.0 gram sample of NaOH?
First convert grams → moles: (80.0 g) / (40.0 g/mole) = 2.00 moles
Then convert moles → molecules: (2.00 moles) X (6.02×10^{23} formula units/mole) = **1.20×10^{24} formula units**

Student Name:________________________________ Grades: _____, _____, _____, _____, _____, _____, _____
TOC FW&N FM % Comp EF MF SOF

1) Types of Compounds Homework

A) Identify the type of compound indicated in the table of properties below and explain why it is that type:

Compound	Melting Point	Boiling Point	Electrical Conductivity (s)	Electrical Conductivity (l)
A	1074 K	2556 K	No	Yes
B	42 K	88 K	No	No
C	303 K	2676 K	Yes	Yes
D	3820 K	5100 K	No	No

A_________________________, because____

B_________________________, because____

C_________________________, because____

D_________________________, because____

B) Multiple Choice and Short-Answer Questions: Place your answer in the space in front of each question.

_____1) Which of the following substances is molecular?
a) NaCl b) CO_2 c) K_2O d) C

Explain how you can tell:____

Explain why one of the wrong choices is incorrect:

Choice:_____ Why:____

_____**2) Which of the following substances has a very high melting point and does not conduct electricity in the liquid phase?**
a) NaCl b) CO_2 c) CH_4 d) SiO_2

Explain how you can tell:____

Explain why one of the wrong choices is incorrect:

Choice:_____ Why:____

3) Carbon dioxide (CO_2) has a boiling point of -57^{o}C and water (H_2O) has a boiling point of 100^{o}C. A molecule of CO_2 has a mass of 44.0 amu and a molecule of H_2O has a mass of 18.0 amu. Why then, if CO_2 is a larger molecule than H_2O, does H_2O have a boiling point that is so much higher? Explain in terms of the polarity of each molecule and its resulting effect on the strength of the attractive forces between the molecules.

C) For each of the following molecules represented by structural formulas indicate

a) **if the molecule is polar or nonpolar.**
b) *******If polar, draw the dipole moment and which side is partially positive and which side is partially negative.*******

   If nonpolar, then skip this step and move on to c).

c) **Identify the shape of the molecule**
d) **identify the type of attractive force that will hold molecules of this substance together in the liquid and solid phase.**

<table>
<tr><th>Molecule (do B in this box)</th><th>Polar or Nonpolar</th><th>Shape</th><th>IMAF Type</th><th>Dot Diagram</th></tr>
<tr><td>H – Cl</td><td></td><td></td><td></td><td></td></tr>
<tr><td>H – H</td><td></td><td></td><td></td><td></td></tr>
<tr><td>H – O
|
H</td><td></td><td></td><td></td><td></td></tr>
<tr><td>H - N – H
|
H</td><td></td><td></td><td></td><td></td></tr>
<tr><td>Cl
|
Cl – C – Cl
|
Cl</td><td></td><td></td><td></td><td></td></tr>
<tr><td>O = C = O</td><td></td><td></td><td></td><td></td></tr>
<tr><td>H - Br</td><td></td><td></td><td></td><td></td></tr>
</table>

2) Naming and Writing Binary and Ternary Formulas Homework

1) <u>Binary Compounds</u>:

a) Write the formulas of the following binary ionic compounds:

Names	Formula
Zinc bromide	
Lead (II) nitride	
Potassium oxide	
Iron (III) chloride	
Lithium fluoride	
Aluminum sulfide	
Chromium (VI) oxide	
Magnesium iodide	
Copper (I) sulfide	
Lead (IV) oxide	

b) Name the following binary compounds:

Formula	Name
$MgCl_2$	
$AgBr$	
$AuCl$	
Al_2O_3	
Cu_2S	
CuS	
Na_2O	
FeO	
Li_2S	
Fe_2O_3	
$SrCl_2$	
PbO_2	
MnO_2	
AlP	
K_3P	
AuP	
ZnO	

2) <u>Ternary Compounds</u>:

a) Write the formulas of the following ternary ionic compounds:

Names	Formula
Zinc acetate	
Lead (II) nitrite	
Potassium oxalate	
Iron (III) chlorate	
Lithium permanganate	
Aluminum sulfate	
Chromium (VI) cyanide	
Ammonium nitrate	
Copper (I) sulfite	
Lead (IV) hypochlorite	

b) Name the following ternary compounds:

Formula	Name
K_2SO_4	
$AgNO_3$	
$Au(ClO_3)_3$	
$Al_2(SO_3)_3$	
Cu_3PO_4	
$Cu_3(PO_4)_2$	
$NaC_2H_3O_2$	
$Fe(OH)_2$	
$Li_2Cr_2O_7$	
$FeCO_3$	
$Sr(ClO_2)_2$	
$Ca(CN)_2$	
$AlPO_4$	
$Pb(SO_4)_2$	
$(NH_4)_2O$	
$Al_2(CO_3)_3$	
KNO_3	
Au_2CrO_4	
$Zn(CH_3COO)_2$	

3) **Both Binary and Ternary Compounds**

a) Indicate whether the formula is binary or ternary, then write the formula of the compound given its name

Formula	Type	Name
NaCl		
Na_2O		
H_2SO_4		
$Fe_3(PO_4)_2$		
HBr		
$CaSO_4$		
K_2O		
$ZnSO_3$		
$(NH_4)_2SO_4$		
HCH_3COO		
$PbCl_2$		
$CuSO_4$		
Fe_2S_3		
$AgNO_2$		

b) Indicate whether the formula is binary or ternary, then write the name of the compound given its formula.

Name	Type	Formula
Lead (IV) chloride		
Potassium oxide		
Lithium dichromate		
Sodium phosphate		
Ammonium nitrate		
Lead (IV) sulfide		
Silver chloride		
Zinc sulfate		
Aluminum iodide		
Magnesium hydroxide		
Tin (II) acetate		
Ammonium oxide		
Potassium permanganate		
Chromium (III) perchlorate		
Iron (III) chloride		
Sodium cyanide		

4) Fill in the blanks for each molecular compound, using the clue given for each:

Formula	Name (Stock System)	Name (Prefix System)
SO_3		
	Phosphorous (V) oxide	
		Phosphorous trichloride
H_2O		
	Hydrogen nitride	
		Carbon tetrabromide
HCl		
	Hydrogen sulfide	

3) Formula Mass Homework

A) Determine the number of moles of each element present in one mole of each of the following compounds:

Formula	Number of:	Number of:	Number of:
NaCl	Na:	Cl:	
$CaBr_2$	Ca:	Br:	
K_2SO_4	K:	S:	O:
$Fe_2(CO_3)_3$	Fe:	C:	O:
SO_3	S:	O:	
CH_3COOH	C:	H:	O:

B) Determine the total number of moles of atoms in the following formulas:

Formula	Total Number of Moles of Atoms	Formula	Total Number of Moles of Atoms	Formula	Total Number of Moles of Atoms
NaCl		2 $NaNO_3$		4 $(NH_3)_2SO_4$	
$MgSO_4$		$Pb_3(PO_4)_2$		3 $K_2Cr_2O_7$	

C) Identify each of the following as an empirical or molecular formula. If a formula is molecular, write its empirical formula.

Formula	Empirical or Molecular?	Simplify if Molecular	Formula	Empirical or Molecular?	Simplify if Molecular
NaCl			N_2O_4		
C_2H_6			$Ra(CN)_2$		
$Ba(NO_3)_2$			$C_6H_{12}O_6$		

D) Identify each statement as having been determined from quantitative or qualitative analysis.

1) The compound consists of carbon, hydrogen, and chlorine.

2) In a molecule of this compound, there are 6 atoms of carbon, four atoms of hydrogen, and two atoms of chlorine.

E) Calculate the formula masses in g/mole of the following compounds (use your Periodic Table to determine the gram atomic masses, and round them off to the nearest tenth).

Formula	Show your work	Gram Formula Mass
$NaCl$		
$NaOH$		
MgF_2		
KNO_3		
Na_2SO_4		
H_2SO_4		
AlF_3		
$ZnSO_4$		
H_2O		
HNO_3		
$Mg(OH)_2$		
CH_4S		

F) Determine how many moles of each substance is represented by the given mass. Use the formula masses you calculated above to solve the problems.

Mass and Formula	Show Your Work Here	Moles of substance
36 grams of H_2O		
29.25 grams of $NaCl$		
60. grams of $NaOH$		
71 grams of Na_2SO_4		
15.75 grams of HNO_3		

G) Determine the mass of the given number of moles of each substance. Use the formula masses you calculated above to solve the problems.

Moles Of Compound	Show Your Work Here	Grams of Substance
2.00 moles of AlF_3		
10.0 moles of $Mg(OH)_2$		
20.0 moles of $NaCl$		
0.500 moles of CH_4S		
4.00 moles of H_2SO_4		

H) Sodium hydroxide and zinc chloride is reacted to form sodium chloride solution with zinc chloride powder suspended in it. The zinc chloride is filtered out and dried. Its mass is 5.887 grams. How many moles is this?

4) Percent Composition, Empirical and Molecular Formulas

A) Determine the percentage of each element in the following compounds. Show all work, make sure answers are properly rounded.

1) CO

%C =__________ %O =__________

2) Na_2O

%Na =_________ %O =_________

3) $KClO_3$

%K =__________ %Cl =__________ %O =__________

4) Na_2CO_3

%Na =__________ %C =__________ %O =__________

5) $(NH_4)_2SO_3$

%N =__________ %H =__________ %S =__________ %O =__________

6) $CuSO_4 \cdot 5\,H_2O$

%Cu =__________ %S =__________ %O =__________ %H =__________

B) Determine the percent by mass of water in each of the following hydrates:

1) $CuSO_4 \cdot 5\,H_2O$

Formula Mass of the ____H_2O: ______________________ Formula Mass Of $CuSO_4 \cdot 5\,H_2O$: _________________

Calculate the % Of Water in $CuSO_4 \cdot 5\,H_2O$:

2) A 8.60 gram sample of hydrated crystal is heated to a constant mass of 6.22 grams. This means all of the water has been driven out by the heat.

a) Calculate the mass of water that was driven out:____

b) Calculate the percent, by mass, of water in this hydrate:

3) An empty evaporating dish is found to have a mass of 32.982 grams. A sample of hydrated crystal is placed into the evaporating dish, the combined mass is 36.553 grams. The evaporating dish is placed on a ringstand and heated gently over a Bunsen burner flame, stirring constantly so that the water does not cause the drying crystal to clump together. The dish is removed from the flame, and weighed, then reheated, then weighed again, until the mass remains a constant 35.378 grams.

a) What was the mass of hydrated crystal that was put into the dish?____

b) What was the mass of water that was driven out of the hydrate?____

c) Calculate the percent, by mass, of water in the hydrate:

C) Determine the empirical formulas for each question below, showing all work.

1) A 30.0 gram sample of substance is found to contain 15.6 g of carbon, 3.9 grams of hydrogen, and 10.5 grams of oxygen. What is the empirical formula of this compound?

2) A compound is found to have the following composition by mass:
25% potassium, 35% manganese, and 40.% oxygen. What is the empirical formula of this compound?

3) A compound contains 11.5 grams of sodium, 7.0 grams of nitrogen, and 1.0 gram of hydrogen. What is the empirical formula of this compound?

4) A compound is found to consist of 11.1% hydrogen and 88.9% oxygen. What is the empirical formula of this compound?

5) A compound is decomposed and discovered to contain 25.9% nitrogen by mass and 74.1% oxygen by mass. What is the empirical formula of this compound?

D) Determine the molecular formulas for each question below, showing all work.

1) The empirical formula of a compound is found to be CH, and the molecular mass has been determined to be 78.0 g/mole. What is the molecular formula of this compound?

2) The empirical formula of a compound is found to be HO, and the molecular mass has been determined to be 34.0 g/mole. What is the molecular formula of this compound?

3) The empirical formula of a compound is found to be NO_2, and the molecular mass has been determined to be 92.0 g/mole. What is the molecular formula of this compound?

4) The empirical formula of a compound is found to be CH_2O, and the molecular mass has been determined to be 180.0 g/mole. What is the molecular formula of this compound?

5) The empirical formula of a compound is found to be CH_3O, and the molecular mass has been determined to be 62.0 g/mole. What is the molecular formula of this compound?

6) The empirical formula of a compound is found to be CH_2 and the molecular mass has been determined to be 84.0 g/mole. What is the molecular formula of this compound?

5) Stoichiometry of Formulas Homework

A) Solve the following mole relationship problems, rounding answers to proper sig figs (SHOW ALL WORK):

1) How many grams will $1.8 \cdot 10^{24}$ molecules of H_2O weigh?

2) How much volume (in L) will 10.0 g of He (g) occupy at STP?

3) How many grams will 11.2 L of O_2 weigh at STP?

4) How many molecules are there in 180. g of SiO_2?

5) How many liters will $3.01 \cdot 10^{23}$ molecules of N_2 occupy at STP?

6) How many grams will 4 moles of AgCl weigh?

7) How many moles are there in a sample of $BaCl_2$ with a mass of 416 g ?

8) 10. grams of CO_2 (s) will sublime to form how many liters of CO_2 at STP(g)?

9) 10 moles of NaCl are needed to carry out a specific reaction. How many

a) grams is this?

b) formula units (the ionic equivalent of *molecules*) is this?

10) A sample of N_2 gas occupies 44.8 liters at STP. How many

a) moles of gas is this?

b) grams of N_2 is this?

c) molecules is this?

d) atoms of N is this? (# atoms of N in the formula N_2 X # of molecules)

11) A sample of CH_4 gas contains $3.01 \cdot 10^{23}$ molecules. How many

a) moles of gas is this?

b) grams of CH_4 is this?

c) liters of gas at STP is this?

d) atoms is this? (Total # atoms in the formula CH_4 X # of molecules)

12) A sample of $CaCO_3$ weighs 200 grams. How many

a) moles of $CaCO_3$ is this?

b) formula units of $CaCO_3$ is this?