

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 (δ^+) end of one polar molecule to the N, O or F (δ^-) 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 kernels.
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. 21, 22)

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 all 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 , $\text{C}_6\text{H}_{12}\text{O}_6$
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!

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

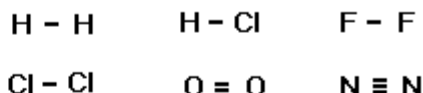
Dot Diagrams and Structural Formulas

When nonmetals atoms form molecules, the unpaired electrons pair up so that each atom ends up with 8 valence electrons (called a STABLE OCTET) except for hydrogen, which is so small that it has to be satisfied with two. Each shared pair of electrons is a bond, and can be represented by a line as shown below:

Dot Diagram		Structural Formula
H : H	replace the two dots with a line and you get	H - H
H : $\ddot{\text{Cl}}$:	replace the two dots between the H and the Cl with a line:	H - Cl
H : $\ddot{\text{O}}$: H	replace the two pairs of dots between the H's and the O with a line:	H - O H
H : $\ddot{\text{N}}$: H H	replace the three pairs of dots between the N and the H's with a line and get:	H - N - H H
: $\ddot{\text{O}}$:: $\ddot{\text{O}}$:	replace the two pairs of dots between the O's with two lines:	O = O

Shapes of Molecules

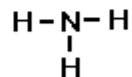
LINEAR: Made of two atoms. Atoms may be bonded by single, double or triple bonds.



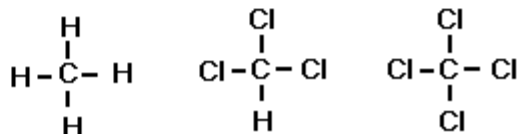
BENT: Made of three atoms, one of which is central and the other two atoms are bonded to it.



PYRAMIDAL: Made of four atoms, one of which is in the center and the other three are bonded to it.



TETRAHEDRAL: Made of five atoms, one of which is in the center and the other four are bonded to it.



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 either 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.

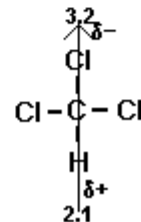
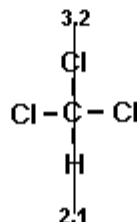
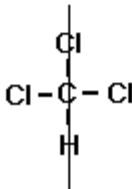
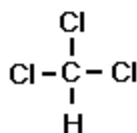
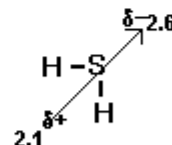
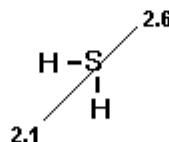
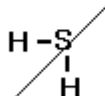
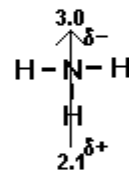
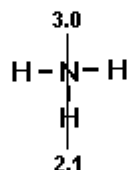
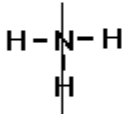
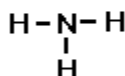
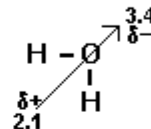
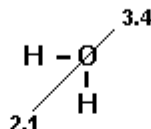
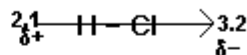
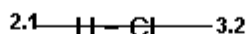
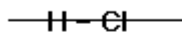
Polar molecules have asymmetric (without symmetry) electron distributions.

Draw an arrowhead on the side with the higher EN. This is the Dipole Moment. The side with higher EN has a greater pull on the electrons, so is charged δ^- . The other side is charged δ^+ . You have completed a polar molecule!

Structural
Formula

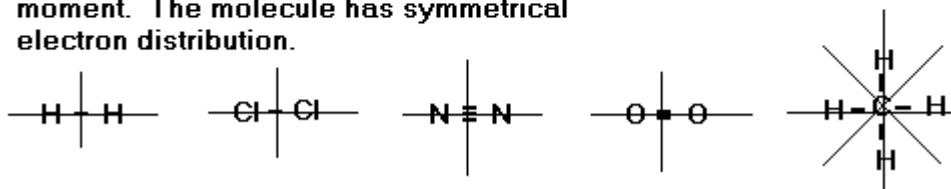
Lines of symmetry:
either side of the line is a
mirror image

Look up the ENs of the
atoms on the ends of
the line of symmetry



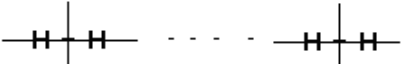
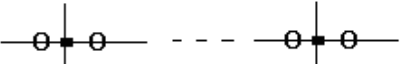
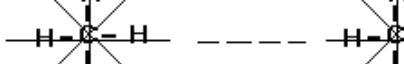
NONPOLAR MOLECULES). If the molecule has a symmetrical shape, then the electrons are distributed evenly through the molecule, and the whole molecule is nonpolar (even if it contains polar bonds). Nonpolar molecules have equal pull of electrons on all sides of the molecule, so no side develops a pole. 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_{10} , 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 ($\text{C}_{60}\text{H}_{12}$, paraffin, which is one molecule that mothballs have been made of).

NONPOLAR MOLECULES have two or more lines of symmetry. The electronegativities along these lines of symmetry are equal, so there is an equal pull on electrons from all sides of the molecule. There is no dipole moment. The molecule has symmetrical electron distribution.

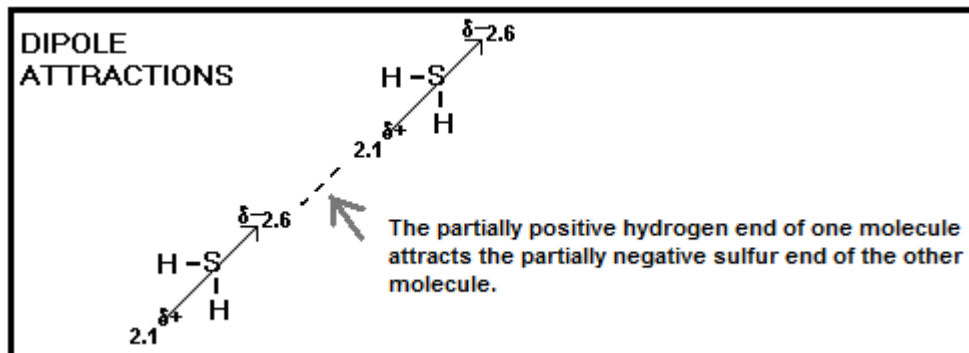


Attractive Forces – what holds molecules together in the solid and liquid phase

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.

LONDON DISPERSION FORCES		
 Hydrogen molecules have no oppositely charged ends. Any attractions between them are due to electron movement within the molecules. Hydrogen boils at 20 K, 20 degrees above absolute zero.	 Oxygen atoms are larger than H atoms, so O_2 is larger than H_2. The electron movement in a larger molecule tends to create stronger London dispersion forces. O_2 has a boiling point of 90 K.	 Methane molecules are larger than the ones above, so methane's London dispersion forces are the strongest of all three molecules here. Its boiling point is 112 K.

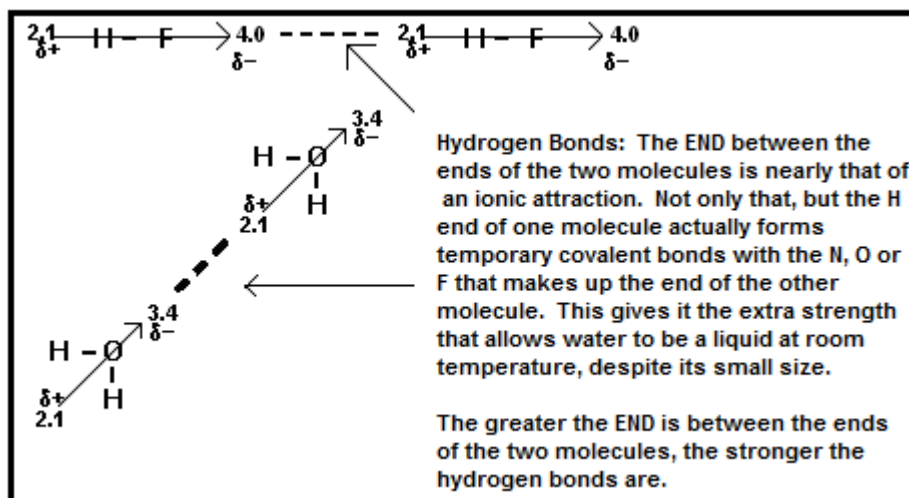
Polar molecules form DIPOLE attraction: the simple attraction of the oppositely charged ends of two molecules. The partially positive end of one molecule attracts to the partially negative end of a different molecule. This attraction allows these substances to exist as solids and liquids at higher temperatures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

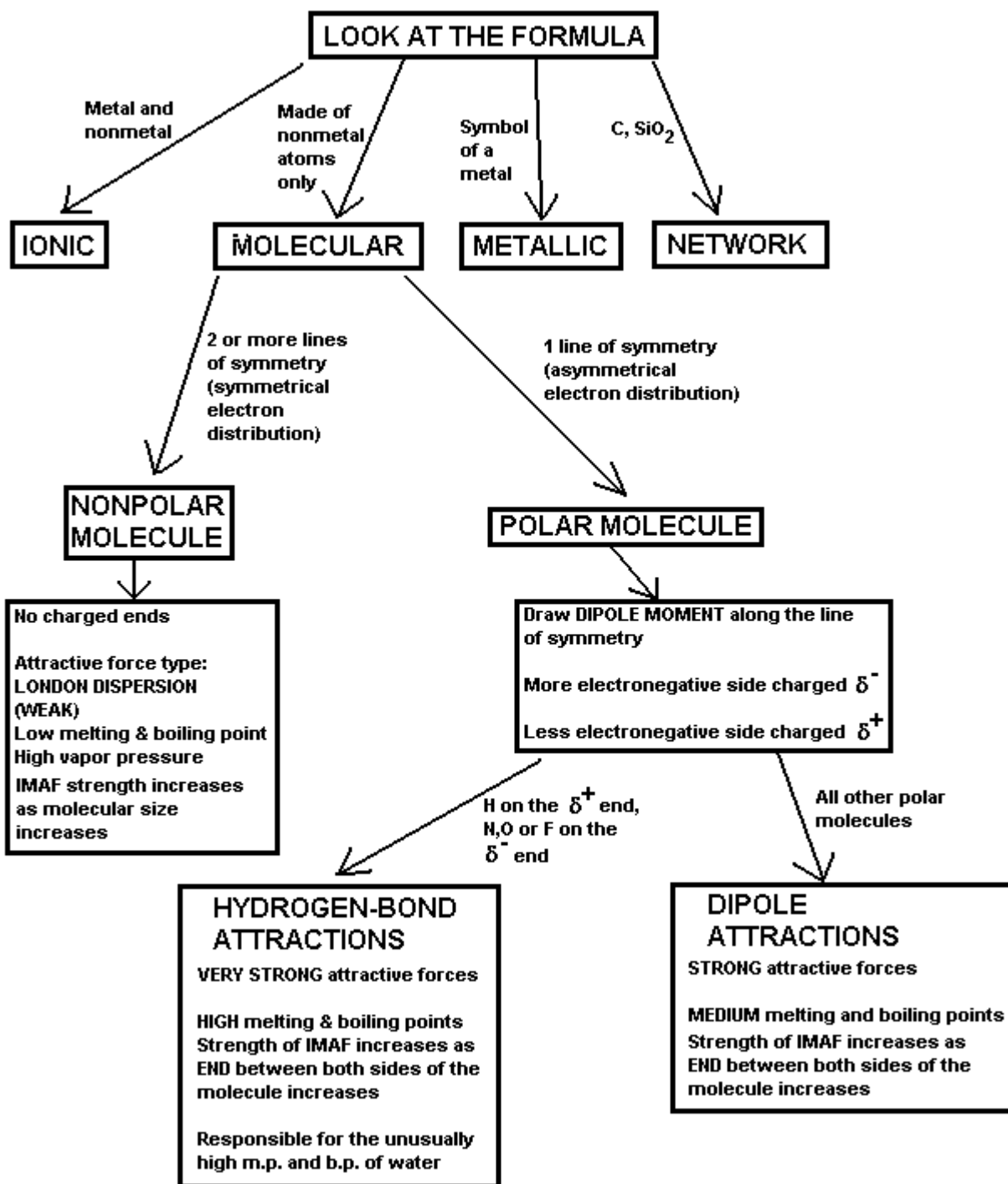
Other atoms that hydrogen can form a hydrogen bond with are the other nonmetal atoms that also have a high electronegativity and very small atomic radius; nitrogen (3.0) and fluorine (4.0). Even though chlorine has a higher electronegativity (3.2) than nitrogen does, chlorine does not form hydrogen bonds. 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. 23-26)

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, Li₂O 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 → CaCl₂
X ⁺³ and Y ⁻¹	It takes three –1 charges to cancel out the +3 charge	XY ₃	Fe ⁺³ and Br ⁻¹ → FeBr₃
X ⁺¹ and Y ⁻²	It takes two +1 charges to cancel out the -2 charge.	X ₂ Y	Li ⁺¹ and O ⁻² → Li₂O
X ⁺¹ and Y ⁻³	It takes three +1 charges to cancel out the -3 charge	X ₃ Y	Na ⁺¹ and P ⁻³ → Na₃P
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 ₃ Y ₂	Cu ⁺² and N ⁻³ → Cu₃N₂
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 ₂ Y ₃	Cr ⁺³ and S ⁻² → Cr₂S₃
X ⁺⁴ and Y ⁻²	It takes two –2 charges to cancel out one +4 charge.	XY ₂	Pb ⁺⁴ and O ⁻² → PbO₂

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. (PbO₂)

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:

39.0983 K 19 2-8-8-1 +1	65.39 Zn 30 2-8-18-2 +2	55.847 Fe 26 2-8-14-2 +2 +3	107.868 Ag 47 2-8-18-18-1 +1	51.996 Cr 24 2-8-13-1 +2 +3 +6	40.08 Ca 20 2-8-8-2 +2	63.546 Cu 29 2-8-18-1 +1 +2
15.9994 O 8 2-6 -2	30.97376 P 15 2-8-5 -3 +3 +5	32.06 S 16 2-8-6 -2 +4 +6	35.453 Cl 17 2-8-7 -1 +1 +3 +5 +7	79.904 Br 35 2-8-18-7 -1 +1 +5		

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 ₂ S	Ag only has one charge, so Ag^{+1} is called silver . S^{-2} is called sulfide .	Silver sulfide
Cr ₂ O 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 ₃ P 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 ₂ S	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)	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$\text{C}_2\text{H}_3\text{O}_2^-$ CH_3COO^-	acetate	NO_2^-	nitrite
CN^-	cyanide	NO_3^-	nitrate
CO_3^{2-}	carbonate	O_2^{2-}	peroxide
HCO_3^-	hydrogen carbonate	OH^-	hydroxide
$\text{C}_2\text{O}_4^{2-}$	oxalate	PO_4^{3-}	phosphate
ClO^-	hypochlorite	SCN^-	thiocyanate
ClO_2^-	chlorite	SO_3^{2-}	sulfite
ClO_3^-	chlorate	SO_4^{2-}	sulfate
ClO_4^-	perchlorate	HSO_4^-	hydrogen sulfate
		$\text{S}_2\text{O}_3^{2-}$	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
Ammonium sulfide	Ammonium is K^+ , sulfide is S^{2-}	$(\text{NH}_4)_2\text{S}$
Cobalt (II) cyanide	Cobalt (II) is Co^{+2} , cyanide is CN^{-1}	$\text{Co}(\text{CN})_2$
Cobalt (III) sulfate	Cobalt (III) is Co^{+3} , sulfate is SO_4^{2-}	$\text{Co}_2(\text{SO}_4)_3$
Tin (II) nitrate	Tin (II) is Sn^{+2} , nitrate is NO_3^{-1}	$\text{Sn}(\text{NO}_3)_2$
Tin (IV) nitrate	Tin (IV) is Sn^{+4} , nitrate is NO_3^{-1}	$\text{Sn}(\text{NO}_3)_4$
Calcium phosphate	Calcium is Ca^{+2} , phosphate is PO_4^{3-}	$\text{Ca}_3(\text{PO}_4)_2$
Aluminum dichromate	Aluminum is Al^{+3} , dichromate is $\text{Cr}_2\text{O}_7^{2-}$	$\text{Al}_2(\text{Cr}_2\text{O}_7)_3$
Silver nitrite	Silver is Ag^{+1} , nitrite is NO_2^{-3}	AgNO_2

39.0983 K 19 2-8-8-1	+1	65.39 Zn 30 2-8-18-2	+2
55.847 Fe 26 2-8-14-2	+2 +3	107.868 Ag 47 2-8-18-18-1	+1
51.996 Cr 24 2-8-13-1	+2 +3 +6	40.08 Ca 20 2-8-8-2	+2
63.546 Cu 29 2-8-18-1			

Table E
Selected Polyatomic Ions

H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	dimercury (I)	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$\text{C}_2\text{H}_3\text{O}_2^-$ CH_3COO^-	acetate	NO_2^-	nitrite
CN^-	cyanide	NO_3^-	nitrate
CO_3^{2-}	carbonate	O_2^{2-}	peroxide
HCO_3^-	hydrogen carbonate	OH^-	hydroxide
$\text{C}_2\text{O}_4^{2-}$	oxalate	PO_4^{3-}	phosphate
ClO^-	hypochlorite	SCN^-	thiocyanate
ClO_2^-	chlorite	SO_3^{2-}	sulfite
ClO_3^-	chlorate	SO_4^{2-}	sulfate
ClO_4^-	perchlorate	HSO_4^-	hydrogen sulfate
		$\text{S}_2\text{O}_3^{2-}$	thiosulfate

NAMING TERNARY COMPOUNDS GIVEN THE FORMULAS

Formula	How to determine name	Name
KNO_3	K only has one charge, so K^+ is called potassium . NO_3^- 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
$\text{Fe}(\text{NO}_2)_2$	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^- is called nitrite .	Iron (II) nitrite
$\text{Fe}(\text{NO}_2)_3$	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^- is called bromide .	Iron (III) nitrite
Ag_2CO_3	Ag only has one charge, so Ag^+ is called silver . CO_3^{2-} is called carbonate .	Silver carbonate
$\text{Cr}_2(\text{SO}_3)_3$	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
$\text{Cr}(\text{SO}_3)_3$	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
$\text{Ca}_3(\text{PO}_4)_2$	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 ₂ O ₅	Nitrogen (V) oxide	Phosphorous (III) oxide	P ₂ O 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 ₂ O ₄
N ₂ O ₅	Dinitrogen pentoxide	diphosphorous trioxide	P ₂ O 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. 27, 28)

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.

$\text{Ca}(\text{NO}_3)_2$ 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	$2+1+4 = 7$
$\text{Al}(\text{NO}_3)_3$	1 Al, 3 N, 9 O	1 Al, 3 N, 9 O	$1+3+9 = 13$
$\text{Al}_2(\text{SO}_4)_3$	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	$2 \times (1 \text{ K}, 1 \text{ Br}) = 2 \text{ K}, 2 \text{ Br}$
MgF_2	1 Mg, 2 F	3 MgF_2	$3 \times (1 \text{ Mg}, 2 \text{ F}) = 3 \text{ Mg}, 6 \text{ F}$
CaSO_4	1 Ca, 1 S, 4 O	4 CaSO_4	$4 \times (1 \text{ Ca}, 1 \text{ S}, 4 \text{ O}) = 4 \text{ Ca}, 4 \text{ S}, 16 \text{ O}$
$\text{Au}(\text{NO}_2)_3$	1 Au, 3 N, 6 O	5 $\text{Au}(\text{NO}_2)_3$	$5 \times (1 \text{ Au}, 3 \text{ N}, 6 \text{ O}) = 5 \text{ Au}, 15 \text{ N}, 30 \text{ O}$
$\text{Co}(\text{OH})_3$	1 Co, 3 O, 3 H	4 $\text{Co}(\text{OH})_3$	$4 \times (1 \text{ Co}, 3 \text{ O}, 3 \text{ H}) = 4 \text{ Co}, 12 \text{ O}, 12 \text{ H}$
$\text{Mg}_3(\text{PO}_4)_2$	3 Mg, 2 P, 8 O	3 $\text{Mg}_3(\text{PO}_4)_2$	$3 \times (3 \text{ Mg}, 2 \text{ P}, 8 \text{ O}) = 9 \text{ Mg}, 6 \text{ P}, 24 \text{ O}$

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 CH₄.
A molecule of benzene contains six atoms of carbon and six atoms of hydrogen, so the molecular formula is C₆H₆.

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. CaCl₂ 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, CaCl₂ has a 1:2 ratio and Al₂(SO₄)₃ has a 2:3:12 ratio, none of which can be simplified any further.

Molecular compounds like CH₄ (1:4), NO₂ (1:2) and H₂O (2:1) cannot be simplified any further. For them, the molecular formula is also the empirical formula. Others can be simplified. C₆H₆ (6:6) can be simplified to a 1:1 ratio, giving the empirical formula CH. H₂O₂ (2:2) can be simplified to 1:1, for an empirical formula of HO. C₄H₈ (4:8) can be simplified to a ratio of 1:2, giving the empirical formula of CH₂.

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:

39.0983 +1 K 19 2-8-8-1 39.1	40.08 +2 Ca 20 2-8-8-2 40.1	26.98154 +3 Al 13 2-8-3 27.0	14.0067 -3 N 7 2-5 14.0	15.9994 -2 O 8 2-6 16.0	35.453 -1 Cl 17 2-8-7 35.5	32.06 -2 S 16 2-8-6 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(NO₃)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**

Al₂(SO₄)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: $(\text{given g}) / (\text{g /mole}) = \text{mole}$

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 $\text{Ca}(\text{NO}_3)_2$ are there in 45.5 grams of $\text{Ca}(\text{NO}_3)_2$?

1) Find the formula mass of $\text{Ca}(\text{NO}_3)_2$: (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: $(\text{given moles} \times \text{g/mole}) = \text{g}$

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 $\text{Ca}(\text{NO}_3)_2$ will have what mass in grams?

1) Find the formula mass of $\text{Ca}(\text{NO}_3)_2$: (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 and Molecular Formulas (HW: p. 29-31)

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:

$(\text{given mass of element in sample} / \text{given mass of total sample}) \times 100\%$

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?

$$\% \text{O} = (56.0/80.0) \times 100 = \mathbf{70.0\%}$$

$$\% \text{N} = (24.0/80.0) \times 100 = \mathbf{30.0\%}$$

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

$(\text{formula mass of element in compound} / \text{formula mass of compound}) \times 100\%$

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

$$\% \text{Si} = (28.1/60.1) \times 100 = \mathbf{46.8\%}$$

$$\% \text{O} = (2 \times 16.0 = 32.0), (32.0/60.1) \times 100 = \mathbf{53.2\%}$$

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

$$\% \text{H} = (2 \times 1.0 = 2.0), (2.0/18.0) \times 100 = \mathbf{11\%}$$

$$\% \text{O} = (16.0/18.0) = \mathbf{88.9\%} \text{ (doesn't add up to 100\% because of sig figs)}$$

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

$$\% \text{H} = (2 \times 1.0 = 2.0), (2.0/98.1) \times 100 = \mathbf{2.0\%} \quad \% \text{S} = (32.1/98.1) \times 100 = \mathbf{32.7\%} \quad \% \text{O} = (4 \times 16.0 = 64.0), (64.0/98.1) \times 100 = \mathbf{65.2\%}$$

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 • $n \text{ H}_2\text{O}$

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.

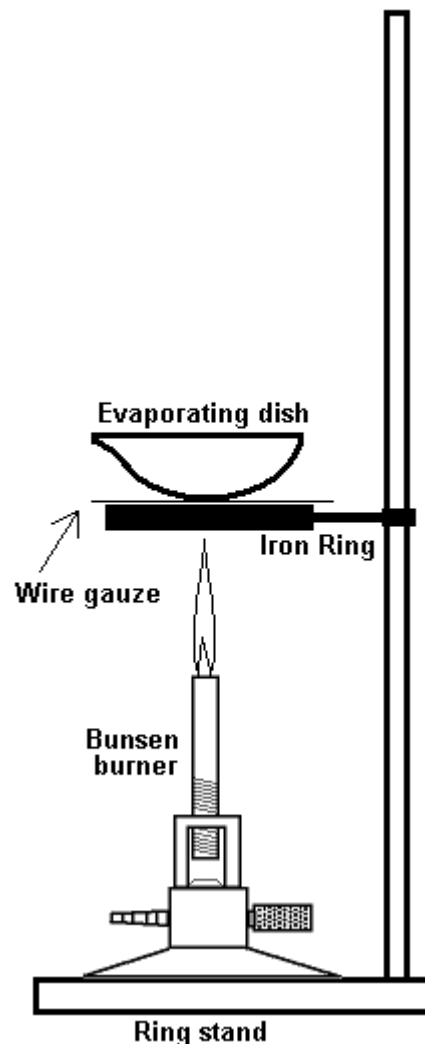
Crystals have open spaces between the atoms that make up the crystal. Just like food can get caught in the spaces between your teeth, water molecules can get trapped in the spaces between the ions in the crystal. Water molecules have charged ends, and they are attracted to the charged ions in the crystal. 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 =
 $(\text{mass of water driven off} / \text{mass of hydrate}) \times 100$.



Determining the % Of Water In A Hydrate Experimentally:

$$(\text{given mass of water in hydrate} / \text{given mass of hydrate}) \times 100\%$$

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:

$(\text{mass of hydrate} - \text{mass of anhydride}) = 10.40 \text{ grams} - 8.72 \text{ grams} = \mathbf{1.68 \text{ grams of water driven off by the heat}}$

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

$(\text{given mass of water in hydrate} / \text{given mass of hydrate}) \times 100 = (1.68 \text{ grams} / 10.40 \text{ grams}) \times 100 = 0.162 \times 100 =$

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?

$\text{Mass of hydrated crystal} = (\text{mass of crystal} + \text{dish}) - (\text{mass of dish}) = 39.486 \text{ g} - 29.993 \text{ g} = \mathbf{9.493 \text{ grams of hydrate}}$

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:

$39.486 \text{ g} - 38.387 \text{ g} = \mathbf{1.099 \text{ grams of water driven off by the heat}}$

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

$(\text{given mass of water in hydrate} / \text{given mass of hydrate}) \times 100 = (1.099 \text{ g} / 9.493 \text{ g}) \times 100 = 0.116 \times 100 =$

11.6% water by mass

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

$$(\text{formula mass of water in hydrate} / \text{formula mass of hydrate}) \times 100\%$$

Water (H₂O) 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 H₂O is 2, then the water weighs (2 X 18.0) = 36.0 grams in the hydrate.

If the coefficient of H₂O 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 CuCO₃ • 2 H₂O?

Neat little trick here. There are 2 molecules of water trapped in each formula unit of CuCO₃. 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 CuCO₃ and two more in the **2 H₂O**.

Why 4 hydrogens? Well, you have four of them in **2 H₂O**.

$$\% \text{Cu} = (63.5/159.5) \times 100 = \mathbf{39.8\%} \quad \% \text{C} = (12.0/159.5) \times 100 = \mathbf{7.52\%}$$

$$\% \text{O} = (80.0/159.5) \times 100 = \mathbf{50.2\%} \quad \% \text{H} = (4.0/159.5) \times 100 = \mathbf{2.5\%}$$

What is the % composition, by mass, of H₂O in CuCO₃ • 2 H₂O?

The formula mass of the hydrate is 159.5 g/mole. There are 2 H₂O's, each one weighs (2 H X 1.0 g/mole) + (1 O X 16.0 g/mole) = 18.0 g/mole. Therefore, % H₂O = [(2.0 X 18.0 = 36.0), (36.0) / (159.5) X 100 = **22.6%**

B) MOLECULAR FORMULAS – Empirical formulas tell you the smallest, whole-number ratio by moles of elements in a compound. Molecular formulas are simply whole-number multiples of those empirical formulas. For example, hydrogen peroxide has a molecular formula of H_2O_2 . The empirical formula is HO, a 1:1 ratio of hydrogen to oxygen. The molecular formula is twice the size of the empirical formula. This means that the mass of the molecular formula is twice that of the mass of the empirical formula. We will take advantage of this fact for what comes next!

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:

- Determine the formula mass of the empirical formula.
- 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
- 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.

$$\text{C}_2\text{H}_3 = (2 \text{ C} \times 12.0 \text{ g/mol}) + (3 \text{ H} \times 1.0 \text{ g/mol}) = \mathbf{27.0 \text{ g/mole}}$$

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

$$(54.0 \text{ g/mol}) / (27.0 \text{ g/mol}) = \mathbf{2.00}$$

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

$$2.00 \times \text{C}_2\text{H}_3 = \mathbf{\text{C}_4\text{H}_6}$$

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.

$$\text{CH} = (1 \text{ C} \times 12.0 \text{ g/mol}) + 1 \text{ H} \times 1.0 \text{ g/mol} = \mathbf{13.0 \text{ g/mole}}$$

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

$$(52.0 \text{ g/mol}) / (13.0 \text{ g/mol}) = \mathbf{4.00}$$

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

$$4.00 \times \text{CH} = \mathbf{\text{C}_4\text{H}_4}$$

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) CO₂ c) K₂O 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) CO₂ c) CH₄ d) SiO₂

Explain how you can tell: _____

Explain why one of the wrong choices is incorrect:

Choice: _____ Why: _____

3) Carbon dioxide (CO₂) has a boiling point of -57°C and water (H₂O) has a boiling point of 100°C. A molecule of CO₂ has a mass of 44.0 amu and a molecule of H₂O has a mass of 18.0 amu. Why then, if CO₂ is a larger molecule than H₂O, does H₂O 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.

Molecule (do B in this box)	Polar or Nonpolar	Shape	IMAF Type	Dot Diagram
H – Cl				
H – H				
$\begin{array}{c} \text{H} - \text{O} \\ \\ \text{H} \end{array}$				
$\begin{array}{c} \text{H} - \text{N} - \text{H} \\ \\ \text{H} \end{array}$				
$\begin{array}{c} \text{Cl} \\ \\ \text{Cl} - \text{C} - \text{Cl} \\ \\ \text{Cl} \end{array}$				
O = C = O				
H - Br				

2) Naming and Writing Binary and Ternary Formulas Homework

1) Binary Compounds:

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 ₂ O ₃	
Cu ₂ S	
CuS	
Na ₂ O	
FeO	
Li ₂ S	
Fe ₂ O ₃	
SrCl ₂	
PbO ₂	
MnO ₂	
AlP	
K ₃ P	
AuP	
ZnO	

2) Ternary Compounds:

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 ₂ O		
H ₂ SO ₄		
Fe ₃ (PO ₄) ₂		
HBr		
CaSO ₄		
K ₂ O		
ZnSO ₃		
(NH ₄) ₂ SO ₄		
HCH ₃ COO		
PbCl ₂		
CuSO ₄		
Fe ₂ S 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 ₂ O		
	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 ₂ SO ₄	K:	S:	O:
Fe ₂ (CO ₃) ₃	Fe:	C:	O:
SO ₃	S:	O:	
CH ₃ COOH	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) ₂ SO ₄	
MgSO ₄		Pb ₃ (PO ₄) ₂		3 K ₂ Cr ₂ O 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 ₂ O ₄		
C ₂ H ₆			Ra(CN) ₂		
Ba(NO ₃) ₂			C ₆ H 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 ₂ SO ₄		
H ₂ SO ₄		
AlF ₃		
ZnSO ₄		
H ₂ O		
HNO ₃		
Mg(OH) ₂		
CH ₄ S		

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 ₂ O		
29.25 grams of NaCl		
60. grams of NaOH		
71 grams of Na ₂ SO 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 ₄ S		
4.00 moles of H ₂ SO 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 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) Na₂O

%Na = _____ %O = _____

3) KClO₃

%K = _____ %Cl = _____ %O = _____

4) Na₂CO₃

%Na = _____ %C = _____ %O = _____

5) (NH₄)₂SO₃

%N = _____ %H = _____ %S = _____ %O = _____

6) $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$

%Cu = _____ %S = _____ %O = _____ %H = _____

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

1) $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$

Formula Mass of the _____ H_2O : _____ Formula Mass Of $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$: _____

Calculate the % Of Water in $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$:

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 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?