VALENCE and OXIDATION NUMBERS

Chemical reactions and the formation of compounds depend on the rearrangement of the electrons on the outer shells of the reacting atoms, so that each atom has a completed outer shell. The electrons in these outer shells are called **valence electrons.** The number of these valence electrons is important in the formation of compounds and their formula representation. When electrons are transferred from one atom to another in a chemical reaction, there will be a charge placed on each atom. The charge, which the atom has or appears to have after the electrons are transferred, is called the **oxidation number (oxidation state).** Formerly these positive and negative oxidation numbers were called **valence numbers.** **Valence** is now defined as the combining capacity of atoms; that is, the number of hydrogen atoms that can be combined with a given atom. Expressed differently **valence** represents the number of single bonds that a given atom can form. Valence therefore, is represented as a number without being positive or negative. Oxidation numbers can be positive or negative.

The following are rules to use in determining the oxidation numbers of elements in compounds. This information is essential in writing formulas.

1. Each atom in a free element has an oxidation number of zero. For example:

2. In ions containing one atom, the oxidation number is equal to the charge on that ion. For example:

3. All elements or metals in Group 1of the periodic table form only +1 ions and their oxidation number is +1 in all compounds. For example:

4. All elements or metals in Group 2 of the periodic table form only +2 ions and their oxidation number is +2 in all compounds. For example:

5. Oxygen has an oxidation number of -2 in most of its compounds with the exception of peroxides, such as H_2O_2, where the oxidation number is -1, and in compounds combined with fluorine, where it can be +1 or +2. For example:

6. Hydrogen has an oxidation number of +1 in most of its compounds with the exception of metal hydrides where its oxidation number is -1. For example:

7. The net sum of all the oxidation numbers in a given compound or neutral molecule must be zero. For example:

8. For polyatomic ions (ions consisting of more than one atom), the oxidation numbers of all the atoms must add up to the charge on the ion. Refer to table E. For example:

Note: Sometimes the oxidation numbers are easy to figure out, while sometimes it is more difficult. Some suggestions are as follows:

- Begin with assigning the oxidation numbers of elements you always know, such as group 1 or 2 metals, oxygen or hydrogen.
- Keep in mind that the total sum must equal zero.

For example:

FORMULAS AND OXIDATION NUMBERS

- The oxidation numbers of elements are used to write formulas of compounds.
- The numerical value of the total oxidation number of the positive part of the formula must be equal in charge to the negative part of the formula. (The sum of the oxidation numbers of the formula must be equal to zero)
- A simple method for writing formulas is to crisscross the oxidation number of each element and write these as subscripts.

For example:

a. Write the formula for a compound made from sodium and bromine.

b. Write the formula for a compound made from magnesium iodide.

c. Write the formula for a compound between calcium and oxygen.

d. Write the formula for a compound between aluminum and chlorine.

e. Write the formula for a compound between barium and the hydroxide ion.

f. Write the formula for a compound between potassium and sulfate.

g. Write the formula for a compound between iron (11) and the phosphate ion.

h. Write the formula for a compound between hydrogen and sulfur.

i. Write the formula for a compound between hydrogen and the sulfate ion.

j. Write the formula for a compound between hydrogen and the sulfite ion.

Name _____ Date _____ Class _____

CHAPTER 7 REVIEW ACTIVITY

Text Reference: Section 7-7

Empirical and Molecular Formulas

Write the empirical formula for each of the following molecular formulas.

1. N_2O_4 — 1. _____
2. NO_2 — 2. _____
3. C_2H_6 — 3. _____
4. CH_4 — 4. _____
5. $C_6H_{12}O_6$ — 5. _____
6. H_2SO_4 — 6. _____
7. $CaCl_2$ — 7. _____
8. C_2H_2 — 8. _____
9. $C_2H_4O_2$ — 9. _____
10. $Hg_2(NO_3)_2$ — 10. _____

Write two molecular formulas that reduce to each of the given empirical formulas. (Do not include the given empirical formula as an answer.)

11. HO — 11. _____
12. CS_2 — 12. _____
13. CH_3O_2 — 13. _____
14. Na_3N — 14. _____
15. $Al_2(SO_4)_3$ — 15. _____
16. $Pb(NO_3)_2$ — 16. _____
17. $Cr(CN)_3$ — 17. _____
18. AgCN — 18. _____
19. $(NH_4)_2S$ — 19. _____
20. NH_4NO_3 — 20. _____

Name _____ Date _____ Class _____

Formulas of Ionic Compounds

Write the correct formula for the compound formed by each of the following pairs of ions.

1. Na^+, F^- 1. _____
2. K^+, S^{2-} 2. _____
3. Ni^{2+}, SO_4^{2-} 3. _____
4. Al^{3+}, O^{2-} 4. _____
5. Ca^{2+}, ClO_3^- 5. _____
6. NH_4^+, P^{3-} 6. _____
7. Cu^+, NO_3^- 7. _____
8. Cu^{2+}, NO_3^- 8. _____

Write the symbols for the ions that form each of the following compounds.

9. $CaCl_2$ 9. _____
10. Na_2CO_3 10. _____
11. $Ga(ClO_3)_3$ 11. _____
12. CuF_2 12. _____
13. $(NH_4)_3PO_4$ 13. _____
14. $FeSO_4$ 14. _____
15. $Hg_2(NO_3)_2$ 15. _____
16. NH_4NO_2 16. _____

WRITING FORMULAS (CRISS-CROSS METHOD)

Name _____

Write the formulas of the compounds produced from the listed ions.

	Cl^-	CO_3^{-2}	OH^-	SO_4^{-2}	PO_4^{-3}	NO_3^-
Na^+						
NH_4^+						
K^+						
Ca^{+2}						
Mg^{+2}						
Zn^{+2}						
Fe^{+3}						
Al^{+3}						
Co^{+3}						
Fe^{+2}						
H^+						

CHEMICAL NOMENCLATURE: Naming chemical compounds

The chemical name of a compound usually indicates the chemical composition of the substance. Once again certain rules are followed in the naming and writing of chemical formulas. Chemical compounds can be divided into two main types:
(1) Binary compounds (compounds composed of only two elements) and
(2) Compounds containing more than two elements.

Binary compounds

1. In compounds composed of a metal and nonmetal the metallic element is written first (least electronegative) and the name of the nonmetal ends in –ide. Exceptions occur when dealing with certain polyatomic ions.
Try:

2. In compounds where the metal has more than one oxidation number, the metallic element is still written first followed by a roman numeral in parenthesis indicating the oxidation number of the metal and then the name of the nonmetal-ide. This method is called the stock system.
Try:

3. The stock system is also employed in naming compounds composed of two nonmetals. The nonmetal with the lower electronegativity is written first followed by the nonmetal with the higher electronegativity-ide. In many cases for simplicity prefixes (mono, di, tri, tetra, penta, hexa, and hepta) are also used to indicate how many atoms of each element are present. If you have only one of the first nonmetal the prefix mono is omitted.
Try:

4. In naming binary acids (compounds made from hydrogen and a nonmetal), the prefix hydro is attached to the front part of the negative ion and the suffix –ic is attached to the end of the negative ion. The word acid is then added.
Try:

Compounds containing more than two elements:

1. In compounds with one or more polyatomic ions, the name of the metal is named first (Remember to use a roman numeral if the metal has more than one oxidation number), followed by the name of the negative polyatomic ion. (Table E)
Things to note: Hg_2^{+2} is treated as mercury (I), and NH_4^{+1} is the ammonium ion. Both of these are positive and are written first.
OH^- the hydroxide ion and CN^-, the cyanide ion both end in –ide however their compounds must contain at least three elements. All the other ions are negative and end in ate or ite.
Try:

2. Hydrated salts (salts containing a specific amount of water are written with the Salt formula * #H_2O) are named according to the specified rules. However the number of attached water molecules is identified by a prefix on the word hydrate.
Try:

3. In naming ternary acids (Hydrogen followed by a negative polyatomic ion). Look up the name of the negative polyatomic ion on table E, if the suffix is –ate change it to –ic acid, and if the suffix -ite change it to –ous acid. Note the prefix hydro (used for naming binary acids) is not used.
Try:

M_xB_y

2 elements in the compound make the compound a **Binary** compound
The name of the compound ends in ***-ide***.
(N.B. #1)

Cation M is a Representative element in Group I, 2, or 13

Name of cation plus prefix of anion +*ide*.

eg. NaCl sodium chloride

eg. Al_2O_3 aluminum oxide

Cation M is a Representative element in Group 14(top) 15, 16, or 17 N.B.#2

Name of cation (with prefix indicating multiples) plus prefix of anion (with multiples indicated in same manner) + *ide*.

eg. N_2O_5 dinitrogen pentoxide

eg. SO_3 sulfur trioxide

Cation M is a Transition Element (Group 3-12)

Name of cation (use Roman Numerals to indicate Oxidation state) plus prefix of anion +*ide*.

eg. FeS Iron (II) sulfide

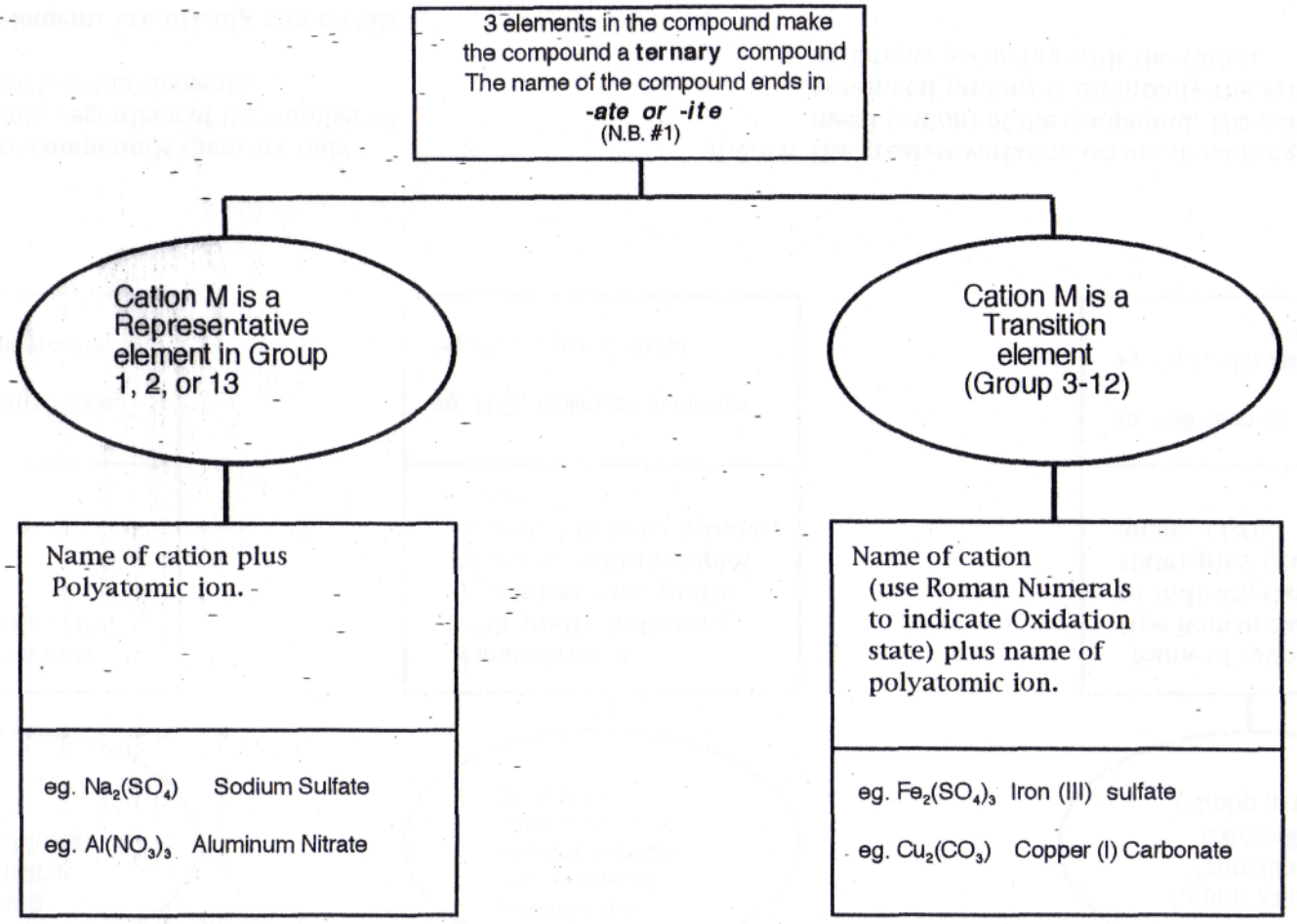
eg. Cr_2O_7 chromium (VII) oxide

N.B.#1: In binary compounds there are only 2 elements, regardless of the number of ions of each of the elements.

N.B. #2 : These elements are usually non-metals but are behaving as metals (cations) in these compounds. These are <u>molecular</u> compounds.

N.B.#3: The greatest variation occurs in naming the metal (cation) of the compound. The naming of the non-metal (anion) is essentially the same for all the variations occurring with the cation.

$M_x(B\ \ _z)_y$



N.B. #1: In ternary compounds there are only 3 elements, regardless of the number of ions of each of the elements,

N.B. #2: There is only one common polyatomic cation, NH_4^{1+} (Ammonium Ion)
Most of the polyatomic anions contain one or more oxygen atoms (the oxoanions)

NAMING IONIC COMPOUNDS

Name _____

Name the following compounds using the Stock Naming System.

1. $CaCO_3$ _____
2. KCl _____
3. $FeSO_4$ _____
4. $LiBr$ _____
5. $MgCl_2$ _____
6. $FeCl_3$ _____
7. $Zn_3(PO_4)_2$ _____
8. NH_4NO_3 _____
9. $Al(OH)_3$ _____
10. $CuC_2H_3O_2$ _____
11. $PbSO_3$ _____
12. $NaClO_3$ _____
13. CaC_2O_4 _____
14. Fe_2O_3 _____
15. $(NH_4)_3PO_4$ _____
16. $NaHSO_4$ _____
17. Hg_2Cl_2 _____
18. $Mg(NO_2)_2$ _____
19. $CuSO_4$ _____
20. $NaHCO_3$ _____
21. $NiBr_3$ _____
22. $Be(NO_3)_2$ _____
23. $ZnSO_4$ _____
24. $AuCl_3$ _____
25. $KMnO_4$ _____

NAMING MOLECULAR COMPOUNDS

Name _____

Name the following covalent compounds.

1. CO_2 _____
2. CO _____
3. SO_2 _____
4. SO_3 _____
5. N_2O _____
6. NO _____
7. N_2O_3 _____
8. NO_2 _____
9. N_2O_4 _____
10. N_2O_5 _____
11. PCl_3 _____
12. PCl_5 _____
13. NH_3 _____
14. SCl_6 _____
15. P_2O_5 _____
16. CCl_4 _____
17. SiO_2 _____
18. CS_2 _____
19. OF_2 _____
20. PBr_3 _____

NAMING ACIDS

Name _____

Name the following acids.

1. HNO_3 _____
2. HCl _____
3. H_2SO_4 _____
4. H_2SO_3 _____
5. $HC_2H_3O_2$ _____
6. HBr _____
7. HNO_2 _____
8. H_3PO_4 _____
9. H_2S _____
10. H_2CO_3 _____

Write the formulas of the following acids.

11. sulfuric acid _____
12. nitric acid _____
13. hydrochloric acid _____
14. acetic acid _____
15. hydrofluoric acid _____
16. phosphorous acid _____
17. carbonic acid _____
18. nitrous acid _____
19. phosphoric acid _____
20. hydrosulfuric acid _____

WRITING FORMULAS FROM NAMES

Name _____

Write the formulas of the following compounds.

1. ammonium phosphate _____
2. iron (II) oxide _____
3. iron (III) oxide _____
4. carbon monoxide _____
5. calcium chloride _____
6. potassium nitrate _____
7. magnesium hydroxide _____
8. aluminum sulfate _____
9. copper (II) sulfate _____
10. lead (IV) chromate _____
11. diphosphorus pentoxide _____
12. potassium permanganate _____
13. sodium hydrogen carbonate _____
14. zinc nitrate _____
15. aluminum sulfite _____

Name _____ Date _____ Class _____

CHAPTER 7 REVIEW ACTIVITY

Text Reference: Section 7-9

Names and Formulas of Compounds

Write either the traditional or Stock system name for each of the following compounds.

1. Na_2S 1. _____
2. NH_4Cl 2. _____
3. CuF 3. _____
4. CuF_2 4. _____
5. $PbSO_4$ 5. _____
6. $Hg(NO_3)_2$ 6. _____
7. Al_2O_3 7. _____
8. N_2O_4 8. _____
9. H_2S (acid name) 9. _____
10. $HClO_3$ (acid name) 10. _____

Write the formula for each of the following compounds.

11. nickel(II) chloride 11. _____
12. cuprous nitrate 12. _____
13. ammonium sulfate 13. _____
14. magnesium nitride 14. _____
15. mercury(I) sulfide 15. _____
16. carbon monoxide 16. _____
17. nitrogen(III) oxide 17. _____
18. diphosphorus pentoxide 18. _____
19. sulfurous acid 19. _____
20. periodic acid 20. _____

Name _____ Class _____ Date _____

Activity 4-3

Practice Drill: Formulas and Names

Write the name for each of the following compounds. Use the Stock system where appropriate.

1. $CaCO_3$ _____
2. FeO _____
3. H_2CO_3 _____
4. $AgCl$ _____
5. $Ca_3(PO_4)_2$ _____
6. $Ba(OH)_2$ _____
7. Na_2S _____
8. $FeCl_2$ _____
9. H_2CrO_4 _____
10. $(NH_4)_2SO_4$ _____
11. H_2SO_4 _____
12. $Zn(NO_3)_2$ _____
13. $CuSO_4$ _____
14. $AlCl_3$ _____
15. $NaOH$ _____
16. $PbCl_2$ _____
17. KNO_3 _____
18. $Mg(OH)_2$ _____
19. $HClO_3$ _____
20. $H_2C_2O_4$ _____

Write the chemical formula for each of the following compounds.

21. sodium nitrite _____
22. iron (III) oxide _____
23. aluminum hydroxide _____
24. ammonium hydroxide _____
25. magnesium chloride _____
26. hydrochloric acid _____
27. cuprous oxide _____
28. potassium sulfate _____
29. zinc oxide _____
30. barium sulfite _____
31. potassium carbonate _____
32. silver sulfide _____
33. nitrous acid _____
34. calcium phosphate _____
35. copper (II) nitrate _____
36. magnesium sulfide _____
37. aluminum oxide _____
38. barium nitride _____
39. lead (II) sulfate _____
40. hypochlorous acid _____

Write the name for each of the following compounds. Use the Stock system where appropriate.

41. NH_4NO_2 _____
42. $Ca(HCO_3)_2$ _____
43. $Ba(ClO_3)_2$ _____
44. Hg_2I_2 _____
45. KCN _____
46. PbO_2 _____
47. $KSCN$ _____
48. $Zn(C_2H_3O_2)_2$ _____
49. K_2SO_4 _____
50. $Hg(OH)_2$ _____
51. K_2SO_3 _____
52. Cu_2S _____
53. $KHSO_4$ _____
54. $ZnBr_2$ _____
55. $Fe_2(CrO_4)_3$ _____
56. $NaClO_4$ _____
57. $KClO$ _____
58. Mg_3N_2 _____
59. Na_2MnO_4 _____
60. $KMnO_4$ _____

rite the chemical formula for each of the following compounds.

61. mercury (I) cyanide _____
62. hydrosulfuric acid _____
63. iron (II) acetate _____
64. potassium chlorate _____
65. lead (II) fluoride _____
66. hydrobromic acid _____
67. ammonium oxalate _____
68. mercury (II) chromate _____
69. silver phosphate _____
70. potassium dichromate _____
71. barium nitride _____
72. sodium peroxide _____
73. cupric bromide _____
74. ammonium sulfide _____
75. calcium nitrate _____
76. zinc hydroxide _____
77. sodium hydrogen carbonate _____
78. lead (IV) oxide _____
79. potassium perchlorate _____
80. mercurous iodide _____

Lab: Formulas and oxidation numbers
Purpose: Use models of ions to assemble formulas.
Write formulas of chemical compounds.
Name chemical compounds.
Materials: Sheets of ion models
Scissors

Procedure:

(1). Cut up the ion model sheets into individual ions.

(2). Separate the ion models into two piles: positively charged ions and negatively charged ions.

(3). Note that the positive ion models have + charges on the right and the negative ions models have – charges on the left. This arrangement allows you to match the charges so that they are neutralized. Note also the +'s and –'s are lined up, the symbols are in the correct order to write chemical formulas.

(4). Using the models, assemble formulas for the compounds formed by each positive ion with each negative ion. Remember the positive ions must equal the negative ions.

(5). Count the number of ions of each charge in each formula and write the formula in data table 1.

(6). Write the name of each compound in data table 2.

(7) Double check your formulas by criss-crossing oxidation numbers to make sure that they are correct. Also check to make sure that polyatomic ions have parentheses around them if there is more than one in a formula.

Data and Observations: Fill in the data tables on the next page.

Questions:

1. What is a chemical formula? What information does a chemical formula provide?
2. Some compounds are described as "binary compound." What does this mean? List the negative ions (from your data table) that are making up binary compounds. What do there names have in compound.
3. Which element from your data table can have two different oxidation numbers? How do you indicate the oxidation number when writing the name of the compound made from this ion?
4. Parentheses must be used to show more than one polyatomic ion. List the negative ions (from your data table) that must be written in parenthesis. What do there names have in compound?
5. Are the formulas you wrote in your data table molecular formulas or empirical formulas? Explain.

Summary: What is an ionic compound? Make a list of steps you can use to name ionic compounds.

Name:_____ Lab # _____

Data and Observations:

Table1: Formulas

	F^-	Cl^-	Br^-	$C_2H_3O_2^-$	SO_4^{-2}	CO_3^{-2}	NO_3^-	PO_4^{-3}	$Cr_2O_7^{-2}$
NH_4^+									
K^+									
Li^+									
Na^+									
Ag^+									
Ca^{+2}									
Mg^{+2}									
Ba^{+2}									
Fe^{+2}									
Al^{+3}									

Data Table 2: Name the compounds from table 1

	F^-	Cl^-	Br^-	$C_2H_3O_2^-$	SO_4^{-2}	CO_3^{-2}	NO_3^-	PO_4^{-3}	$Cr_2O_7^{-2}$
NH_4^+									
K^+									
Li^+									
Na^+									
Ag^+									
Ca^{+2}									
Mg^{+2}									
Ba^{+2}									
Fe^{+2}									
Al^{+3}									

Ion Models — p. 1

Ba^{2+} + +	− − − PO_4^{3-}	Al^{3+}
Fe^{2+} + +	− F^-	− F^-
Ca^{2+} + +	− F^-	− F^-
	− Cl^-	− Br^-
− − $Cr_2O_7^{2-}$	− $C_2H_3O_2^-$	− − CO_3^{2-}
	− $C_2H_3O_2^-$	
− − $Cr_2O_7^{2-}$	Ag^+ +	Ag^+ +
	Ag^+ +	− NO_3^-
Ca^{2+} + +	− − − PO_4^{3-}	Al^{3+} + + +
Mg^{2+} + +	− NO_3^-	− NO_3^-
Mg^{2+} + +	− − SO_4^{2-}	− − CO_3^{2-}

Ion Models — p. 2

+ Ca^{2+} +	− SO_4^{2-} −	− CO_3^{2-} −
+ Ba^{2+} +	NH_4^+ +	NH_4^+ +
	Na^+ +	Na^+ +
+ Ba^{2+} +	K^+ +	K^+ +
	Li^+ +	Li^+ +
+ Mg^{2+} +	− F^-	− F^-
	Li^+ +	Li^+ +
+ Fe^{2+} +	− NO_3^-	− $C_2H_3O_2^-$
	− NO_3^-	− $C_2H_3O_2^-$
+ Fe^{2+} +	− Br^-	− Cl^-
	− Br^-	− Br^-
− $Cr_2O_7^{2-}$ −	− Cl^-	− Cl^-
	Ag^+ +	Ag^+ +
− $Cr_2O_7^{2-}$ −	− SO_4^{2-} −	− SO_4^{2-} −

Ion Models — p. 3

Ca^{2+} + +	− SO_4^{2-} −	− CO_3^{2-} −
Ba^{2+} + +	− Cl^-	Na^+ +
	− Cl^-	Na^+ +
− $Cr_2O_7^{2-}$ −	− Br^-	K^+ +
	− Br^-	K^+ +
Mg^{2+} + +	− $C_2H_3O_2^-$	− $C_2H_3O_2^-$
	− $C_2H_3O_2^-$	− Cl^-
Fe^{2+} + +	− Br^-	Na^+ +
	Li^+ +	K^+ +
− − − PO_4^{3-}	+ + + Al^{3+}	NH_4^+ +
		NH_4^+ +
		NH_4^+ +
− NO_3^-	+ Al^{3+} + +	− − − PO_4^{3-}
− − CO_3^{2-}		

Topic: Mathematics of Chemistry: Problems involving Formulas

[illegible]1: Calculations Involving Gram Atomic Mass, Gram Formula Mass and the Mole.

DO NOW: What is a mole?

1. GRAM ATOMIC MASS (GAM):

EXAMPLES:

Element	no. of atoms	atomic mass units	moles of atoms	GAM	no. of atoms per mole
1. Zn					
2. Si					
3. Cu					
4 Pb					
Al					
6. S					

2. GRAM MOLECULAR MASS (GMM):
note: You use the term gram formula mass (GFM) if the compound is ionic.

EXAMPLES:

1. NaCl:

2. CuO:

3. $FeSO_4$:

4. $CuSO_4$:

5. H_2O:

6. $C_{12}H_{22}O_{11}$:

3. THE MOLE:

mole = or mole=

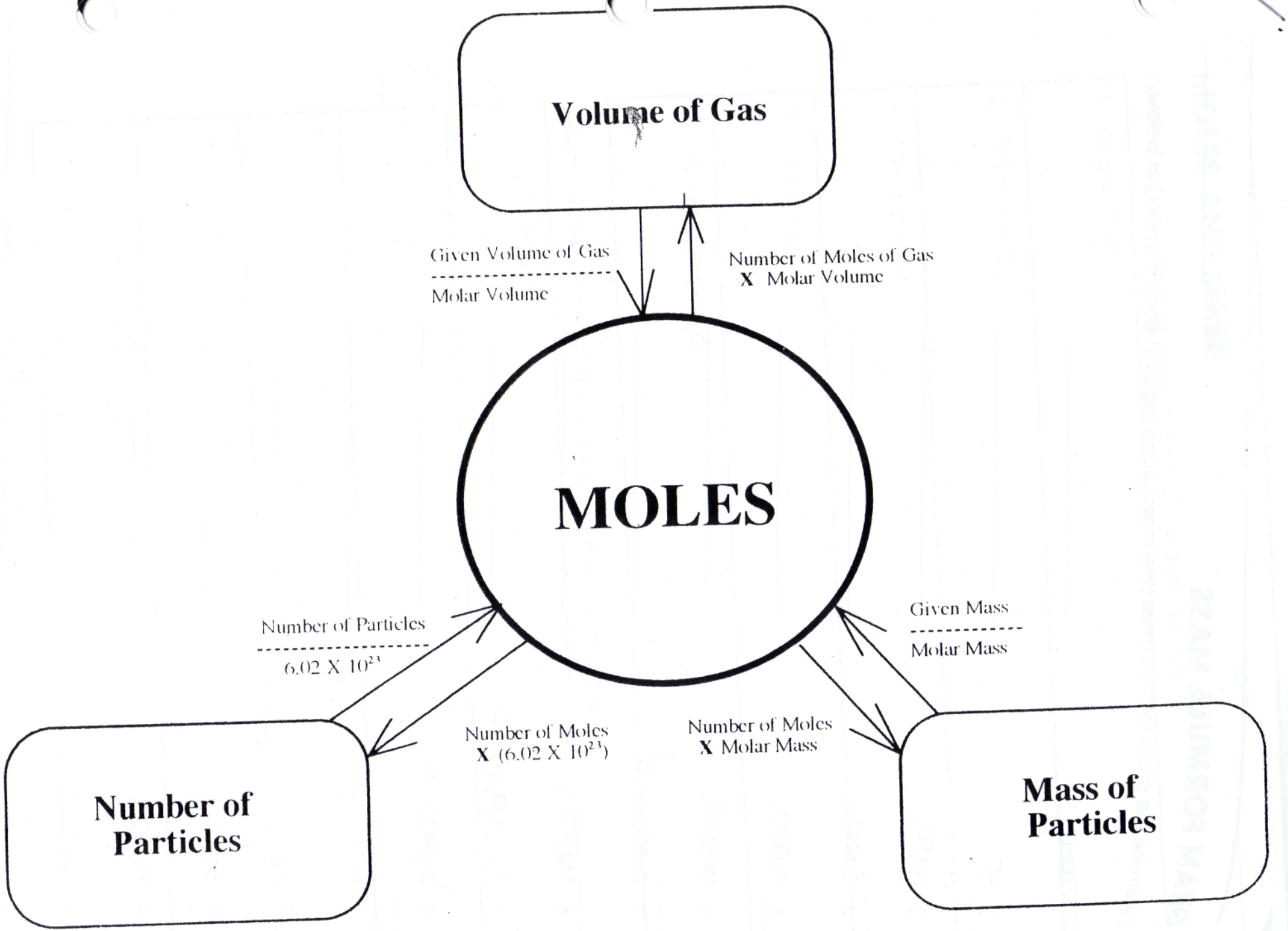
EXAMPLES:

1. 300 grams of Na_2SO_4 are to be used to make a solution. How many moles of Na_2SO_4 are needed?

2. What is the mass of 3 moles of $AlCl_3$?

3. How many molecules are contained in 14 grams of N_2?

How many moles are there in 45 liters of methane gas?


Volume of Gas
Given Volume of Gas
Molar Volume
Number of Moles of Gas
X Molar Volume
MOLES
Number of Particles
6.02 X 10^23
Number of Moles
X (6.02 X 10^23)
Number of Moles
X Molar Mass
Given Mass
Molar Mass
Number of Particles
Mass of Particles

GRAM FORMULA MASS

Name _____

Determine the gram formula mass (the mass of one mole) of each compound below.

1. $KMnO_4$ _____
2. KCl _____
3. Na_2SO_4 _____
4. $Ca(NO_3)_2$ _____
5. $Al_2(SO_4)_3$ _____
6. $(NH_4)_3PO_4$ _____
7. $CuSO_4 \bullet 5H_2O$ _____
8. $Mg_3(PO_4)_2$ _____
9. $Zn(C_2H_3O_2)_2 \bullet 2H_2O$ _____
10. $Zn_3(PO_4)_2 \bullet 4H_2O$ _____
11. H_2CO_3 _____
12. $Hg_2Cr_2O_7$ _____
13. $Ba(ClO_3)_2$ _____
14. $Fe_2(SO_3)_3$ _____
15. $NH_4C_2H_3O_2$ _____

MOLES AND MASS

Name _____

Determine the number of moles in each of the quantities below.

1. 25 g of NaCl _____
2. 125 g of H_2SO_4 _____
3. 100. g of $KMnO_4$ _____
4. 74 g of KCl _____
5. 35 g of $CuSO_4 \bullet 5H_2O$ _____

Determine the number of grams in each of the quantities below.

1. 2.5 moles of NaCl _____
2. 0.50 moles of H_2SO_4 _____
3. 1.70 moles of $KMnO_4$ _____
4. 0.25 moles of KCl _____
5. 3.2 moles of $CuSO_4 \bullet 5H_2O$ _____

THE MOLE AND VOLUME

Name _____

For gases at STP (273 K and 1 atm pressure), one mole occupies a volume of 22.4 L. What volume will the following quantities of gases occupy at STP?

1. 1.00 mole of H_2
2. 3.20 moles of O_2
3. 0.750 mole of N_2
4. 1.75 moles of CO_2
5. 0.50 mole of NH_3
6. 5.0 g of H_2
7. 100. g of O_2
8. 28.0 g of N_2
9. 60. g of CO_2
10. 10. g of NH_3

THE MOLE AND AVOGADRO'S NUMBER

Name _____

One mole of a substance contains Avogadro's Number (6.02×10^{23}) of molecules.

How many molecules are in the quantities below?

1. 2.0 moles
2. 1.5 moles
3. 0.75 mole
4. 15 moles
5. 0.35 mole

How many moles are in the number of molecules below?

1. 6.02×10^{23}
2. 1.204×10^{24}
3. 1.5×10^{20}
4. 3.4×10^{26}
5. 7.5×10^{19}

MIXED MOLE PROBLEMS

Name _____

Solve the following problems.

1. How many grams are there in 1.5×10^{25} molecules of CO_2?

_____

2. What volume would the CO_2 in Problem 1 occupy at STP?

_____

3. A sample of NH_3 gas occupies 75.0 liters at STP. How many molecules is this?

_____

4. What is the mass of the sample of NH_3 in Problem 3?

_____

5. How many atoms are there in 1.3×10^{22} molecules of NO_2?

_____

6. A 5.0 g sample of O_2 is in a container at STP. What volume is the container?

_____

7. How many molecules of O_2 are in the container in Problem 6? How many atoms of oxygen?

_____

_____

Name _____ Class _____ Date _____

Formula Mass

Calculate the formula mass of each of the following compounds, given its formula and the atomic mass of each element in the compound. Show your work.

1. Na_2S (Na = 23.0 u; S = 32.1 u) 1. _____

2. $Ba(NO_3)_2$ (Ba = 137.3 u; N = 14.0 u; O = 16.0 u) 2. _____

3. $(NH_4)_3P$ (N = 14.0 u; H = 1.0 u; P = 31.0 u) 3. _____

4. CH_3Cl (C = 12.0 u; H = 1.0 u; Cl = 35.5 u) 4. _____

5. SiF_4 (Si = 28.1 u; F = 19.0 u) 5. _____

6. Cu_2SO_4 (Cu = 63.5 u; S = 32.1 u; O = 16.0 u) 6. _____

7. $NaHCO_3$ (Na = 23.0 u; H = 1.0 u; C = 12.0 u; O = 16.0 u) 7. _____

8. H_2SO_3 (H = 1.0 u; S = 32.1 u; O = 16.0 u) 8. _____

Formula Mass (*Continued*)

For each of the molecular compounds given below, the empirical formula and the molecular mass in atomic mass units have been determined. Write the correct molecular formula for each compound. Show your work.

Example

Empirical Formula	Molecular Mass	Atomic Masses	Molecular Formula
HO	34.0 u	H = 1.0 u; O = 16.0 u	H_2O_2

The mass of each formula unit (HO) is 1.0 u + 16.0 u = 17.0 u. The molecular mass if 34.0 u.

$$\frac{34.0\ \text{u}}{17.0\ \text{u}} = 2.0$$

There are 2 formula units in each molecule. The molecular formula is therefore 2 × HO, or H_2O_2.

Empirical Formula	Molecular Mass	Atomic Masses	Molecular Formula
9. CH	78.0 u	C = 12.0 u; H = 1.0 u	**9.** _____
10. SiC	40.1 u	Si = 28.1 u; C = 12.0 u	**10.** _____
11. NO_2	92.0 u	N = 14.0 u; O = 16.0 u	**11.** _____
12. CO_2	44.0 u	C = 12.0 u; O = 16.0 u	**12.** _____
13. CH_2O	60.0 u	C = 12.0 u; H = 1.0 u; O = 16.0 u	**13.** _____

Name _____ Class _____ Date _____

Stoichiometry of Chemical Formulas

Solve the following problems. Show your work.

1. What is the mass of 7.5 moles of sulfur dioxide (SO_2)? (S = 32.1 g/mole; O = 16.0 g/mole) **1.** _____

2. How many moles are there in 250.0 g of sodium phosphate (Na_3PO_4)? (Na = 23.0 g/mole; P = 31.0 g/mole; O = 16.0 g/mole) **2.** _____

3. What is the volume occupied by 4.2 moles of oxygen gas (O_2) at STP? **3.** _____

4. How many moles are there in 45.0 liters of methane gas (CH_4) measured at STP? **4.** _____

5. How many atoms are there in 3.00 moles of sodium (Na)? **5.** _____

6. How many moles are there in 15.5×10^{23} molecules of carbon dioxide (CO_2)? **6.** _____

Class_____ Date_____

Activity 4-10
Percentage Composition

Finding percentage composition

The percentage composition, by mass, of a chemical compound can be found from experimental evidence.

Sample Problem 1 From laboratory measurements, a sample of a pure compound is known to have a mass of 3.74 grams. Analysis of the sample shows 1.10 g calcium, 0.880 g sulfur, and 1.76 g oxygen. What is the percentage composition of this compound?

Solution

$$\frac{1.10 \text{ g calcium}}{3.74 \text{ g compound}} \times 100 = 29.4\% \text{ (by mass) calcium}$$

$$\frac{0.880 \text{ g sulfur}}{3.74 \text{ g compound}} \times 100 = 23.5\% \text{ (by mass) sulfur}$$

$$\frac{1.76 \text{ g oxygen}}{3.74 \text{ g compound}} \times 100 = 47.1\% \text{ (by mass) oxygen}$$

Percentage composition may also be found by calculation from a known chemical formula.

Sample Problem 2 Calculate the percentage composition for copper (II) nitrate, $Cu(NO_3)_2$, from its formula.

Solution

1 Cu atom	1 × 63.5 u/atom	=	63.5 u
2 N atoms	2 × 14.0 u/atom	=	28.0 u
6 O atoms	6 × 16.0 u/atom	=	96.0 u
$Cu(NO_3)_2$	formula mass	=	187.5 u

Cu: $\frac{63.5 \text{ u}}{187.5 \text{ u}} \times 100 = 33.9\%$ copper by mass

N: $\frac{28.0 \text{ u}}{187.5 \text{ u}} \times 100 = 14.9\%$ nitrogen by mass

O: $\frac{96.0 \text{ u}}{187.5 \text{ u}} \times 100 = 51.2\%$ oxygen by mass

Practice problems

For each of the following, find the percentage composition from the analysis determined by experiment. Show a labeled setup in the space below each problem, and write your answers in the spaces provided.

1. A 14.80-g sample contains 3.83 g iron and 10.97 g bromine.

1. _____% Fe
_____% Br

2. A 9.14-g sample contains 4.77 g carbon, 1.19 g hydrogen, and 3.18 g oxygen.

2. _____% C
_____% H
_____% O

3. A 2.85-g sample contains 0.82 g magnesium, 0.41 g carbon, and 1.62 g oxygen.

3. _____% Mg
_____% C
_____% O

For each of the following, calculate the percentage composition from the formula of the compound. Show a labeled setup in the space below the problem, and write your answers in the spaces provided.

4. NaBr

4. _____% Na
_____% Br

5. H_2O_2

5. _____% H
_____% O

6. CaC_2O_4

6. _____% Ca
_____% C
_____% O

7. $Al_2(SO_4)_3$

7. _____% Al
_____% S
_____% O

PERCENTAGE COMPOSITION

Name _____

Determine the percentage composition of each of the compounds below.

1. $KMnO_4$

   K = _____

   Mn = _____

   O = _____

2. HCl

   H = _____

   Cl = _____

3. $Mg(NO_3)_2$

   Mg = _____

   N = _____

   O = _____

4. $(NH_4)_3PO_4$

   N = _____

   H = _____

   P = _____

   O = _____

5. $Al_2(SO_4)_3$

   Al = _____

   S = _____

   O = _____

Solve the following problems.

6. How many grams of oxygen can be produced from the decomposition of 100. g of $KClO_3$? _____

7. How much iron can be recovered from 25.0 g of Fe_2O_3? _____

8. How much silver can be produced from 125 g of Ag_2S? _____

COMPOSITION OF HYDRATES

Name _____

A hydrate is an ionic compound with water molecules loosely bonded to its crystal structure. The water is in a specific ratio to each formula unit of the salt. For example, the formula $CuSO_4 \bullet 5H_2O$ indicates that there are five water molecules for every one formula unit of $CuSO_4$. Answer the questions below.

1. What percentage of water is found in $CuSO_4 \bullet 5H_2O$?

_____

2. What percentage of water is found in $Na_2S \bullet 9H_2O$?

_____

3. A 5.0 g sample of a hydrate of $BaCl_2$ was heated, and only 4.3 g of the anhydrous salt remained. What percentage of water was in the hydrate?

_____

4. A 2.5 g sample of a hydrate of $Ca(NO_3)_2$ was heated, and only 1.7 g of the anhydrous salt remained. What percentage of water was in the hydrate?

_____

5. A 3.0 g sample of $Na_2CO_3 \bullet H_2O$ is heated to constant mass. How much anhydrous salt remains?

_____

6. A 5.0 g sample of $Cu(NO_3)_2 \bullet nH_2O$ is heated, and 3.9 g of the anhydrous salt remains. What is the value of n?

_____

Name _____ Class_____ Date_____

Activity 3-7
Empirical Formulas

The empirical formula of a compound expresses the simplest whole number ratio of elements in that compound. The empirical formula can be calculated from the percentage by mass for each element in the compound.

Sample Problem The percentage composition by mass of a compound is 56.6% potassium, 8.7% carbon, and 34.7% oxygen. Find its empirical formula.
Solution Assume that there are 100 grams of the compound. Then calculate the number of moles of atoms of each element in the sample.

$$100 \text{ g compound} \times \frac{56.6 \text{ g K}}{100 \text{ g compound}} \times \frac{1 \text{ mole K}}{39.1 \text{ g K}} = 1.45 \text{ moles K}$$

$$100 \text{ g compound} \times \frac{8.7 \text{ g C}}{100 \text{ g compound}} \times \frac{1 \text{ mole C}}{12.0 \text{ g C}} = 0.73 \text{ mole C}$$

$$100 \text{ g compound} \times \frac{34.7 \text{ g O}}{100 \text{ g compound}} \times \frac{1 \text{ mole O}}{16.0 \text{ g O}} = 2.17 \text{ moles O}$$

Use the number of moles as subscripts for each element to express the ratio of atoms:

$$K_{1.45}C_{0.73}O_{2.17}$$

Convert these values to a simple whole-number ratio by dividing each subscript by the smallest subscript.

$$\frac{K_{1.45}}{0.73}\frac{C_{0.73}}{0.73}\frac{O_{2.17}}{0.73} = K_2C_1O_3 \text{ or } K_2CO_3$$

If the results of this step are half-integers, multiply by 2 to convert to whole numbers. In the same way, if the results are third-integers, multiply by 3.

Practice problems

From percentage composition information in each of the following, calculate empirical formulas. Show a labeled setup below the problem, and write your answers in the spaces provided.

1. 69.6% barium, 6.1% carbon, 24.3% oxygen 1. _____

2. 40.5% zinc, 19.9% sulfur, 39.6% oxygen 2. _____

3. 88.8% copper, 11.2% oxygen 3. _____

4. 79.9% copper, 20.1% oxygen 4. _____

5. 36.7% potassium, 33.3% chlorine, 30.0% oxygen 5. _____

6. 28.2% potassium, 25.6% chlorine, 46.2% oxygen 6. _____

7. 40.2% potassium, 26.8% chromium, 33.0% oxygen 7. _____

8. 26.6% potassium, 35.3% chromium, 38.1% oxygen 8. _____

9. 56.3% oxygen, 43.7% phosphorus 9. _____

10. 90.7% lead, 9.33% oxygen 10. _____

DETERMINING EMPIRICAL FORMULAS

Name _____

What is the empirical formula (lowest whole number ratio) of the compounds below?

1. 75% carbon, 25% hydrogen

_____

2. 52.7% potassium, 47.3% chlorine

_____

3. 22.1% aluminum, 25.4% phosphorus, 52.5% oxygen

_____

4. 13% magnesium, 87% bromine

_____

5. 32.4% sodium, 22.5% sulfur, 45.1% oxygen

_____

6. 25.3% copper, 12.9% sulfur, 25.7% oxygen, 36.1% water

_____

DETERMINING MOLECULAR FORMULAS (TRUE FORMULAS)

Name _____

Solve the problems below.

1. The empirical formula of a compound is NO_2. Its molecular mass is 92 g/mol. What is its molecular formula?

_____

2. The empirical formula of a compound is CH_2. Its molecular mass is 70 g/mol. What is its molecular formula?

_____

3. A compound is found to be 40.0% carbon, 6.7% hydrogen and 53.5% oxygen. Its molecular mass is 60. g/mol. What is its molecular formula?

_____

4. A compound is 64.9% carbon, 13.5% hydrogen and 21.6% oxygen. Its molecular mass is 74 g/mol. What is its molecular formula?

_____

5. A compound is 54.5% carbon, 9.1% hydrogen and 36.4% oxygen. Its molecular mass is 88 g/mol. What is its molecular formula?

_____