Chapter 3: Molecules, Compounds and Chemical Equations:
(continue and finish chapter 3: 8-11)

Homework: Read Chapters 3. Work out sample/practice exercises.
Check MasteringChemistry.com assignment and complete before due date

Molar Mass:

The masses of molecules and compounds can be calculated from atomic masses found on the periodic table that are based on carbon-12 being exactly 12 amu/atom or 12 grams/mol. Molar mass may also be called... atomic mass or weight, molecular weight or mass, formula weight or mass.

Solve for the molar mass of CuSO₄

Avogadro’s Number: 6.022 x 10²³ parts/mole

6.022 x 10²³ amu = 1 gram, therefore we can take the weighted average mass of elements as units of either amu/atom or grams/mole, we will commonly use g/mol. With molar mass in combination with Avogadro’s number one can determine the number of formulas or atoms from grams.

Solve for the number of formula units in 5.00 g of CuSO₄, solve for the number of oxygen atoms.

Percent Composition of Compounds: Percent of each element by mass

Percent composition solves for the mass of the part divided by the mass of the total times 100%. Experiments can be performed to solve for the percent composition of each element in a compound. This information is useful when learning what a substance is composed of and useful in solving for empirical formulas.

Empirical Formula:

Empirical formulas are the simplest whole-number ratio of atoms in a compound and can be determined from elemental analysis. Often ionic compounds are already written as an empirical formula, but molecules like the organic compounds are written in a molecular formula. Several molecules may share the same empirical formula. The empirical formula C₂H₄O is shared by the following molecules: C₂H₄O, C₄H₈O₂, C₆H₁₂O₃, C₈H₁₆O₄, C₁₀H₂₀O₅ and so on.
Empirical formulas from combustion train:
Organic compounds are often tested through the use of combustion train analysis. A small organic sample is burned in a furnace that has oxygen added to insure complete combustion of all nonmetals. The flow of product gasses and the extra oxygen is then directed into a “train” during which the CO₂ gas is absorbed and weighed in one “train car” and H₂O vapor is absorbed and weighed in another “train car”. The mass information is gathered by the scientist and from that information an empirical formula may be determined.

The Chemical Equation:
Reactants → Products
Phases: solid (s), liquid (l), gas (g), aqueous (aq)

Balancing Chemical Equations:
Balance the same number of atoms on either side
Once the formulas are correctly written, add numbers to coefficients (never change a subscript to balance)

Process:
Write skeletal equation with correct formulas
Start with the substance with the most atoms and subscripts
Balance a free element last
Remove fractions by multiplying through by the denominator
Check
Practice Balancing Reactions:

1. \( \text{Na} \text{(s)} + \text{Br}_2 \text{(l)} \rightarrow \text{NaBr} \text{(s)} \)

2. \( \text{CH}_4 \text{(g)} + \text{O}_2 \text{(g)} \rightarrow \text{CO}_2 \text{(g)} + \text{H}_2\text{O} \text{(l)} \)

3. \( \text{FeCl}_3 \text{(aq)} + \text{AgNO}_3 \text{(aq)} \rightarrow \text{AgCl} \text{(s)} + \text{Fe(NO}_3)_3 \text{(aq)} \)

4. Solid potassium chlorate will decompose when heated into solid potassium chloride and oxygen gas.

5. Water vapor and sulfur dioxide exhaust from a coal burning plant may react to cause acid rain when sulfurous acid is formed.

Stoichiometry Example Problems:

1. Solve for the number of molecules or formula units, given moles:
   \[
   \text{Formula units} = \text{moles} \times \text{Avogadro's number}
   \]
   **Example 1:**
   How many formula units are in 3.20 moles of \( \text{Al}_2\text{S}_3 \)?
   \[
   (3.20 \text{ moles of Al}_2\text{S}_3) \times (6.022 \times 10^{23} \text{ formula units/mole})
   \]
   \[
   = 1.93 \times 10^{24} \text{ formula units of Al}_2\text{S}_3.
   \]

2. Solve for the number of atoms of a particular element in a given number of molecules or formula units:
   \[
   \text{atoms} = \text{molecules} \times \text{the number of atoms in each molecule}
   \]
   **Example 2:**
   How many Al atoms are in \( 1.93 \times 10^{24} \text{ formula units of Al}_2\text{S}_3 \)?
   \[
   (1.93 \times 10^{24} \text{ molecules Al}_2\text{S}_3) \times (2 \text{ Al atoms per 1 formula unit Al}_2\text{S}_3)
   \]
   \[
   = 3.86 \times 10^{24} \text{ atoms Al}
   \]
3. Find the average atomic mass of an element:
look up the number on the periodic table.

Example 3:
atomic mass of Na = 22.99 or atomic mass of Cl = 35.45
The units we use are grams/mole, rather than amu/atom.

4. Solve for the molar mass (also known as molecular weight or formula weight
for a molecule or compound, atomic mass or weight if its only one element)
given the formula:
MW = the summation of the number of atoms of each element times their
atomic mass. **Always keep a minimum of one place past the decimal.**

Example 4:
What is the molar mass of Al₂S₃?
2 Al = 2(26.98) = 53.96 g/mole
3 S = 3(32.06) = 96.18 g/mole
MW of Al₂S₃ = 150.14 g/mole

5. Solve for the mass of a compound given the number of moles:
mass = MW x moles (MW is molar weight)

Example 5:
What is the mass of 2.31 moles of Al₂S₃? (from example 4 we know the MW
is 150.14 g/mole)
150.14 g/mole x 2.31 moles = 347 grams

6. Solve for the number of moles of a compound given the mass:
moles = mass/MW

Example 6.
How many moles of Al₂S₃ are contained in 564 grams?
(MW still is 150.14 g/mol)
(564g)/(150.14g/mol) = 3.76 moles of Al₂S₃.
7. Coefficients The number before the formula or element in balanced equations: Coefficients represent the ratio of the number of molecules or moles involved, it does NOT represent grams.

Example 7:
\[2H_2 + O_2 \rightarrow 2 \text{H}_2\text{O}\]
this means it takes 2 hydrogen molecules or moles and 1 oxygen molecule or mole to produce 2 water molecules or moles.

8. Solve for a mass percent of an element in a given compound:
mass \% = (mass of element in 1 mole of compound / MW) \times 100\%

Example 8:
What is the mass percent of S in Al₂S₃?
The molar mass is 150.14g/mole; the mass of S is 3(32.06) = 96.18g/mole
\[
\left(\frac{96.18 \text{g/mol S}}{150.14 \text{g/mol Al}_2\text{S}_3}\right) \times 100\% = 64.06\%
\]

9. Empirical formulas are the simplest whole-number ratio of atoms in a compound.
To get the empirical formula divide the formula subscripts by the lowest common multiple.
Example 9:
C₆H₆ divide the subscript numbers by 6 = CH
C₈H₆O₄ divide the subscript numbers by 2 = C₄H₃O₂
10. Calculate the molecular formula when you know the empirical formula and the molecular weight use:
Molar mass / empirical formula mass = n
multiply subscripts in the empirical mass by n
Example 10:
A compound has an empirical formula of NO\textsubscript{2} and is found to have a molar mass of 92g. What is the molecular formula?
from method in example 4 the empirical formula mass = 46.0g
\[
\frac{92g \text{ total}}{46.0g \text{ NO}_{2}} = 2
\]
molecular formula = N\textsubscript{2}O\textsubscript{4}

11. Calculate empirical formulas when you know the mass of each element in the compound,
1. If you have % by mass; change the % sign to grams
   (64.1% = 64.1g)
2. Convert the mass of each element to moles (example 6)
3. Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all mole numbers are now integers, use these numbers as the subscripts in the empirical formula.
4. If the numbers are not within 1 or 2 decimals of a whole number, then multiply the numbers by the smallest integer necessary (such as 2, 3, 4, 5) to convert them to whole numbers. Use mole numbers as subscripts.

Example 11:
A compound was analyzed to have 11.98g Al and 21.35g of sulfur. What is the empirical formula?
1. We already have grams
2. Convert to moles using example 6 method (moles = mass/MW)
   atomic mass Al = 26.98g/mole; atomic mass of S = 32.06g/mole
   \[
   \frac{(11.98g \text{ Al})}{(26.98g/mol \text{ Al})} = 0.4440 \text{ mol Al}
   \]
   \[
   \frac{(21.35g \text{ S})}{(32.06g/mol \text{ S})} = 0.6659 \text{ mol S}
   \]
3. Divide each by 0.4440 mol (smallest number of moles)
   0.4440mol Al / 0.4440 mol = 1 mol Al
   0.6659 mol S / 0.4440 mol = 1.50 mol S
   Numbers are not all integers
4. Multiply each mole number by 2 to convert them to whole numbers
   1mole Al x 2 = 2 mol Al
   1.5 mol S x 2 = 3 mol S
   Empirical formula = Al\textsubscript{2}S\textsubscript{3}
12. Calculate empirical formulas using combustion train analysis.

1. The initial sample weight before combustion is known and the final mass of CO\textsubscript{2} and H\textsubscript{2}O products are known. All C has been converted to CO\textsubscript{2} and all H has been converted to H\textsubscript{2}O.

2. Convert the mass of CO\textsubscript{2} into moles of CO\textsubscript{2} and then to moles and mass of just the C alone.

3. Convert the mass of H\textsubscript{2}O into moles of H\textsubscript{2}O and then to moles and mass of just the H alone.

4. If the original sample contained only C, H, and O, the mass of O may be determined by subtracting the masses of C and H from the original mass of the sample. Convert the mass of oxygen to moles of oxygen.

5. Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all mole numbers are now integers, use these numbers as the subscripts in the empirical formula.

6. If the numbers are not within 1 or 2 decimals of a whole number, then multiply the numbers by the smallest integer necessary (such as 2, 3, 4, 5) to convert them to whole numbers. Use mole numbers as subscripts.

Example 12:

Glyceraldehyde is made up of only carbon, hydrogen, and oxygen. A 30.50 mg sample of glyceraldehyde was placed into a combustion train with extra oxygen flowing and yielded 18.30 mg H\textsubscript{2}O and 44.73 mg CO\textsubscript{2}. Its molecular weight is 90.0 g/mol. Solve for its empirical formula and its molecular formula.

1. Molar masses are generally written as g/mol, but it is sometimes useful to think in terms of the equivalent units mg/mmol.

2. Convert CO\textsubscript{2} to mmoles and mg of C
   atomic mass: C=12.01 mg/mmol; atomic mass of O=16.00 mg/mmol
   \[44.73 \text{ mg CO}_2 \left(\frac{1 \text{ mmol CO}_2}{44.01 \text{ mg CO}_2}\right) \left(\frac{1 \text{ mmol C}}{1 \text{ mmol CO}_2}\right) = 1.016 \text{ mmol C}\]
   \[1.016 \text{ mmol CO}_2 \left(\frac{12.01 \text{ mg C}}{1 \text{ mmol C}}\right) = 12.21 \text{ mg C}\]

3. Convert H\textsubscript{2}O to mmoles and mg of H
   atomic mass: H=1.008 mg/mmol; atomic mass of O=16.00 mg/mmol
   \[18.30 \text{ mg H}_2O \left(\frac{1 \text{ mmol H}_2O}{18.016 \text{ mg H}_2O}\right) \left(\frac{2 \text{ mmol H}}{1 \text{ mmol H}_2O}\right) = 2.032 \text{ mmol H}\]
   \[2.032 \text{ mmol H} \left(\frac{1.008 \text{ mg H}}{1 \text{ mmol H}}\right) = 2.048 \text{ mg H}\]
4. Solve for mass and moles of O

\[ 30.50 \text{ mg original sample} - (12.21 \text{ mg C} + 2.048 \text{ mg H}) = 16.24 \text{ mg O} \]

\[ 16.24 \text{ mg O} \left( \frac{1 \text{ mmol O}}{16.00 \text{ mg O}} \right) = 1.015 \text{ mmol O} \]

5. Divide all by smallest moles.

\[ \left( \frac{1.016 \text{ mmol C}}{1.015 \text{ mmol O}} \right) = 1 \text{ mmol C}, \left( \frac{2.032 \text{ mmol H}}{1.015 \text{ mmol O}} \right) = 2 \text{ mmol H}, 1 \text{ mmol O} \]

The empirical formula is \( \text{CH}_2\text{O} \).

6. The problem is not finished. Its molecular weight is 90.0 g/mol and we need to find the molecular formula.

\[ \left( \frac{90.0 \text{ mg MW}}{30.0 \text{ mg CH}_2\text{O}} \right) = 3, \text{ so the molecular formula = C}_3\text{H}_6\text{O}_3 \]

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**More Problems**

1. a) Calculate the formula weight of \( \text{Ca(ClO}_3)_2 \)
   b) Calculate the number of moles in 1.56 g of calcium chlorate
   c) Calculate the number of formula units/molecules in 1.56 g of \( \text{Ca(ClO}_3)_2 \)
   d) Calculate the number of oxygen atoms in 1.56 g of \( \text{Ca(ClO}_3)_2 \)
   e) Calculate the mass in grams of \( 3.011 \times 10^{22} \) formula units of \( \text{Ca(ClO}_3)_2 \)

2. Calculate the number of moles of \( \text{H}_3\text{PO}_4 \) found in 350 ml of dilute \( \text{H}_3\text{PO}_4 \) solution. The density of the solution is 1.24 g/ml and it is 65.1% \( \text{H}_3\text{PO}_4 \) by weight.
3. a) Balance the following unbalanced equation:
   \[ \text{BaO} (s) + \text{HCl} (aq) \rightarrow \text{BaCl}_2 (aq) + \text{H}_2\text{O} (l) \]
b) If there is 2 moles of BaO and 2 moles of HCl, which is the limiting reactant?

4. A compound consists of 2.00 g carbon, 0.333 g hydrogen and 2.66 g oxygen.
a) Calculate the number of moles carbon, hydrogen and oxygen in the sample.
b) Calculate the empirical formula of this compound.
c) If the molecular weight is found to be approximately 120, what is the molecular formula?
d) Calculate the percent composition of carbon, hydrogen and oxygen in this compound.

5. A sample of dimethlyhydrazine, a colorless liquid used as rocket fuel, is found to contain 39.9% carbon, 13.4% hydrogen, and 46.6% nitrogen.
a) Calculate the empirical formula of the sample.
b) If the molecular weight is found to be 60.0 g/mol, what is its molecular formula?