General Chemistry

202-SN1-RE

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Naming Hydrates

Hydrates of Salts

- Some salts exist with molecules of water bound to the crystal structure
- These are referred to as hydrated salts
- If no water is present, then the salt is described as anhydrous
- Depending on how hygroscopic (how readily it absorbs moisture) a compound may be difficult to find non-hydrated
- Sometimes this difference can be visually interesting

Hydrates of Salts – Cobalt(II) Chloride

- Each of these salts are cobalt(II) chloride

$CoCl_2$ anhydrous

$CoCl_2 \bullet 6H_2O$
(named cobalt(II) chloride hexahydrate)

Hydrates of Salts – Cobalt(II) Chloride

Remember the Prefixes!

Prefix	Number Indicated
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Naming Hydrates (Exercise)

- Provide the name for the following salts
  - $Fe(NO_3)_3 \bullet 9H_2O$
  - $CuSO_4 \bullet 5H_2O$

Naming Hydrates (Exercise)

- Provide the name for the following salts
  - $Fe(NO_3)_3 \bullet 9H_2O$: Iron(III) nitrate nonahydrate
  - $CuSO_4 \bullet 5H_2O$: Copper(II) sulfate pentahydrate

Nomenclature Summary

- Ionic Compounds = ionic bond
  - TRANSFER of electrons between a metal and nonmetal
  - ALWAYS empirical formula
  - “NEVER” prefixes
    - Most transition metals have specifications [ex. Iron (II) oxide]
    - Hydrates have prefixes [ex. *ionic compound* **penta**hydrate]

- Molecular Compounds = covalent bond
  - SHARING of electrons between nonmetal and nonmetal
  - ALWAYS prefixes
    - Exception: if the first element is 1, we don’t add mono (ex. CO is carbon monoxide and NOT monocarbon monoxide)
  - MOLECULAR formula

Exercise

Examine the following table of formulas and names. Which of the compounds are named correctly?

	Formula	***Name***
I	P_2O_5	Diphosphorus pentoxide
II	H_2SO_4	Hydrosulfuric acid
III	PbI_4	Lead iodide
IV	$CuSO_4$	Copper(I) sulfate

a) I, II
b) I, III, IV
c) I, IV
d) I only

Exercise

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Type II Binary Compounds (exercise)

What is the name of the compound CrO_2?

a) chromium oxide
b) chromium(II) oxide
c) chromium(IV) oxide
d) chromium dioxide

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Exercise

What is the name of the compound $SrBr_2$?

a) strontium bromine
b) sulfur bromide
c) strontium dibromide
d) strontium bromide

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a) strontium bromine
b) sulfur bromide
c) strontium dibromide
d) strontium bromide

Exercise

For both columns, indicate what type of compound each one is.

Provide the name of the following compounds:

- NaH
- Pb_3N_2
- K_2HPO_4
- H_2S
- $Ba(C_2H_3O_2)_2$

Give the formula of each compound:

Nitrous acid

Potassium carbonate

Dinitrogen pentoxide

Cobalt(III) nitrate

Exercise

Provide the name of the following compounds:

- NaH Sodium hydride
- Pb_3N_2 Lead (II) nitride
- K_2HPO_4 Potassium hydrogen phosphate
- H_2S Hydrosulfuric acid
- $Ba(C_2H_3O_2)_2$ Barium acetate

Writing Formulas from Names

Nitrous acid	HNO_2
Potassium carbonate	K_2CO_3
Dinitrogen pentoxide	N_2O_5
Cobalt(III) nitrate	$Co(NO_3)_3$

Three Fundamental Laws of Chemistry

1. **Law of Conservation of Mass**
2. **Law of Definite Proportions**
3. **Law of Multiple Proportions**

Fundamental Laws (video 10:58)

Chemical Composition

Counting by Weight

Counting by Weighing

- Objects do not need to have identical masses to be counted by weighing. All we need to know is the average mass of the objects.

Instead of counting each and every jelly bean, you can determine the # of jellybeans by weighing them.

Counting by Weighing

Averaging the Mass of Similar Objects

- Example: What is the mass of 1000 jelly beans?
  1. Not all jelly beans have the same mass.
  2. Suppose we weigh 10 jelly beans and find:

Bean	1	2	3	4	5	6	7	8	9	10
Mass	5.1 g	5.2 g	5.0 g	4.8 g	4.9 g	5.0 g	5.0 g	5.1 g	4.9 g	5.0 g

  3. Now we can find the average mass of a bean.

$$\text{Average mass} = \frac{\text{total mass of beans}}{\text{number of beans}}$$

$$= \frac{5.1\text{ g} + 5.2\text{ g} + 5.0\text{ g} + 4.8\text{ g} + 4.9\text{ g} + 5.0\text{ g} + 5.0\text{ g} + 5.1\text{ g} + 4.9\text{ g} + 5.0\text{ g}}{10}$$

$$= \frac{50.0}{10} = 5.0\text{ g}$$

← Average mass of 1 jelly bean

  4. Finally, we can multiply to find the mass of 1000 beans!

Mass of 1000 beans = 5.0 g x 1000 beans = 5000 g or 5 kg

Atomic Masses

Atomic Masses: Counting Atoms by Weighing

Atoms have very tiny masses, so scientists made a unit to avoid using very small numbers:

1 atomic mass unit (amu) = 1.66×10^{-24} g

- To count the atoms in a sample of a given element by weighing, we must know the mass of the sample and the average mass for that element.

- The average atomic mass for an element is the weighted average of the masses of all the isotopes of an element (the atomic mass).

Example: H = 1.008 amu, C = 12.011 amu, N = 14.007 amu

Atomic Masses

In nature, most elements contain mixtures of isotopes. The atomic mass is actually the ‘relative atomic mass’.

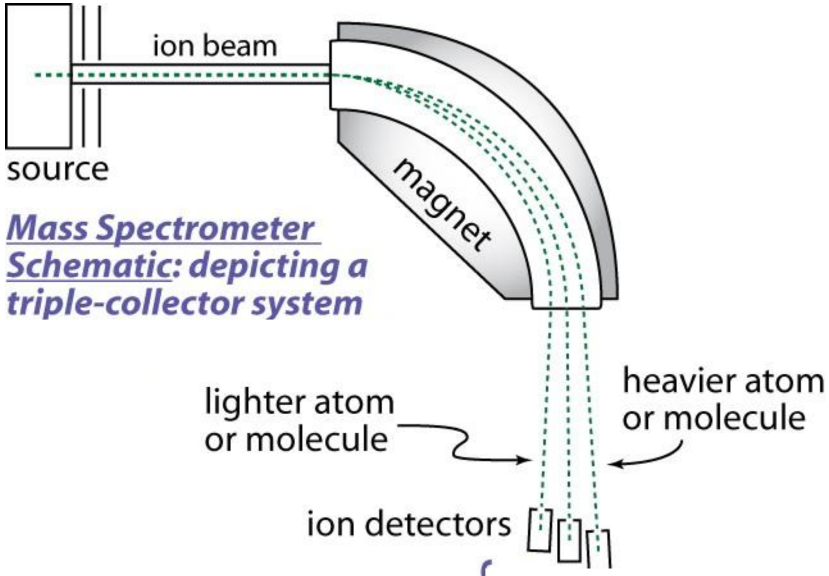
The relative atomic mass (***A_R***) of an element is the *weighted average* of the atomic masses of all isotopes, relative to the atomic mass unit (amu).

For the sample below, $A_R = \dfrac{(62 \times 12) + (1 \times 13) + (1 \times 14)}{64} = 12.04$ amu

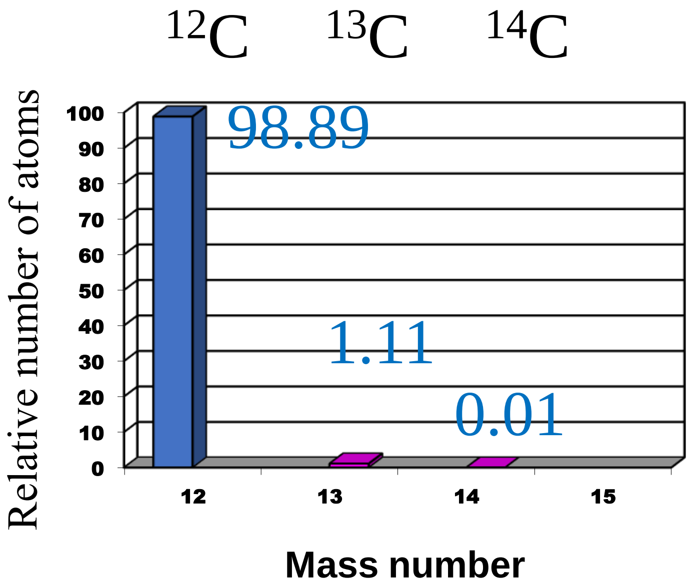
Actual A_R of Carbon is 12.011 amu.

Atomic Masses

- Relative Atomic Masses (A_R) can be measured in a mass spectrometer



C atomic masses { 12 13 14



^{12}C is defined as exactly 12 amu

http://serc.carleton.edu/research_education/geochemsheets/techniques/gassourcemassspec.html

Atomic Masses

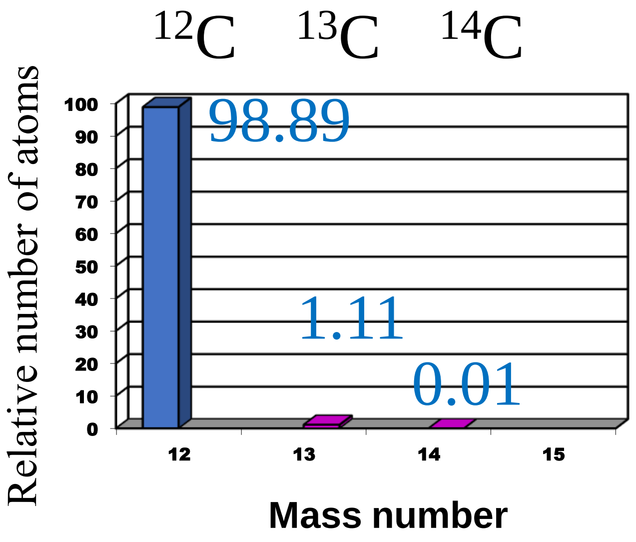
- Calculating the (average) atomic mass of an element.
- For carbon:

^{12}C ^{13}C ^{14}C

(98.88% of 12 amu) + (1.11% of 13.0001) + (0.01% of 14.0002) = 12.0113 amu

$$\left(\frac{98.88}{100} \times 12 \text{ amu}\right) + \left(\frac{1.11}{100} \times 13.0001\right) + \left(\frac{0.01}{100} \times 14.0002\right) = 12.0113 \text{ amu}$$

Note:
- No individual carbon atom has an A_R of 12.0113 amu!
- 12.0113 is an **average** value used for stoichiometric purposes!



The Mole

The Mole

- Because samples of matter typically contain so many atoms, the **mole** unit of measure was established for use in counting atoms.

- The **mole** is the number equal to the number of carbon atoms in exactly 12.011 grams of carbon.
  - Scientists chose carbon because it is regarded as the most significant element as all life is based on this atom.

- 1 mole of anything = 6.022×10^{23} units of that thing (Avogadro's number).
  - Just as a dozen eggs is 12 eggs, a mole of eggs is 6.022×10^{23} eggs! One mole of water contains 6.022×10^{23} H_2O molecules.

- 1 mole C = 6.022×10^{23} C atoms = 12.011 g C

The Mole: A Counting Unit for Chemistry

6.022×10^{23} atoms

How large is Avogadro's number?

1 mol of seconds represents a span of time 4 MILLION times as long as the earth has already existed!

1 mol of marbles is enough to cover the entire earth to a depth of 50 miles!

A computer counting at a rate of 1 billion atoms per second would take 19 million years to count a mole of atoms!

The Mole

- A sample of an element with a mass equal to that element's average atomic mass (expressed in g) contains one mole of atoms (6.022×10^{23} atoms).

- All samples below of pure elements contain the same number (a mole) of atoms: 6.022×10^{23} atoms.

Lead bar
207.2 g

Silver bar
107.9 g

Pile of copper
63.55 g

1 amu/atom
= 1 g/mol

Notice mass of each samples = atomic mass of the element!

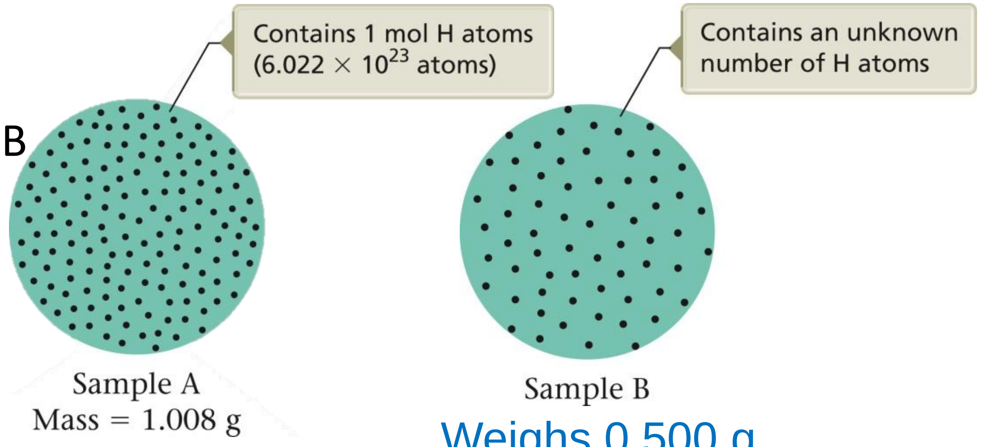
The Mole: Counting By Weighing

How do you find out how many atoms of H are in Sample B?

1. Determine mass of Sample B

2. Calculate number of moles in Sample B

3. Calculate the number of atoms



Weighs 0.500 g

1 mol H atoms = 1.008 g

mol H atoms in Sample B = 0.500 ~~g H~~ x $\frac{\text{1 mol H}}{\text{1.008 } \not{\text{g H}}}$ =

1 mol H atoms = 6.022×10^{23} atoms of H

H atoms in Sample B	= 0.496 ~~mol H~~ x $\frac{6.022 \times 10^{23} \text{ atoms of H}}{\text{1 } \not{\text{mol H}}}$ =

When in doubt, calculate the amount in moles and work from there.

The Mole (example)

A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68 mg. How many silicon (Si) atoms are present in this chip?

The average atomic mass for Si is 28.09 amu/atom.

Milligrams of Si atoms → Grams of Si atoms → Moles of Si atoms → Number of Si atoms

$$5.68 \text{ mg Si} \times \frac{1 \text{ g}}{1000 \text{ mg}} \rightarrow 5.68 \times 10^{-3} \text{ g Si}$$

$$5.68 \times 10^{-3} \text{ g Si} \times \frac{1 \text{ mol}}{28.09 \text{ g}} \rightarrow 2.02 \times 10^{-4} \text{ mol Si}$$

$$2.02 \times 10^{-4} \text{ mol Si} \times \frac{6.022 \times 10^{23} \text{ Si atoms}}{\text{mol}} \rightarrow 1.22 \times 10^{20} \text{ Si atoms}$$

Exercise

Calculate the number of iron atoms in a 4.48 mole sample of iron.

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Calculate the number of iron atoms in a 4.48 mole sample of iron.

$$\text{\# of Fe atoms} = 4.48\ \text{mol Fe} \times \frac{6.022 \times 10^{23}\ \text{Fe atoms}}{1\ \text{mol Fe}}$$

$$\text{\# of Fe atoms} = 2.70 \times 10^{24}\ \text{Fe atoms}$$

Exercise

Which silver ring would be worth more?

1. A ring that weighs 10.0 grams.

2. A ring that consists of 1.0×10^{22} atoms of Ag.

Exercise

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2. A ring that consists of 1.0×10^{22} atoms of Ag.

Always find your way to moles and then do conversion!

Choice #1 is the correct answer. The ring that weighs 10.0 g would be worth more. We can tell by using dimensional analysis to figure out how much the 1.0×10^{22} atoms of Ag would weigh:

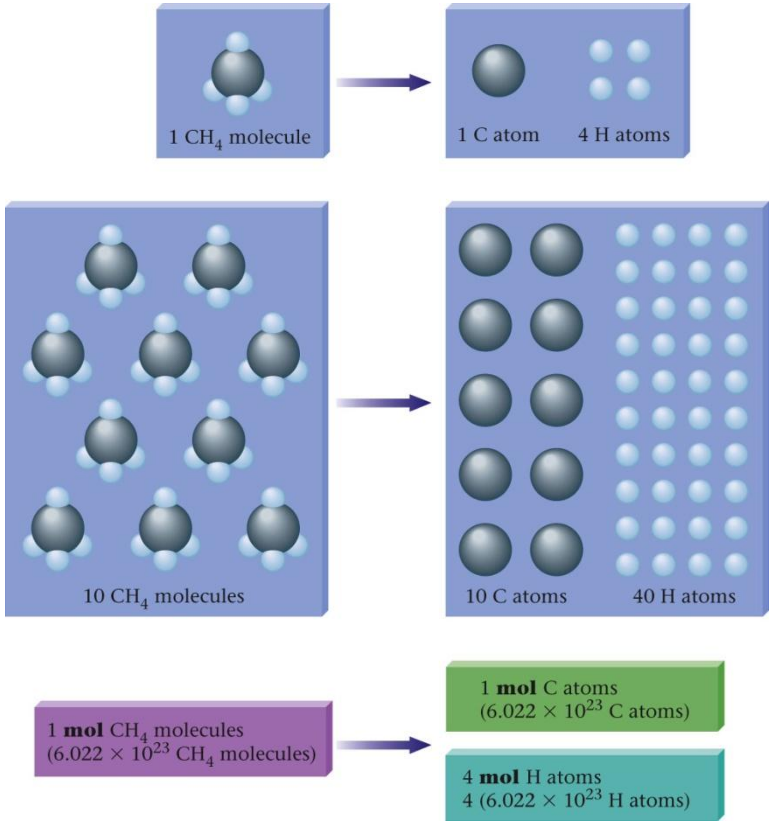
1.0×10^{22} atoms Ag x (1 mole Ag/6.022 $\times 10^{23}$ atoms Ag) x (107.868g Ag/1 mole Ag) = 1.8 g Ag

Molar Mass

Molar Mass

Molar mass (also referred to as the molecular weight) of any substance is the mass (in grams) of 1 mole of the substance.

Obtained from the sum of the individual atomic molar masses of the atoms that make up the compound.



For CH_4:

Mass of 1 mol C = 1 x 12.011 g = 12.011 g
Mass of 4 mol H = 4 x 1.008 g = 4.032 g
Mass of 1 mol CH_4 = 16.043 g

Molar mass for CH_4 is 16.043 g/mol

Unit is $\frac{g}{mol}$

Molar Mass

Examples:

Molar Mass of N = 14.007 g/mol

Molar Mass of H_2O = 18.015 g/mol

(2 × 1.008 g/mol) + 15.999 g/mol

Molar Mass of $Ba(NO_3)_2$ = 261.336 g/mol

137.328 g/mol + (2 × 14.007 g/mol) + (6 × 15.999 g/mol)

Molar Mass (example)

Calculate the molar mass for sodium sulfate.

1. Where are we going?

Determine molar mass of sodium sulfate in units of g/mol.

2. What do we know?

Formula for sodium sulfate is Na_2SO_4.

The atomic masses of each atom.

3. How do we get there?

1 Mole of Na_2SO_4 contains 2 moles Na+ ions and 1 mole SO_4^{2-} ion.

Mass of 2 mol Na+ = 2 x 22.990	=	45.980 g
Mass of 1 mol SO_4^{2-} = 32.066 + 4(15.999)	=	96.062 g
Mass of 1 mol Na_2SO_4	=	142.042 g

4. Reality Check

The molar mass of Na_2SO_4 is 142.042 g/mol which is greater than the atomic masses of the individual atoms. The units (g/mol) are correct, and the answer is reported to the correct number of significant figures (three decimal places).

Molar Mass (example – Continued)

A sample of sodium sulfate with a mass of 300.0 g represents what number of moles of sodium sulfate?

1. Where are we going?

Determine the number of moles of 300.0 g sodium sulfate.

2. What do we know?

The molar mass of Na_2SO_4 (142.042 g/mol).

We have 300.0 g of Na_2SO_4.

3. How do we get there?

Determine the number of moles by using the molar mass.

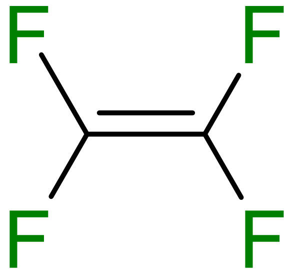
$$300.0 \text{ g } \cancel{Na_2SO_4} \times \frac{1 \text{ mol } Na_2SO_4}{142.042 \cancel{\text{ g } Na_2SO_4}} =$$

4. Reality Check

The number of moles of Na_2SO_4 is 2.112 mol. We know one mole of Na_2SO_4 is 142.042 g and since there was 300.0 g of Na_2SO_4, we expect at least twice the number of moles, which was what we see. The number of significant figures should be 4.

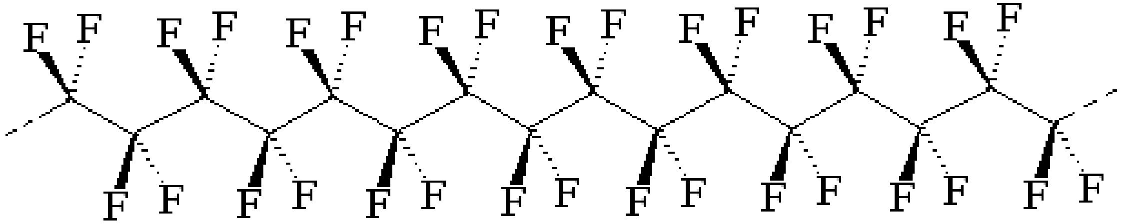
Calculating Number of Molecules

Example: The substance Teflon, the slippery coating on many frying pans, is made from the C_2F_4 molecule. Calculate the number of C_2F_4 units present in 135 g of Teflon.



Gas

If this is a gas, how can it be used to coat objects to be non-stick like this pan?



Polymerization!

Calculating Number of Molecules - Continued

Example: The substance Teflon, the slippery coating on many frying pans, is made from the C_2F_4 molecule. Calculate the number of C_2F_4 units present in 135 g of Teflon.

C_2F_4 : 2 x 12.011g = 24.022g
4 x 18.998g = 75.992g
24.022g + 75.992g = 100.014 g/mol

$$135 \text{ g } C_2F_4 \times \frac{1 \text{ mol } C_2F_4}{100.014 \text{ g } C_2F_4} = 1.35 \text{ mol } C_2F_4$$

Exercise

How many grams of fluorine are contained in one molecule of boron trifluoride?

a) 3.155×10^{-23} g
b) 9.465×10^{-23} g
c) 6.022×10^{23} g
d) 3.433×10^{25} g

Exercise

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b) 9.465×10^{-23} g
c) 6.022×10^{23} g
d) 3.433×10^{25} g

Exercise

You have 27.0 g of an unknown substance A and 75.0 g of chlorine gas. Substance A contains 1.5 times as many moles as the chlorine gas. What is the identity of A?

a) HClO
b) NO_2
c) BF_3
d) NH_3

Exercise

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a) HClO
b) NO_2
c) BF_3
d) NH_3

Choice #4 is the correct answer. To identify A, you need to know the molar mass (# grams/# moles). The number of grams of A is already given. To find the moles of A:

75.0 g Cl_2 x (1mol Cl_2/70.906 g Cl_2) x (1.5 mol A/1 mol Cl_2) = 1.6 mol A

Therefore, the molar mass of A is 17 g/mol (27.0 g/1.6 mol), which is NH_3.

Percent Composition

Percent Composition of Compounds

- A common way of describing the composition of a compound, based on the percentages (by mass) of each of its elements.

- Easiest to determine when 1 mole of the compound is considered.

- Mass percent of an element:

$$mass\ \% = \frac{mass\ of\ element\ in\ compound}{mass\ of\ compound} \times 100\%$$

Percent Composition of Compounds

- Mass percent of an element:

$$mass\ \% = \frac{mass\ of\ \ element\ in\ compound}{mass\ of\ \ compound} \times 100\%$$

- For example:
  - Mass percent of H in H_2O

$$\frac{\text{Mass H in 1 mole of } H_2O = \mathbf{2} \text{ x } 1.008 \text{ g}}{\text{Mass of 1 mole of } H_2O = 18.016 \text{ g}} \text{ X } 100\% = 11.19\%$$

  - Mass percent of O in H_2O

$$\frac{\text{Mass O in 1 mole of } H_2O = 1 \text{ x } 15.999 \text{ g}}{\text{Mass of 1 mole of } H_2O = 18.016 \text{ g}} \text{ X } 100\% = 88.80\%$$

Total % → **100%**

Percent Composition of Compounds

- Mass percent of an element:

$$mass\ \% = \frac{mass\ of\ \ element\ in\ compound}{mass\ of\ \ compound} \times 100\%$$

- For iron in iron(III) oxide:
  - Assume you take 1 mole

$$\text{Mass \% Fe} = \frac{2(55.845\text{ g})}{2(55.845\text{ g}) + 3(15.999\text{ g})} = \frac{111.690\text{ g}}{159.687\text{ g}} \text{ X } 100\% = 69.94\%$$

Percent Composition (Exercise)

Morphine, derived from opium plants, has the potential for use and abuse. Its formula is $C_{17}H_{19}NO_3$. What percent, by mass, is the carbon in this compound?

a) 12.0 %
b) 54.8 %
c) 67.9 %
d) 71.6 %

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a) 12.0 %
b) 54.8 %
c) 67.9 %
d) 71.6 %

Percent Composition (Exercise)

Consider separate 1 mole samples of each of the following:

N_2O, $C_3H_6O_2$, CO_2

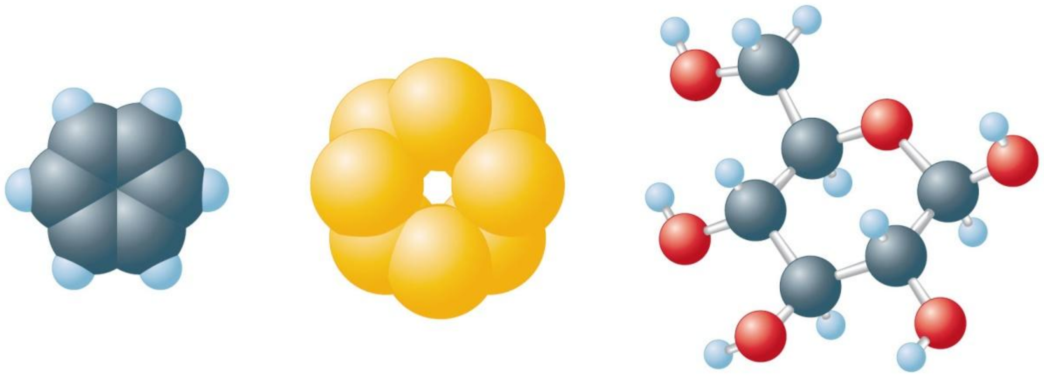
Rank them from highest to lowest percent oxygen by mass.

Determining the Formula of a Compound

Formulas of Compounds

Empirical Formulas

- The empirical formula of a compound is the simplest whole number ratio of the atoms present in the compound.



$C_6H_6 = (CH)_6$ $S_8 = (S)_8$ $C_6H_{12}O_6 = (CH_2O)_6$

- The empirical formula can be found from the percent composition of the compound. That is, we use the measured masses of the elements present to determine the formula.

Calculation of Empirical Formulas

4 Steps:

1. Obtain the mass of each element present (in grams).

2. Determine the number of moles of each type of atom present.

3. Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all of the numbers so obtained are integers, these are the subscripts in the empirical formula. If one or more of these numbers are not integers, go on to step 4.

4. Multiply the numbers you derived in step 3 by the smallest integer that will convert all of them to whole numbers. This set of whole numbers represents the subscripts in the empirical formula.

Empirical Formula (Example)

A gaseous compound containing carbon and hydrogen was analyzed and found to consist of 83.65% carbon by mass. Determine the empirical formula of the compound.

1. Obtain the mass of each element present (in grams).

Assume you have 100g of the compound.

83.65% C = 83.65 g C

% H = 100.00 – 83.65

% H = 16.35

16.35% H = 16.35 g H

Empirical Formula (Example –Continued)

2. Determine the number of moles of each type atom present.

$$83.65 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} =$$

$$16.35 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} =$$

Empirical Formula (Example –Continued)

3. Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all of the numbers so obtained are integers, these are the subscripts in the empirical formula. If one or more of these numbers are not integers, go on to step 4.

$$\frac{6.964 \text{ mol C}}{6.964 \text{ mol}} = 1 \qquad \frac{16.22 \text{ mol H}}{6.964 \text{ mol}} = 2.329$$

Can’t have a partial H bonded to a C atom. Must be a whole atom!

Empirical Formula (Example –Continued)

4. Multiply the numbers you derived in step 3 by the smallest integer that will convert all of them to whole numbers. This set of whole numbers represents the subscripts in the empirical formula.

$$\text{C: } 1 \times 3 = 3$$

$$\text{H: } 2.329 \times 3 = 6.987$$

The empirical formula is C_3H_7.

Empirical Formula (Example)

The most common form of nylon (Nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.14% oxygen. Calculate the empirical formula.

1. Where are we going?

   Determine $C_aN_bH_cO_d$ where we solve for *a, b, c, d*

2. What do we know?
   - The compound composition (mass percent) – 63.68% C, 12.38% N, 9.80% H, and 14.14% O
   - The atomic masses of C (12.011 g/mol), N (14.007 g/mol), H (1.008 g/mol) and O (15.999 g/mol).

3. What do we need to know?
   - *a,b,c* and *d* represent moles of atoms in 1 mole of the compound, we need to determine the relative number of moles of C, N, H and O.
   - We have mass percent data, and to get to the number of moles we need to know the mass of each element (g) in the sample.

Empirical Formula (Example –Continued)

The most common form of nylon (Nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.14% oxygen. Calculate the empirical formula.

4. How do we get there?

Step 1: Determine the mass of each element present (in 100 g sample).

C = 63.68 g

N = 12.38 g

H = 9.80 g

O= 14.14 g

Empirical Formula (Example –Continued)

The most common form of nylon (Nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.14% oxygen. Calculate the empirical formula.

4. How do we get there?

Step 2: Determine the number of moles of each type of atom. (Use atomic masses)

$$63.68\ \cancel{g\ C} \times \frac{1\ mol\ C}{12.011\ \cancel{g\ C}} = 5.302\ mol\ C$$

$$12.38\ \cancel{g\ N} \times \frac{1\ mol\ N}{14.007\ \cancel{g\ N}} = 0.8838\ mol\ N$$

$$9.80\ \cancel{g\ H} \times \frac{1\ mol\ H}{1.008\ \cancel{g\ H}} = 9.72\ mol\ H$$

$$14.14\ \cancel{g\ O} \times \frac{1\ mol\ O}{15.999\ \cancel{g\ O}} = 0.8838\ mol\ O$$

Empirical Formula (Example –Continued)

The most common form of nylon (Nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.14% oxygen. Calculate the empirical formula.

4. How do we get there?

Step 3: Divide through by the smallest number of moles.

$$\frac{5.302 \text{ mol C}}{0.8838} = 5.999 \text{ mol C}$$

$$\frac{0.8838 \text{ mol N}}{0.8838} = 1.000 \text{ mol N}$$

$$\frac{9.72 \text{ mol H}}{0.8838} = 11.0 \text{ mol H}$$

$$\frac{0.8838 \text{ mol O}}{0.8838} = 1.000 \text{ mol O}$$

The empirical formula for Nylon-6 is **$C_6H_{11}NO$**

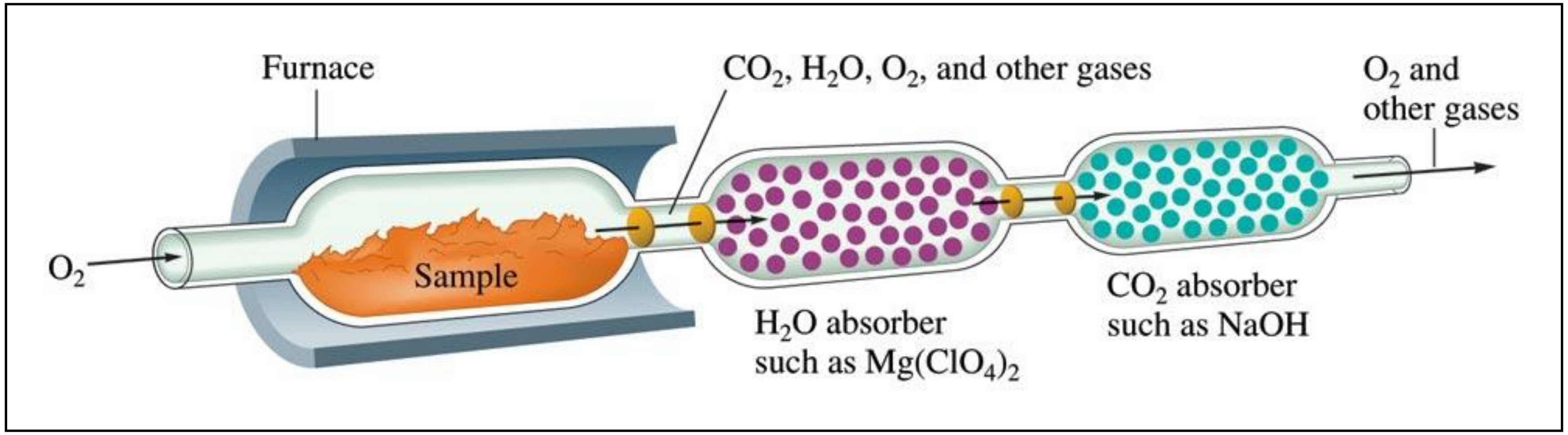
Exercise

Determine the empirical formula for a compound that gives the following mass percentages on analysis:

71.65% Cl   24.27% C   4.07% H

Combustion Analysis

- One way of carrying out an elemental analysis
- Requires a device to determine the mass percent of each element in a compound
- The data collected is used to determine the empirical formula.



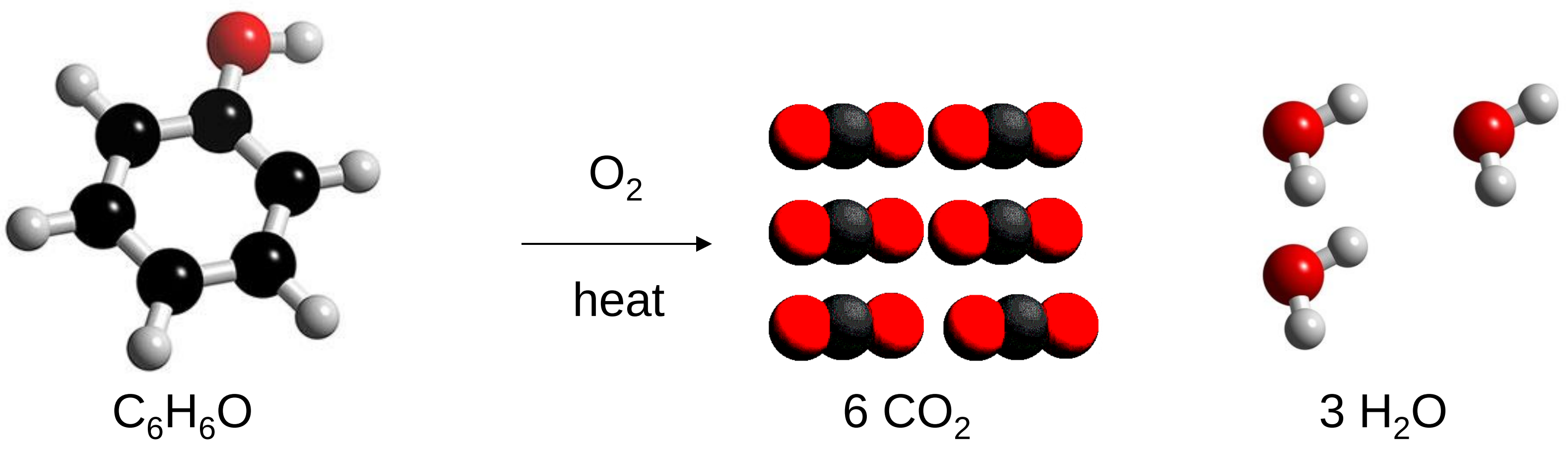
Elemental Analysis Procedure

Elemental Analysis (using combustion):

An accurately weighed compound is burned in the presence of excess oxygen so that:

- All carbon in the sample becomes CO_2
- All hydrogen in the sample becomes H_2O
- All nitrogen in the sample becomes N_2

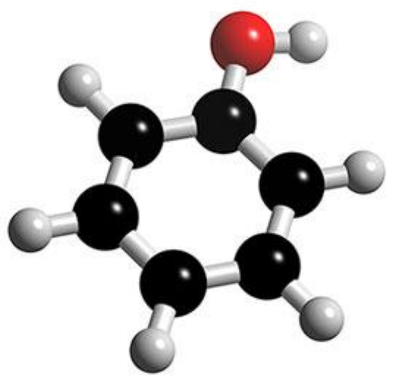
} Weighed separately (applies to CO_2 and H_2O)



Elemental Analysis Procedure

To determine the empirical formula:

- Involves the same four steps except first step requires additional calculations.



C_6H_6O

$\xrightarrow[\text{heat}]{O_2}$

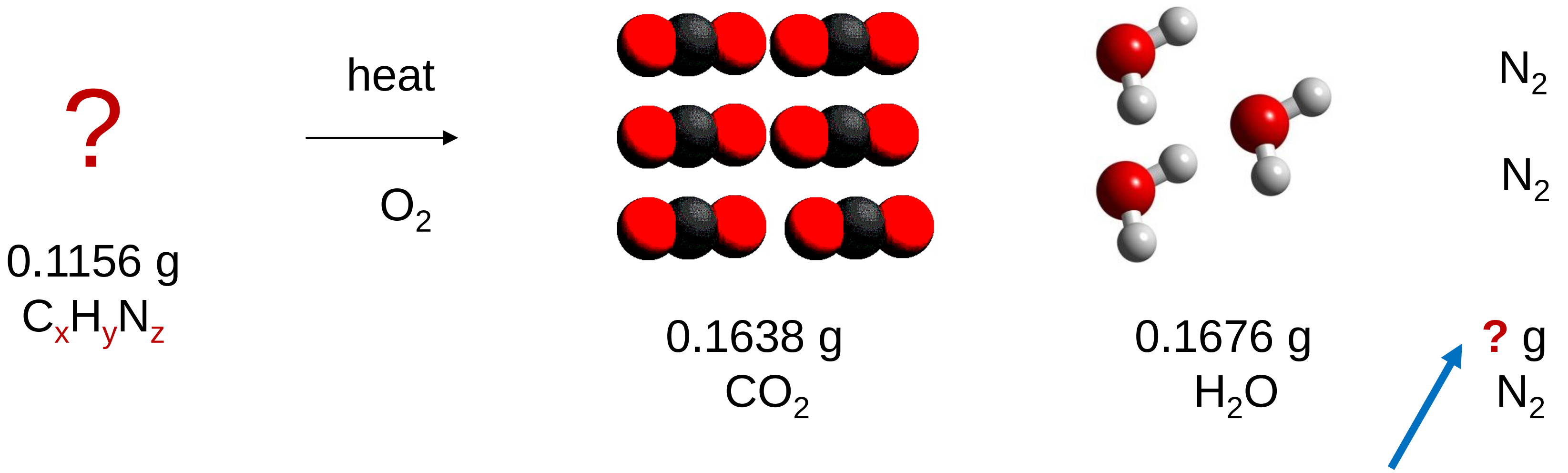
- All carbon in the sample becomes CO_2
- All hydrogen in the sample becomes H_2O
- All nitrogen in the sample becomes N_2

- Must determine how much carbon is in the CO_2, how much hydrogen is in the H_2O.

Example

A substance that is composed of carbon, hydrogen, and nitrogen has undergone an elemental analysis. When 0.1156 g of the substance is burned in the presence of excess oxygen, 0.1638 g of CO_2 and 0.1676 g of H_2O are collected. Determine the empirical formula.

1. Obtain the mass of each element present (in grams).



NOTE: we need to be told if the compound contains O or N atoms because we do not measure them directly

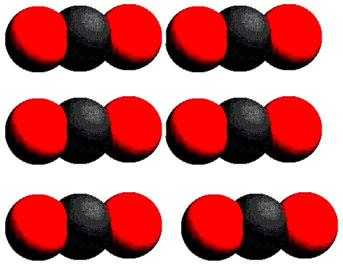
Example - Continued

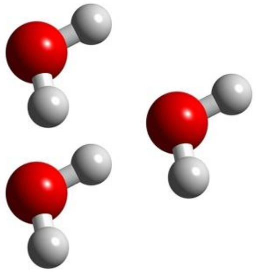
1. Obtain the mass of each element present (in grams).

?

0.1156 g $C_xH_yN_z$

heat → O_2





N_2

N_2

0.1638 g CO_2

0.1676 g H_2O

? g N_2

Determining mass of C in CO_2 first by calculating mass percent of C (assume 1 mole)

1. Mass % of C in CO_2

$$\frac{12.011 \text{ g C}}{44.009 \text{ g } CO_2} \times 100\% = 27.29\% \text{ C}$$

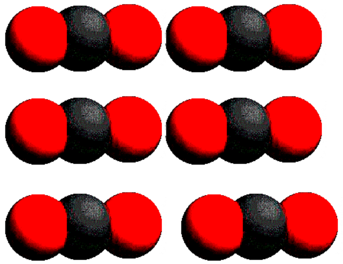
2. Calculate mass of C in CO_2, which represents mass of C from the original sample (0.1156 g).

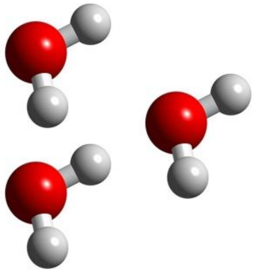
$$0.1638 \text{ g } CO_2 \times \frac{27.29\% \text{ C}}{100\% \; CO_2} = 0.04470 \text{ g C}$$

Example - Continued

1. Obtain the mass of each element present (in grams).

? heat $\xrightarrow[O_2]{}$





N_2

N_2

0.1156 g $C_xH_yN_z$ | 0.1638 g CO_2 | 0.1676 g H_2O | ? g N_2

Determining mass of H in H_2O first by calculating mass percent of H (assume 1 mole)

1. Mass % of H in H_2O

$$\frac{2.016 \text{ g H}}{18.015 \text{ g } H_2O} \times 100\% = 11.19\% \text{ H}$$

2. Calculate mass of H in H_2O, which represents mass of H from the original sample (0.1156 g).

$$0.1676 \text{ g } H_2O \times \frac{11.19\% \text{ H}}{100\% \text{ } H_2O} = 0.01875 \text{ g H}$$

Example - Continued

1. Obtain the mass of each element present (in grams).

?

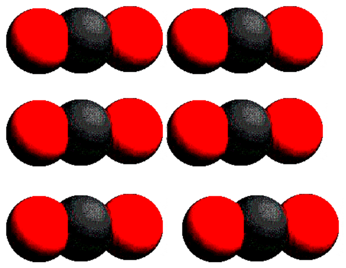
heat

$\xrightarrow{}$

O_2

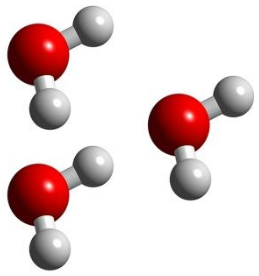
0.1156 g

$C_xH_yN_z$



0.1638 g

CO_2



0.1676 g

H_2O

N_2

N_2

? g

N_2

Determining mass of N

In combustion analysis, N gets converted to N_2.

Total Mass of Sample = Mass of C + Mass of H + Mass of N

Mass of N = Total Mass of Sample - Mass of C - Mass of H

= 0.1156 g – 0.04470 g – 0.01875 g

= 0.0521 g

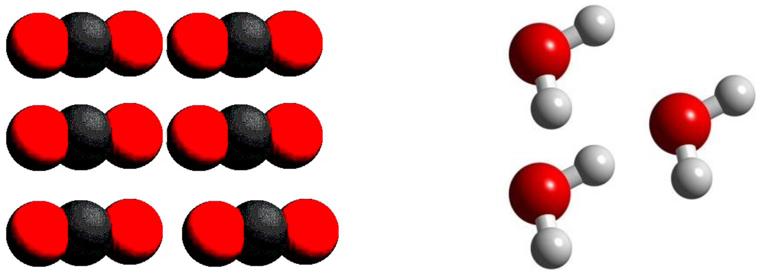
Example - Continued

1. Obtain the mass of each element present (in grams).

?

0.1156 g $C_xH_yN_z$

heat → O_2



N_2

N_2

0.1638 g CO_2

0.1676 g H_2O

? g N_2

The mass of each element in the original sample that was burned:

0.1156 g $C_xH_yN_z$ ➡

Mass of C = 0.04470 g
Mass of H = 0.01875 g
Mass of N = 0.0521 g

Example - Continued

2. Determine the number of moles of each type atom present.

$$0.04470 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 0.003722 \text{ mol C}$$

$$0.01875 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.01860 \text{ mol H}$$

$$0.0521 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 0.00372 \text{ mol N}$$

Example - Continued

3. Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all of the numbers so obtained are integers, these are the subscripts in the empirical formula. If one or more of these numbers are not integers, go on to step 4.

$$\text{C:} \quad \frac{0.003722}{0.00372} = 1$$

$$\text{H:} \quad \frac{0.01860}{0.00372} = 5$$

$$\text{N:} \quad \frac{0.00372}{0.00372} = 1$$

Empirical Formula = CH_5N

Formulas of Compounds

- Empirical formula = CH
  - Simplest whole-number ratio

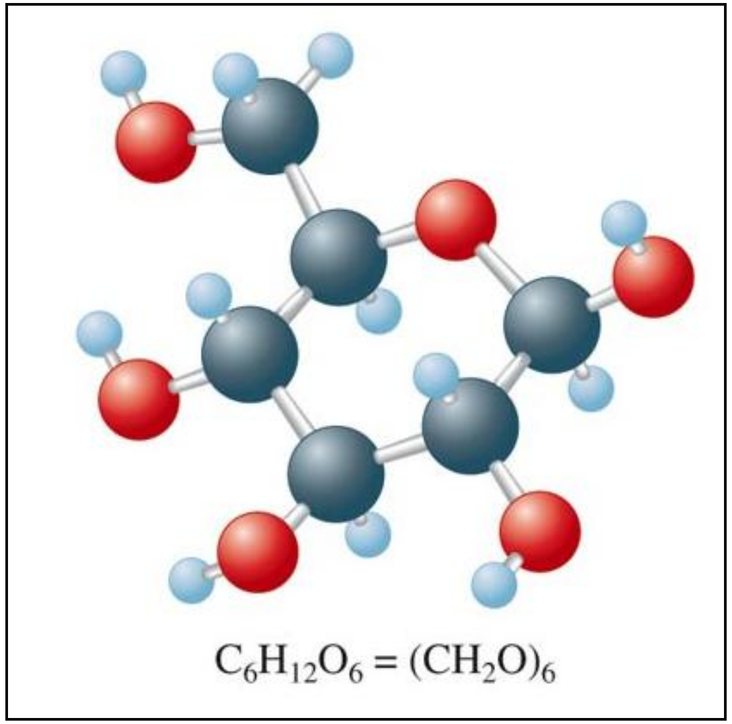
- Molecular formula = C_6H_6 = $(CH)_6$
  - The molecular formula is the exact formula of the molecules present in a substance.
  - The molecular formula is always an integer multiple of the empirical formula.

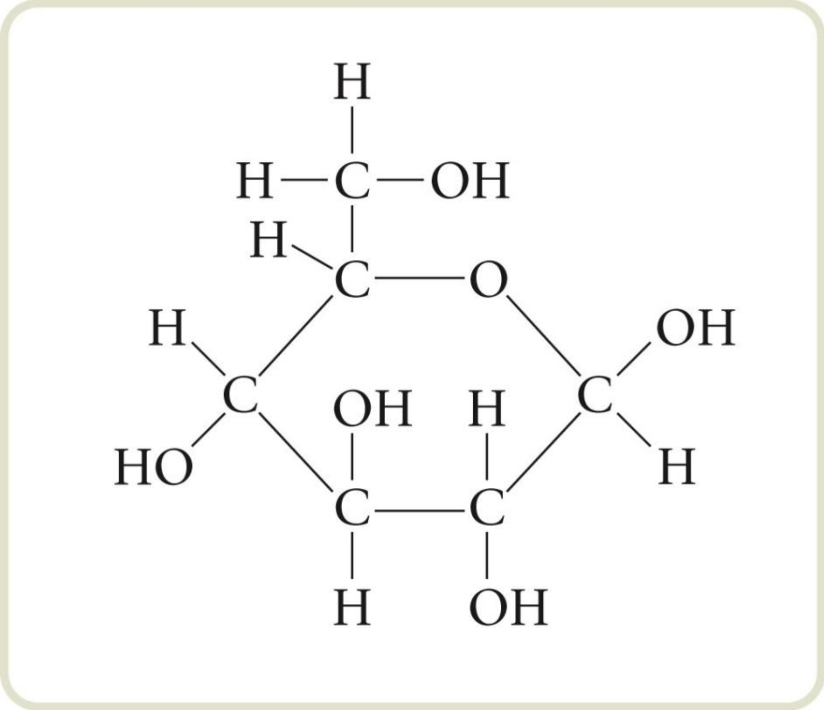
Molecular formula = $(\text{empirical formula})_n$

[n = integer]

Empirical vs Molecular Formula

Glucose Molecule





Molecular formula = $C_6H_{12}O_6$

= $(\text{Empirical formula})_n$

= $(CH_2O)_6$

Determining Molecular Formula from Empirical Formula

4 Steps:

1. Obtain the empirical formula.

2. Compute the mass corresponding to the empirical formula.

3. Calculate the ratio

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

4. The integer from step 3 represents the number of empirical formula units in one molecule. When the empirical formula subscripts are multiplied by this integer, the molecular formula results.

$$\text{Molecular formula} = (\text{empirical formula}) \times \frac{\text{Molar mass}}{\text{Empirical formula mass}}$$

Example

A gaseous compound containing carbon and hydrogen was analyzed and found to consist of 83.65% carbon by mass. The molar mass of the compound is 86.2 g/mol. You determined the empirical formula to be C_3H_7. What is the molecular formula of the compound?

Molar mass of C_3H_7 = 43.089 g/mol

$$\frac{86.2\ \cancel{\text{g/mol}}}{43.089\ \cancel{\text{g/mol}}} = 2$$

Molecular formula = (empirical formula)n

= $(C_3H_7)_2$

= **C_6H_{14}**