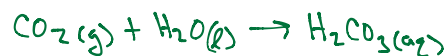




## Types of chemical reactions:



two or more compounds  
react to form one



- opposite of synthesis
- one compound breaks apart to form two or more new compounds



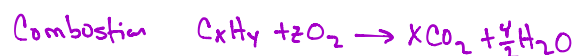
- one component of a binary compound exchanges with a monoatomic molecule



} commonly ion exchange



- Cations in ionic compounds switch place



- Hydrocarbon reacts with molecular oxygen;  $\text{CO}_2$  &  $\text{H}_2\text{O}$  are product



So for the rest of this class, we'll be focusing on chemical reactions. One MAJOR thing that chemists need to do on a regular basis is to predict how much product will form given certain amounts of reactants

But here is the catch: masses of reactants and products are NOT directly related

For example:



If we start with 1.00 g Na(s) and excess KCl, there is NO straight forward way to calculate the amount of NaCl will be produced

← excess, so we won't run out

We could use the number of atoms:

10 Na atoms will produce 10 NaCl molecules

but we don't have a good way to measure atoms & - only mass is measurable

- So we're going to introduce a new unit that allows us to relate atoms to grams - the mole -

conceptually, a mole is absolutely no different than a dozen

1 dozen = 12 units

If you have 2 dozen donuts, how many donuts?

$$\frac{2 \text{ dozen} | 12 \text{ donuts}}{1 \text{ dozen}} = 24 \text{ donuts}$$

36 cars → how many dozen?

$$\frac{36 \text{ cars} | 1 \text{ dozen}}{12 \text{ cars}} = 3 \text{ dozen}$$

Treat moles the same way: (Just a much bigger number)

$$1 \text{ mol} = 6.022 \times 10^{23} \text{ units}$$

46 moles of donuts → how many donuts?

$$\frac{46 \text{ mol} | 6.022 \times 10^{23} \text{ donuts}}{1 \text{ mol}} = 2.77 \times 10^{25} \text{ donuts}$$

$1 \times 10^{20}$  atoms of carbon. How many moles?

$$\frac{1 \times 10^{20} \text{ atoms} | 1 \text{ mol}}{6.022 \times 10^{23}} = 1.66 \times 10^{-4} \text{ mol}$$

Since moles are DIRECTLY related to number of atoms, we can use this unit to relate atoms + molecules to other atoms/molecules

If I have 4.5 moles of Carbon Dioxide: ① How many moles of carbon?

② How many moles of Oxygen?

① We can use the subscripts in  $\text{CO}_2$  as conversion factors:

1 mol  $\text{CO}_2$  = 1 mol Carbon

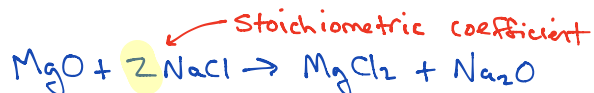
$$4.5 \text{ mol CO}_2 \left| \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right. = 4.5 \text{ mol Carbon}$$

$$4.5 \text{ mol Carbon} \left| \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \right. = 2.71 \times 10^{24} \text{ atoms carbon}$$

②  $4.5 \text{ mol CO}_2 \left| \frac{2 \text{ mol O}}{1 \text{ mol CO}_2} \right. = 9 \text{ mol Oxygen}$

$$9 \text{ mol O} \left| \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \right. = 5.42 \times 10^{24} \text{ atoms oxygen}$$

We can also use the mol unit to relate the amount of products and reactants in a chemical reaction:



If I have 10 moles of  $\text{MgO}$  and excess  $\text{NaCl}$ , how much  $\text{MgCl}_2$  can we produce?

- use the stoichiometric as a conversion factor: 1 mol  $\text{MgO}$  = 1 mol  $\text{MgCl}_2$

$$10 \text{ mol MgO} \left| \frac{1 \text{ mol MgCl}_2}{1 \text{ mol MgO}} \right. = 10 \text{ mol MgCl}_2$$

What if we have 10 moles of  $\text{NaCl}$  and excess  $\text{MgO}$ , how much  $\text{MgCl}_2$  will be produced

$$10 \text{ mol NaCl} \left| \frac{1 \text{ mol MgCl}_2}{2 \text{ mol NaCl}} \right. = 5 \text{ mol MgCl}_2$$

Since it takes 2 moles of  $\text{NaCl}$  to produce 1 mol of  $\text{MgCl}_2$ , we only get 5 moles

But... we cannot measure moles or atoms! We CAN measure mass

One of the most important conversions we'll see this term

$g \rightarrow mol$  This conversion factor comes from the periodic table.

- The average isotopic mass of an element (e.g. Carbon is 12.011)

has the unit  $\frac{\text{grams}}{1 \text{ mol}}$ . This is the **Molecular Weight**

**Molar mass**

**Atomic mass**

**Atomic weight**

- these terms tend to be used interchangeably. The most correct one **highlighted**

moles can be calculated from the mass of a substance using the molecular weight

How many moles of carbon are present in 1.00 g?  $MW = 12.011 \text{ g/mol}$  (from periodic table)

$$\frac{1.00 \text{ g C}}{12.011 \text{ g}} \times \frac{1 \text{ mol}}{1 \text{ mol}} = 0.0833 \text{ mol C}$$

How many carbon atoms?

$$\frac{0.0833 \text{ mol C}}{1 \text{ mol}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 5.01 \times 10^{22} \text{ atoms}$$

16.82 g  $\text{CO}_2 \rightarrow$  How many grams of oxygen?

$$\text{CO}_2 \rightarrow MW = 12.011 + 2(16.) = 44.011 \text{ g/mol}$$

\* need to go through moles! \*

$$\frac{16.82 \text{ g CO}_2}{44.011 \text{ g CO}_2} \times \frac{1 \text{ mol CO}_2}{1 \text{ mol CO}_2} \times \frac{2 \text{ mol O}}{1 \text{ mol CO}_2} \times \frac{16 \text{ g}}{1 \text{ mol O}} = 12.23 \text{ g Oxygen}$$

## Class Problem:

Using  $\times$  dot to aid in calculations.

Consider  $\text{CH}_4$ . If you are told that you have 172 grams of carbon in  $\text{CH}_4$ , how many atoms (H + C) exist in this sample?

Strategy: There are several ways to tackle this problem. My approach is to recognize that there are 4 Hydrogens for every 1 Carbon. If we are able to determine how many C atoms, we can determine H

mass C  $\rightarrow$  moles C  $\rightarrow$  atoms C  $\rightarrow$  atoms H  $\rightarrow$  total atoms

$$\frac{172 \text{ g C}}{12.011 \text{ g C}} \times \frac{1 \text{ mol C}}{1 \text{ mol C}} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol}} = 8.62 \times 10^{24} \text{ C atoms}$$

$$\frac{8.62 \times 10^{24} \text{ C atoms}}{1 \text{ C atom}} \times \frac{4 \text{ H atoms}}{1 \text{ C atom}} = 3.45 \times 10^{25} \text{ H atoms}$$

$$8.62 \times 10^{24} + 3.45 \times 10^{25} = 4.31 \times 10^{25}$$

Now calculate the mass of  $\text{CH}_4$  which contains 172 g Carbon

Strategy: mass C  $\rightarrow$  moles C  $\rightarrow$  mol  $\text{CH}_4$   $\rightarrow$  mass  $\text{CH}_4$

$$\frac{172 \text{ g C}}{12.011 \text{ g C}} \times \frac{1 \text{ mol C}}{1 \text{ mol C}} \times \frac{1 \text{ mol CH}_4}{1 \text{ mol C}} \times \frac{16.05 \text{ g CH}_4}{1 \text{ mol CH}_4} = 229.86 \text{ g CH}_4$$

IN 229.86 grams of  $\text{CH}_4$ , carbon accounts for 172 g

$$\text{The \%} \rightarrow \frac{172}{229.86} \times 100 = 74.83 \%$$

**Mass %**  $\rightarrow$  The percentage of a molecule's mass that comes from one atom

$\uparrow$  Elemental analysis data is MUCH more common in the form of mass %  
 $\uparrow$  this is JUST a conversion factor!

$\text{CO}_2$  is 72.71% oxygen by mass. If you have 42.86 g  $\text{CO}_2$ , what mass of oxygen?

$$\frac{72.71 \text{ g oxygen}}{100 \text{ g CO}_2}$$

$$\frac{42.86 \text{ g CO}_2}{100 \text{ g CO}_2} \times \frac{72.71 \text{ g O}}{100 \text{ g CO}_2} = 31.16 \text{ g oxygen}$$

calculating mass% → these are just mole conversions! What is the mass% C in C<sub>2</sub>H<sub>4</sub>O<sub>2</sub>?  
 Let's assume 100g C<sub>2</sub>H<sub>4</sub>O<sub>2</sub>.

$$\text{mass \%} = \frac{x}{100 \text{ g C}_2\text{H}_4\text{O}_2} \times 100$$

$$x = \text{mass C in } 100 \text{ g C}_2\text{H}_4\text{O}_2 \quad \begin{matrix} 12.01(2) + 4(1.01) + 2(16) \\ 24.02 \quad 4.04 \quad 32 \\ 60.06 \text{ g/mol} \end{matrix}$$

$$\frac{39.96 \text{ g C}}{100 \text{ g C}_2\text{H}_4\text{O}_2} \times 100 = 39.96\%$$

$$\frac{100 \text{ g C}_2\text{H}_4\text{O}_2}{60.06 \text{ g}} \times \frac{1 \text{ mol C}_2\text{H}_4\text{O}_2}{1 \text{ mol C}_2\text{H}_4\text{O}_2} \times \frac{2 \text{ mol C}}{1 \text{ mol C}_2\text{H}_4\text{O}_2} \times \frac{12.01 \text{ g}}{1 \text{ mol}} = 39.96 \text{ g C}$$

Now calculate H + O  
 ↑  
 7.19%      53.28%

How else can we use mol conversions? Determining molecular Formulas!

First, we need to categorize compound formulas:

Molecular Formula: the exact atomic composition of a molecule

Empirical Formula: The lowest possible ratio of atoms in a compound

<u>Molecular Formula</u>	→	<u>Empirical Formula</u>
CO <sub>2</sub>	→	CO <sub>2</sub>
C <sub>2</sub> O <sub>4</sub>	→	CO <sub>2</sub>
C <sub>3</sub> H <sub>6</sub>	→	CO <sub>2</sub>

Lots of compounds can have the same Empirical Formula!

Decomposition of a sample reveals that it is composed of: 3.6 g Carbon

0.61 g Hydrogen

4.8 g Oxygen

① Determine the Empirical Formula.

② Determine the molecular formula if

this compound has a MW = 90.09

① 1. convert to moles

$$\text{C: } \frac{3.6 \text{ g C}}{12.01 \text{ g}} \times \frac{1 \text{ mol}}{1} = \frac{0.2998 \text{ mol C}}{0.2998} = 1 \text{ mol C}$$

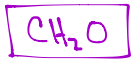
2. Divide by lowest number of moles

$$\text{H: } \frac{0.61 \text{ g H}}{1.01 \text{ g}} \times \frac{1 \text{ mol}}{1} = \frac{0.604 \text{ mol H}}{0.2998} = 2.01 \text{ mol H}$$

3. divide by lowest decimal again if necessary.

$$\text{O: } \frac{4.8 \text{ g O}}{16. \text{ g}} \times \frac{1 \text{ mol}}{1} = \frac{0.3 \text{ mol O}}{0.2998} = 1 \text{ mol O}$$

this assumes that at least one atom has a coeff. of 1



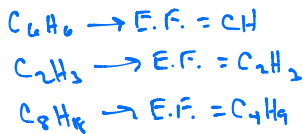
② - Determine MW of Empirical Formula —  $12.01 + 2(1.01) + 16 = 30.03 \text{ g/mol}$   
 - Since E.F. is M.F.  $\div$  an integer, this is also true for M.W.

$$\frac{\text{M.W.}}{\text{E.F.W.}} = n \quad \frac{90.09}{30.03} = 3 \quad \text{E.F.} = \text{CH}_2\text{O} \quad \leftarrow \text{3x}$$

Molecular Formula = C<sub>3</sub>H<sub>6</sub>O<sub>3</sub>

Empirical Formula Weight

Molecular Formula vs. Empirical Formula  
 actual formula      lowest integer ratio of atoms



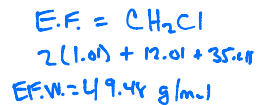
This E.F. is useful because it lets us quickly determine the identity of a compound from experimental data!  
 • Elemental analysis determines that a compound contains 24.27 g Carbon, 4.075g Hydrogen + 71.65 g Chlorine and has a molecular weight of 98.96 g/mol

\*mass is NOT reliable → convert to moles!\*

$$\frac{24.27 \text{ g C}}{12.01 \text{ g}} \cdot \frac{1 \text{ mol}}{1} = \frac{2.02 \text{ mol C}}{2.02} = 1$$

$$\frac{4.075 \text{ g H}}{1.01 \text{ g}} \cdot \frac{1 \text{ mol}}{1} = \frac{4.03 \text{ mol H}}{2.02} = 2$$

$$\frac{71.65 \text{ g Cl}}{35.45 \text{ g}} \cdot \frac{1 \text{ mol}}{1} = \frac{2.02 \text{ mol Cl}}{2.02} = 1$$



$$\frac{\text{MW}}{\text{E.F.W.}} = \frac{\text{number of times the Molecular formula is bigger than the empirical formula}}{1} = \frac{98.96}{49.44} = 2$$



Consider Trifluoroacetic acid:

$$21.6\% \text{ C} \left| \frac{\text{mol}}{12.011 \text{ g}} = \frac{1.75}{0.876} = 2$$

$$49.9\% \text{ F} \left| \frac{\text{mol}}{19 \text{ g}} = \frac{2.63}{0.876} = 3$$

$$0.87\% \text{ H} \left| \frac{\text{mol}}{1.01 \text{ g}} = \frac{0.876}{0.876} = 1$$

$$28.06\% \text{ O} \left| \frac{\text{mol}}{16 \text{ g}} = \frac{1.75}{0.876} = 2$$

