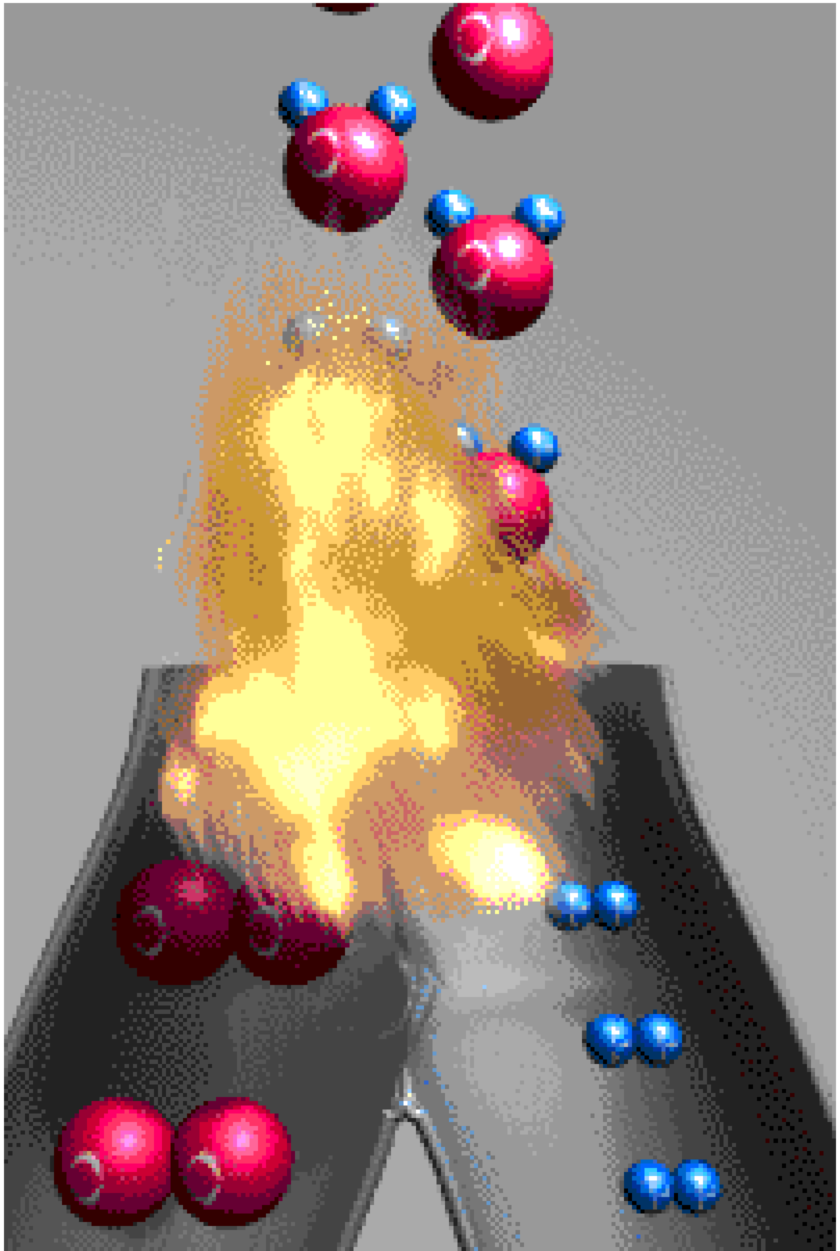


CHEMICAL REACTIONS

Reactants: $Zn + I_2$ $\longrightarrow$ **Product: $Zn I_2$**

Chemical Equations

Their Job: Depict the kind of reactants and products and their relative amounts in a reaction.

$4\ Al_{(s)} + 3\ O_{2\ (g)} \text{ ---> } 2\ Al_2O_{3\ (s)}$

The letters (s), (g), and (l) are the physical states of compounds.

Parts of a Reaction Equation

- Chemical equations show the conversion of reactants (the molecules shown on the left of the arrow) into products (the molecules shown on the right of the arrow).
  - A + sign separates molecules on the same side
  - The arrow is read as “yields”
  - Example

    $C + O_2 \rightarrow CO_2$
  - This reads “carbon plus oxygen react to yield carbon dioxide”

- The charcoal used in a grill is basically carbon. The carbon reacts with oxygen to yield carbon dioxide. The chemical equation for this reaction, $C + O_2 \rightarrow CO_2$, contains the same information as the English sentence but has a quantitative meaning as well.

Chemical Equations

Because of the principle of the conservation of matter,

an equation must be balanced.

It must have the same number of atoms of the same kind on both sides.

Lavoisier, 1788

Symbols Used in Equations

- **Solid (c)**
- **Liquid (l)**
- **Gas (g)**
- **Aqueous solution (aq)**
- **Catalyst** $\xrightarrow{H_2SO_4}$
- **Escaping gas (↑)**
- **Change of temperature (Δ)**

Law of Conservation of Mass

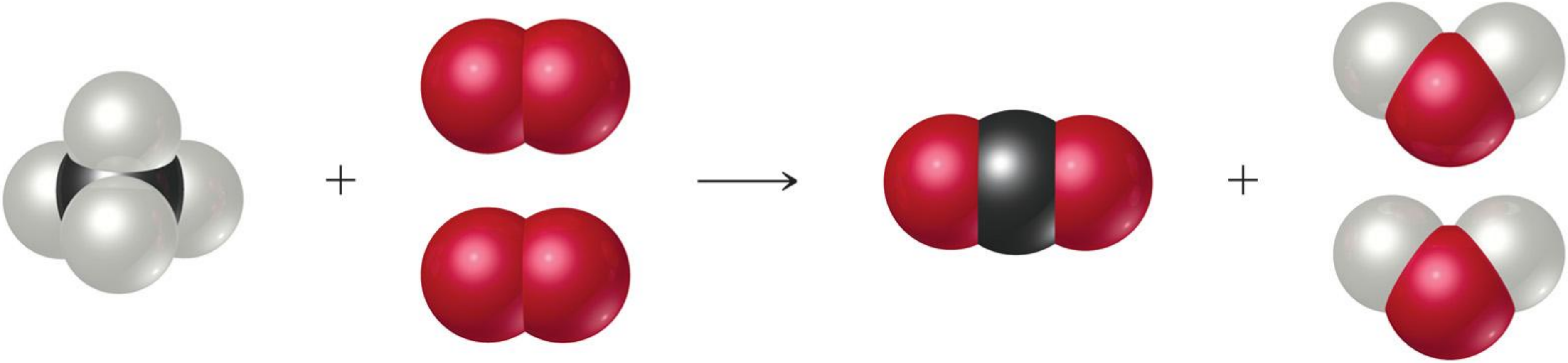
"We may lay it down as an incontestable axiom that, in all the operations of art and nature, nothing is created; an equal amount of matter exists both before and after the experiment. Upon this principle, the whole art of performing chemical experiments depends."

--Antoine Lavoisier, 1789

Chemical Equations

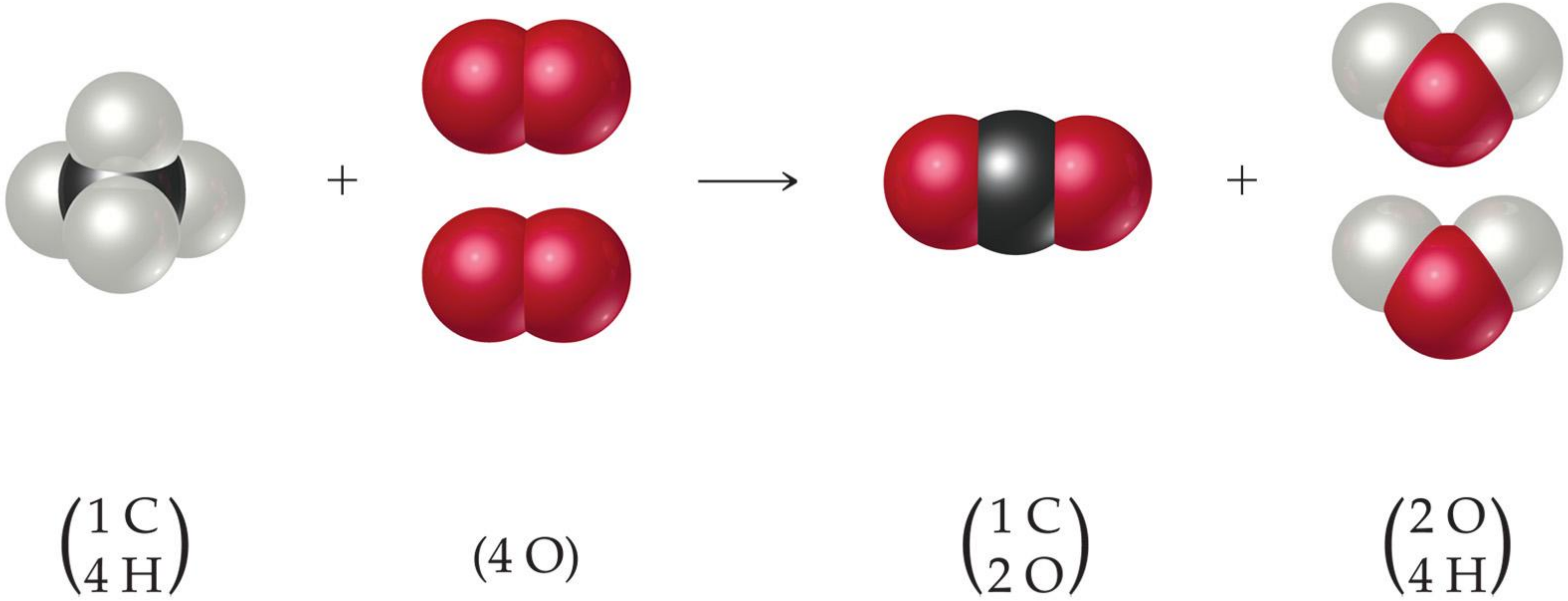
Concise representations of chemical reactions

“Chemical Sentences”



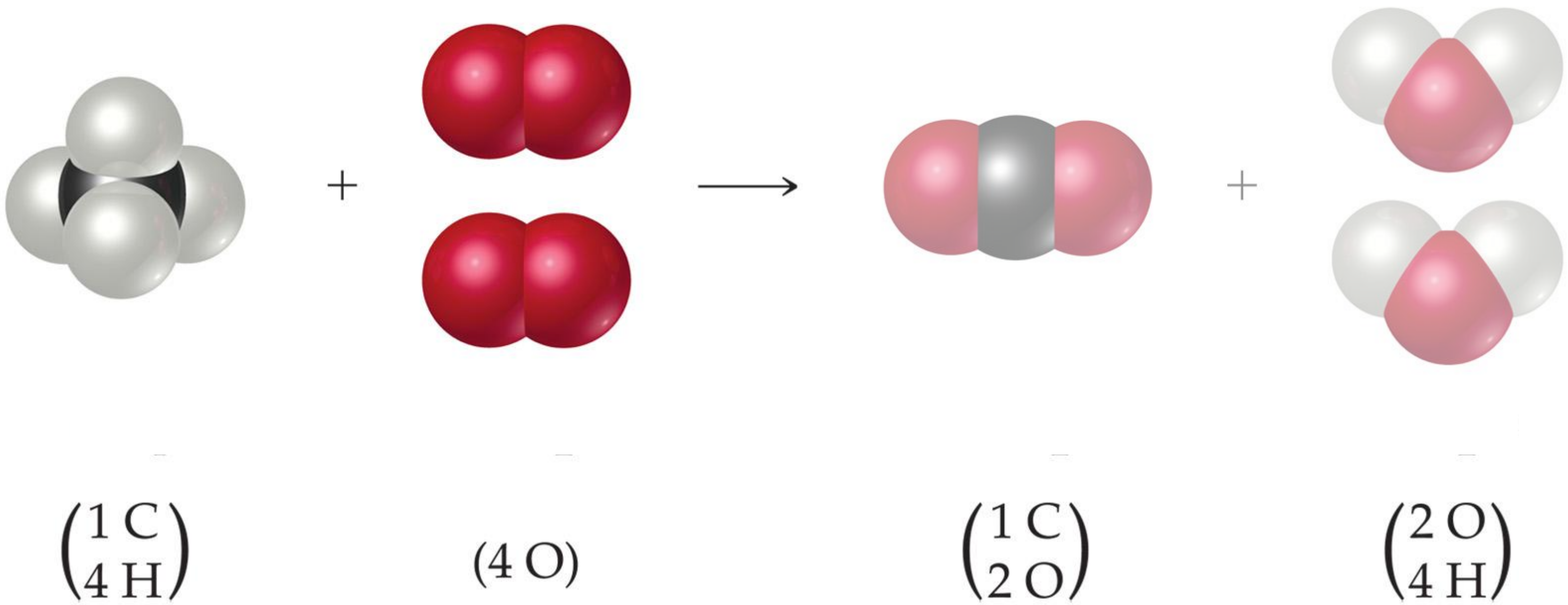
Anatomy of a Chemical Equation

$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$



Anatomy of a Chemical Equation

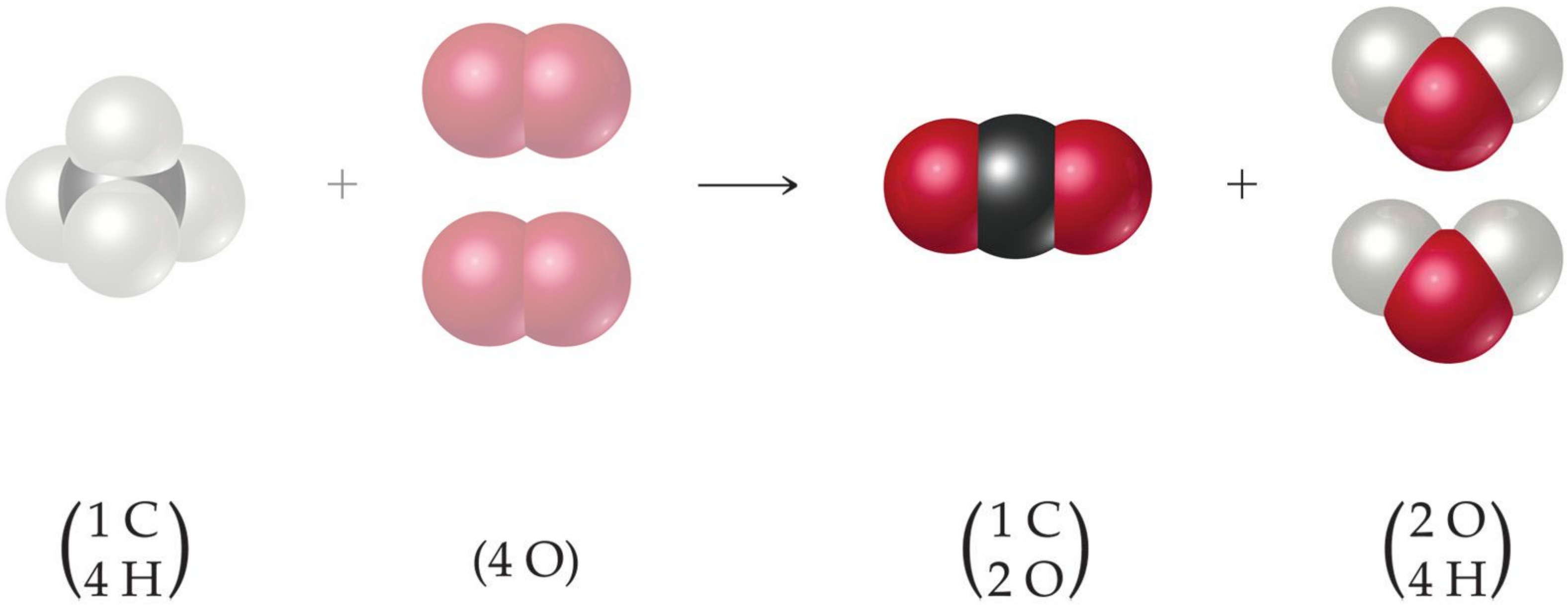
$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$



Reactants appear on the left side of the equation.

Anatomy of a Chemical Equation

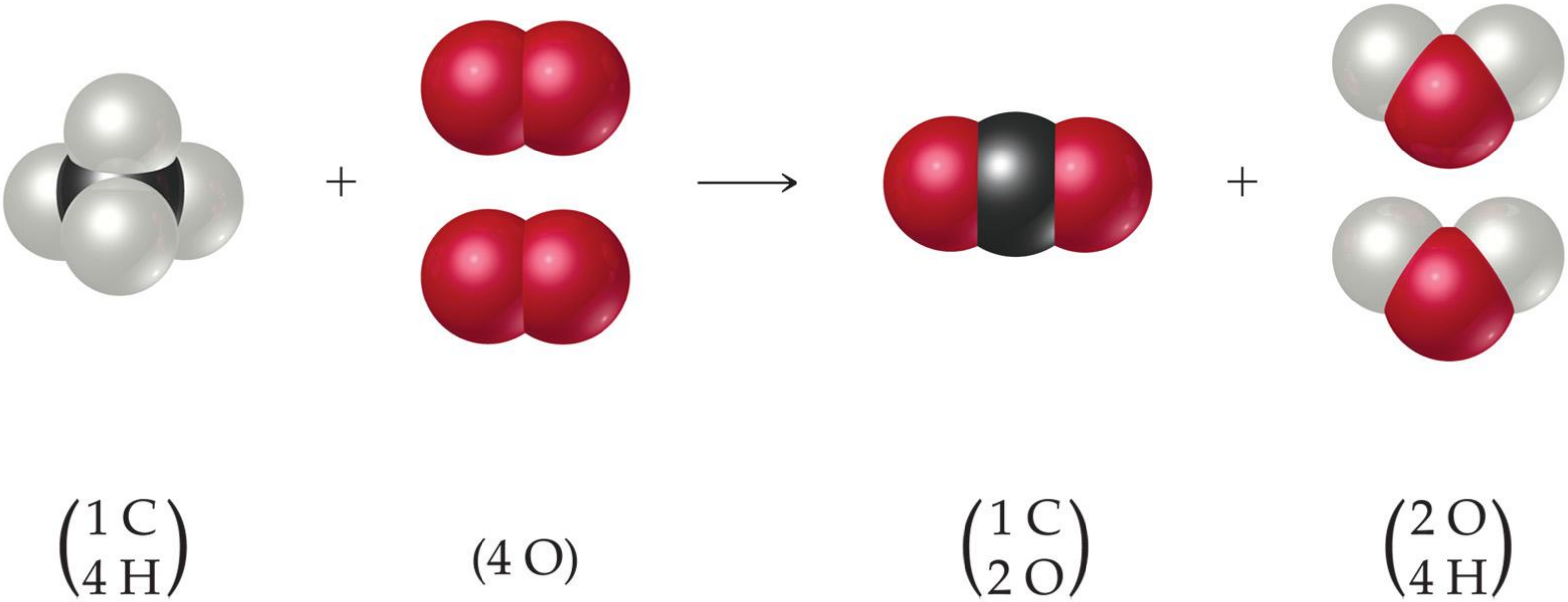
$$CH_{4\,(g)} + 2\ O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\ H_2O_{\,(g)}$$



Products appear on the right side of the equation.

Anatomy of a Chemical Equation

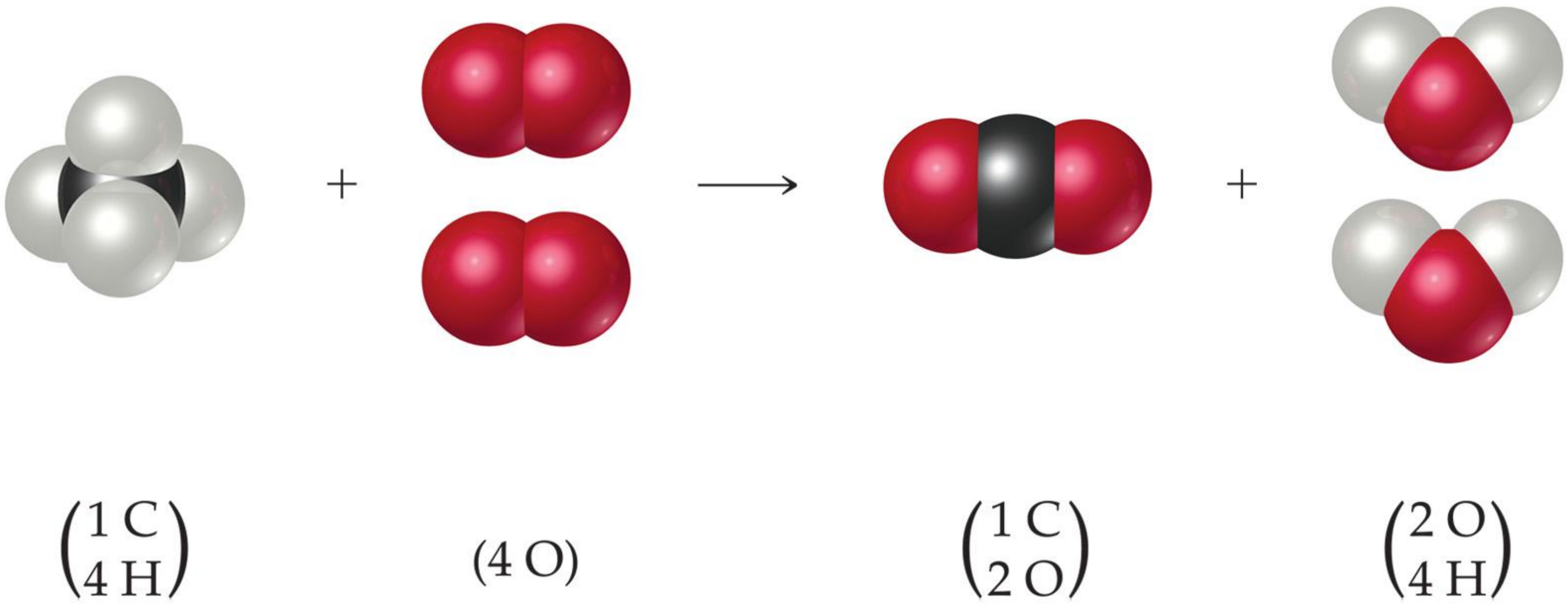
$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$



The states of the reactants and products are written in parentheses to the right of each compound.

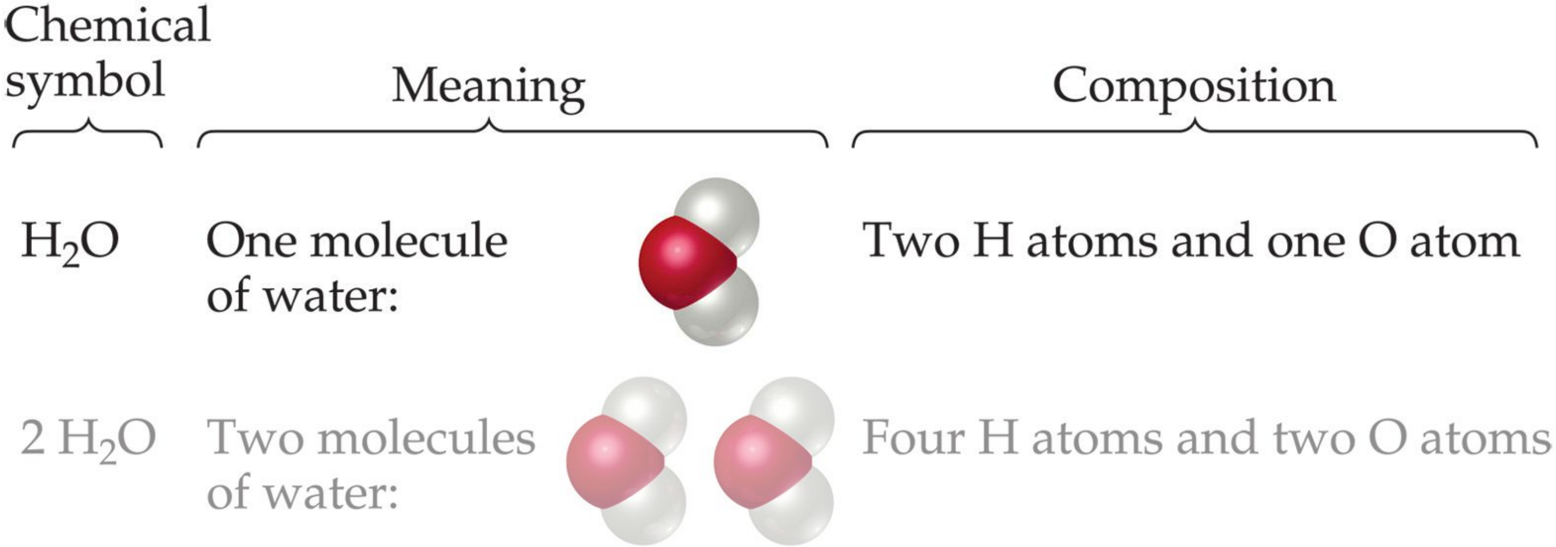
Anatomy of a Chemical Equation

$$CH_{4\,(g)} + 2\ O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\ H_2O\ _{(g)}$$



Coefficients are inserted to balance the equation.

Subscripts and Coefficients Give Different Information



Chemical symbol	Meaning	Composition
H_2O	One molecule of water:	Two H atoms and one O atom
$2\ H_2O$	Two molecules of water:	Four H atoms and two O atoms

- Subscripts tell the number of atoms of each element in a molecule

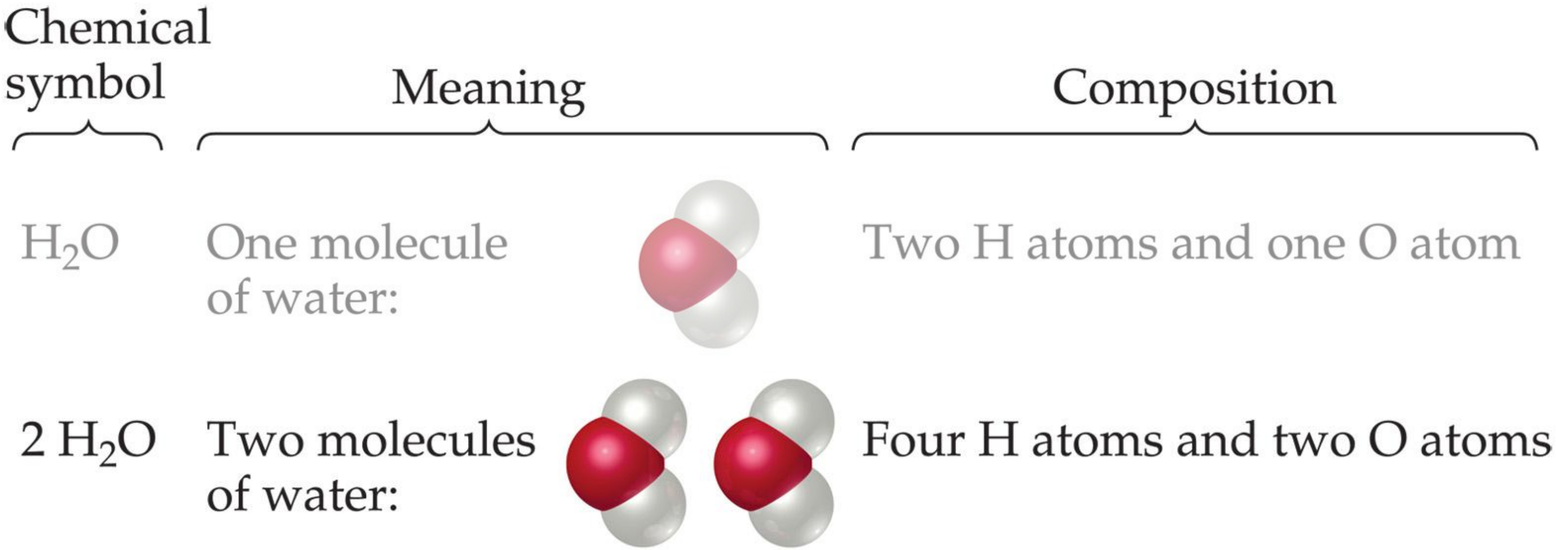
Balancing Equations

- When balancing a chemical reaction you may add coefficients in front of the compounds to balance the reaction, but

  you may **not** change the subscripts.

  - Changing the subscripts changes the compound. Subscripts are determined by the valence electrons (charges for ionic or sharing for covalent)

Subscripts and Coefficients Give Different Information



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H_2O	One molecule of water:	Two H atoms and one O atom
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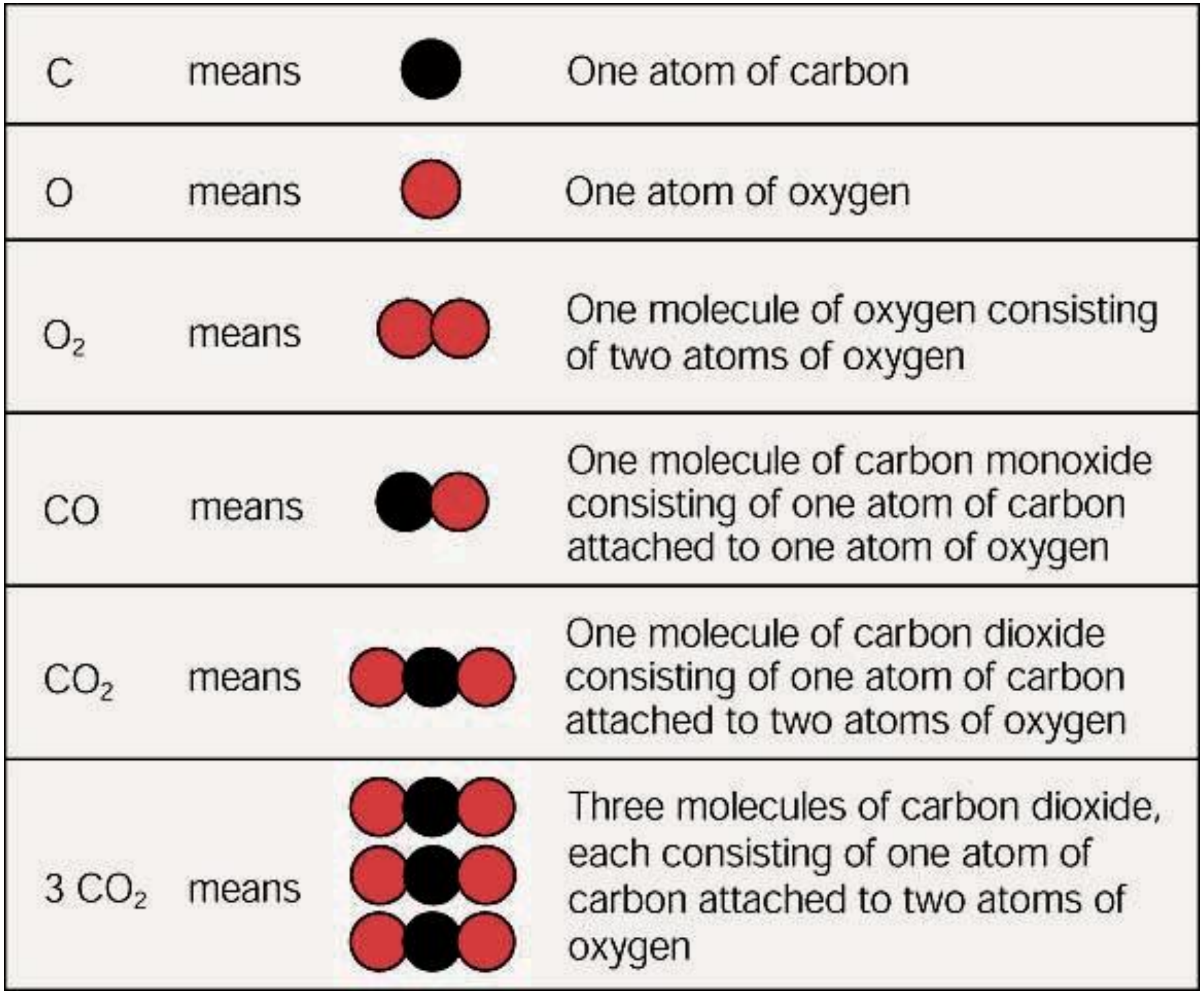
- Subscripts tell the number of atoms of each element in a molecule
- Coefficients tell the number of molecules

Subscripts vs. Coefficients



C	means	⚫	One atom of carbon
O	means	🔴	One atom of oxygen
O_2	means	🔴🔴	One molecule of oxygen consisting of two atoms of oxygen
CO	means	⚫🔴	One molecule of carbon monoxide consisting of one atom of carbon attached to one atom of oxygen
CO_2	means	🔴⚫🔴	One molecule of carbon dioxide consisting of one atom of carbon attached to two atoms of oxygen
3 CO_2	means	🔴⚫🔴 🔴⚫🔴 🔴⚫🔴	Three molecules of carbon dioxide, each consisting of one atom of carbon attached to two atoms of oxygen

The subscripts tell you how many atoms of a particular element are in a compound. The coefficient tells you about the quantity, or number, of molecules of the compound.

Examples:

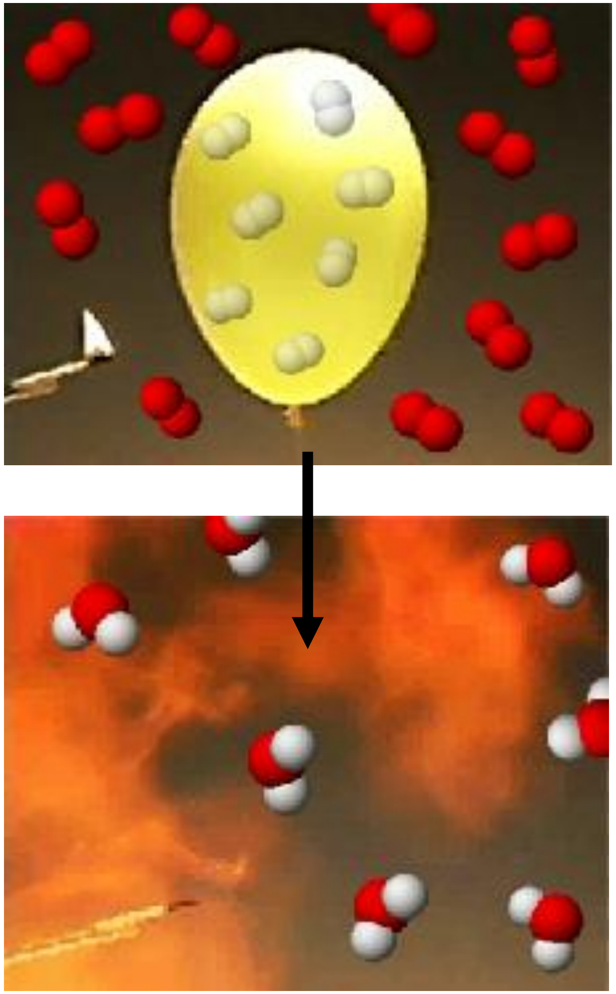
- Subscripts tell the number of atoms of each element in a molecule
- Coefficients tell the number of molecules

Balancing Equations

$$\underline{\ \ 2\ \ }\ H_2(g)\ +\ \underline{\qquad}\ O_2(g) \rightarrow \underline{\ \ 2\ \ }\ H_2O(l)$$

What Happened to the Other Oxygen Atom?????

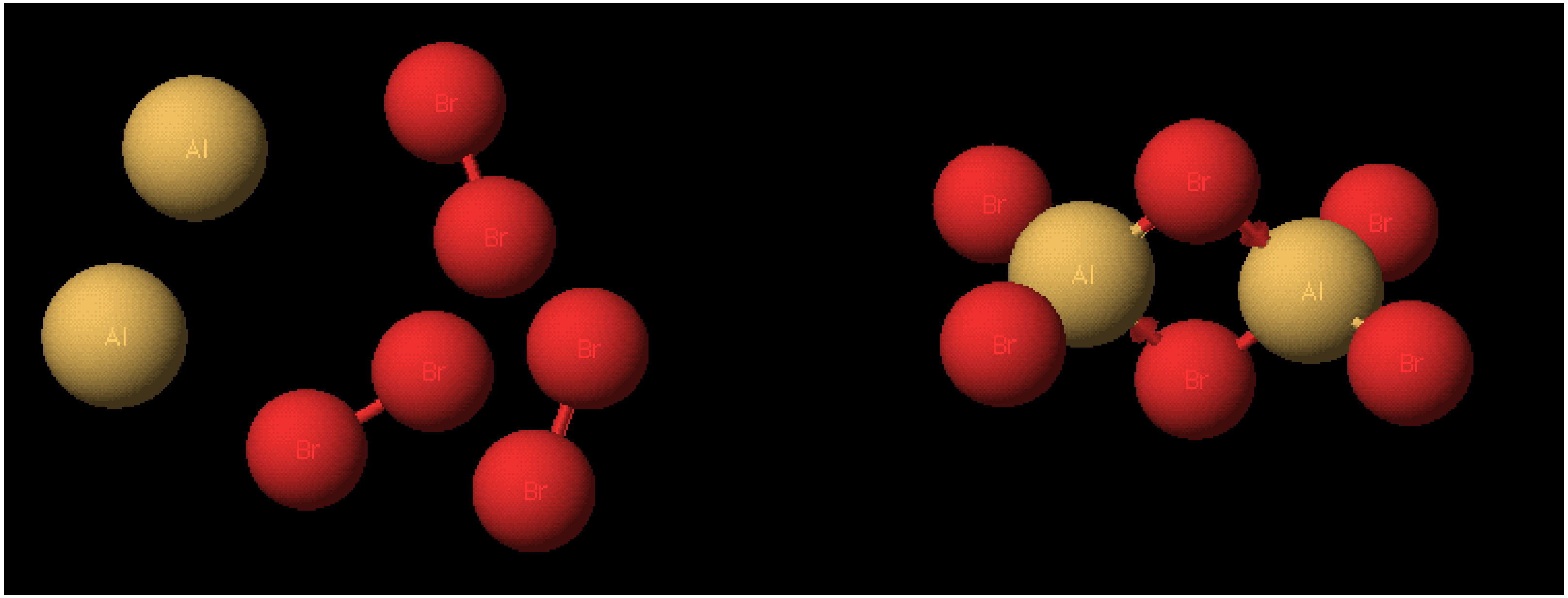
This equation is not balanced!

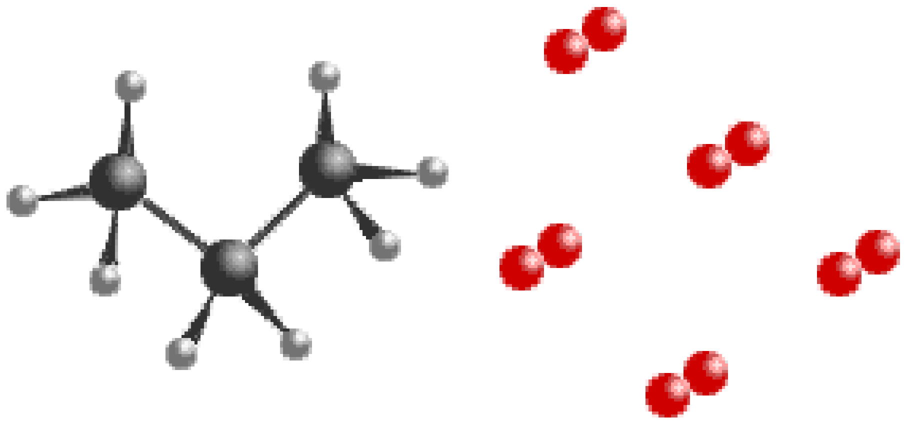


Two hydrogen atoms from a hydrogen molecule (H_2) combines with one of the oxygen atoms from an oxygen molecule (O_2) to form H_2O. Then, the remaining oxygen atom combines with two more hydrogen atoms (from another H_2 molecule) to make a second H_2O molecule.

Balancing Equations

$\underline{\ 2\ }\ Al(s)\ +\ \underline{\ 3\ }\ Br_2(l) \rightarrow \underline{\quad\quad}\ Al_2Br_6(s)$





Balancing Equations

_____$C_3H_8(g)$ + ______ $O_2(g)$ →

______$CO_2(g)$ + ______ $H_2O(g)$

_____$B_4H_{10}(g)$ + ______ $O_2(g)$ →

___ $B_2O_3(g)$ + ______ $H_2O(g)$

Balancing Equations

Sodium phosphate + iron (III) oxide → sodium oxide + iron (III) phosphate

Na_3PO_4 + Fe_2O_3 ---->

Na_2O + $FePO_4$

Steps to Balancing Equations

There are four basic steps to balancing a chemical equation.

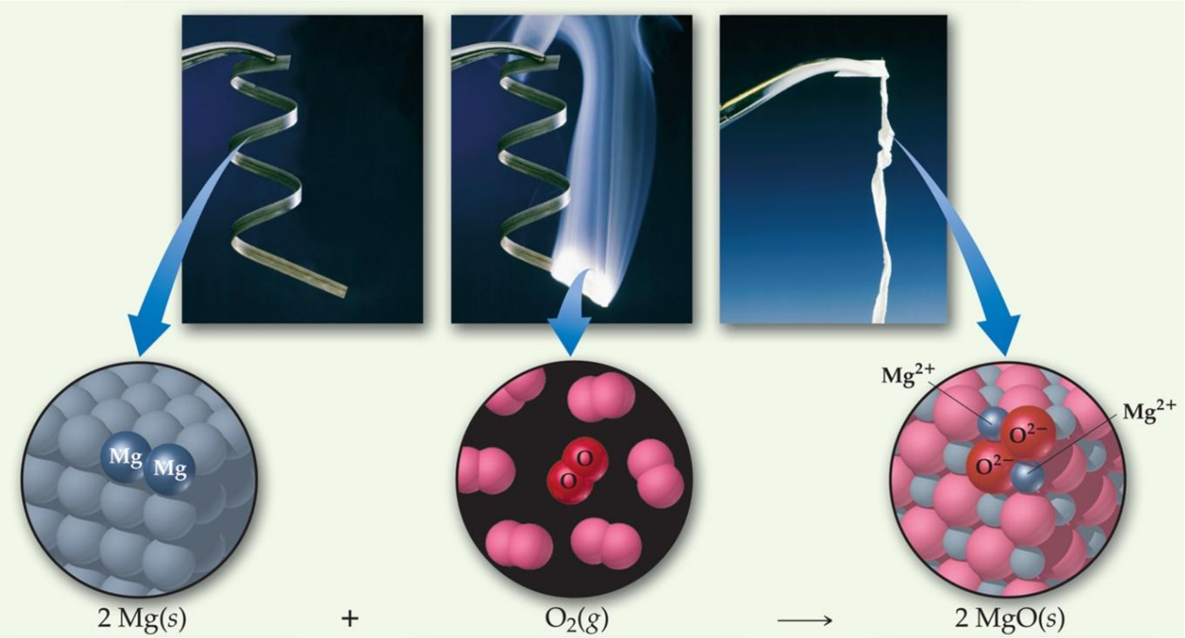
1. **Write the correct formula for the reactants and the products. DO NOT TRY TO BALANCE IT YET! You must write the correct formulas first. And most importantly, once you write them correctly DO NOT CHANGE THE FORMULAS!**
2. **Find the number of atoms for each element on the left side. Compare those against the number of the atoms of the same element on the right side.**
3. **Determine where to place coefficients in front of formulas so that the left side has the same number of atoms as the right side for EACH element in order to balance the equation.**
4. **Check your answer to see if:**
    - **The numbers of atoms on both sides of the equation are now balanced.**
    - **The coefficients are in the lowest possible whole number ratios. (reduced)**

Reaction Types

Classifying Chemical Changes

- There are four general types
  - Single displacement – a reaction where one element displaces another in a compound.
    $Cl_{2(g)} + 2KBr_{(aq)} \rightarrow 2\ KCl_{(aq)} + Br_{2(g)}$
  - Double displacement - a reaction where positive and negative portions are interchanged
    $PbCl_{2(cr)} + LiSO_{4(aq)} \rightarrow 2LiCl_{(aq)} + PbSO_{4(cr)}$
  - Decomposition - where more complex substance breaks down into simpler substances.
    $CdCO_3 \rightarrow CdO + CO_2$
  - Synthesis - two or more substances combine to form 1 new substance $NH_{3(g)} + HCl_{(g)} \rightarrow NH_4Cl_{(cr)}$

Synthesis Reactions



- Two or more substances react to form one product

- Examples:

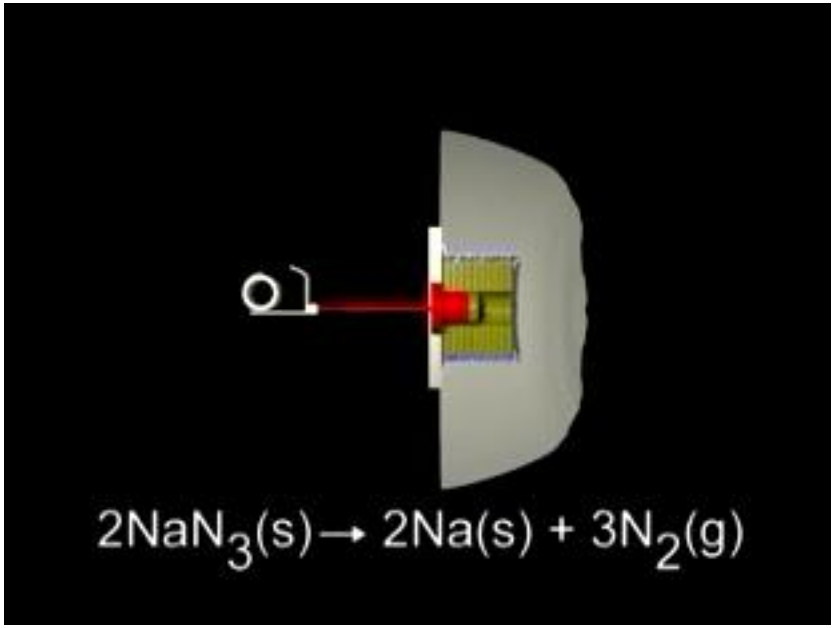
$$N_{2\,(g)} + 3\,H_{2\,(g)} \longrightarrow 2\,NH_{3\,(g)}$$

$$C_3H_{6\,(g)} + Br_{2\,(l)} \longrightarrow C_3H_6Br_{2\,(l)}$$

$$2\,Mg_{\,(s)} + O_{2\,(g)} \longrightarrow 2\,MgO_{\,(s)}$$

$$2\ \mathrm{Mg}_{(s)} + \mathrm{O}_{2\,(g)} \longrightarrow 2\ \mathrm{MgO}_{(s)}$$

Decomposition Reactions



- One substance breaks down into two or more substances

- Examples:

$CaCO_{3\,(s)} \longrightarrow CaO_{\,(s)} + CO_{2\,(g)}$

$2\ KClO_{3\,(s)} \longrightarrow 2\ KCl_{\,(s)} + O_{2\,(g)}$

$2\ NaN_{3\,(s)} \longrightarrow 2\ Na_{\,(s)} + 3\ N_{2\,(g)}$

Combustion Reactions

- Rapid reactions that produce a flame
- Most often involve hydrocarbons reacting with oxygen in the air

- Examples:

$$CH_{4\,(g)} + 2\ O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\ H_2O_{\,(g)}$$

$$C_3H_{8\,(g)} + 5\ O_{2\,(g)} \longrightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{\,(g)}$$

Some Suggestions to Help You

Some Helpful Hints for balancing equations:

- **Take one element at a time, working left to right except for H and O. Save H for next to last, and O until last.**
- **IF everything balances except for O, and there is no way to balance O with a whole number, double all the coefficients and try again. (Because O is diatomic as an element)**
- **(Shortcut) Polyatomic ions that appear on both sides of the equation should be balanced as independent unit**

Chemical Reactions Trends

- When two salts combine it is a double displacement reaction.
  Example: $NaNO_3 + AgCl \rightarrow NaCl + AgNO_3$
- Metals and nonmetals combine to form a salt.
  Example: $2Ca + O_2 \rightarrow 2CaO$
- When highly reactive metals like alkaline metal combine with water the form a base and hydrogen gas.
  Example: $Na + H_2O \rightarrow NaOH + H_2$
- When a metal oxide is put in water it forms a base.
  Example: $CaO + H_2O \rightarrow Ca(OH)_2$
- When a nonmetal oxide is put in water it forms an acid.
  Example: $CO_2 + H_2O \rightarrow H_2CO_3$
- When an organic compound is burned it yields CO_2 & H_2O.
  $C_3H_6 + 6O_2 \rightarrow 3CO_2 + 6H_2O$

More Chemical Trends

- When a metal and an acid combine a salt and hydrogen gas is formed.
  Example: $2\ Zn + 2\ HCl \rightarrow 2\ ZnCl + H_2$
- When an acid and a base are put together the result is salt and water.
  Example: $HCl + NaOH \rightarrow NaCl + H_2O$
- Hydroxides break down into metal oxides and water.
  Example: $Cu(OH)_2 \rightarrow CuO + H_2O$

Stoichiometry

Stiochiometry & Chocolate Chip Cookies!!

1 cup butter

1/2 cup white sugar

1 cup packed brown sugar

1 teaspoon vanilla extract

2 eggs

2 1/2 cups all-purpose flour

1 teaspoon baking soda

1 teaspoon salt

2 cups semisweet chocolate chips

Makes 3 dozen

How many eggs are needed to make 3 dozen cookies?

How much butter is needed for the amount of chocolate chips used?

How many eggs would we need to make 9 dozen cookies?

How much brown sugar would I need if I had 1 ½ cups white sugar?

Cookies and Chemistry…Huh!?!?

- Just like chocolate chip cookies have recipes, chemists have recipes as well
- Instead of calling them recipes, we call them reaction equations
- Furthermore, instead of using cups and teaspoons, we use moles
- Lastly, instead of eggs, butter, sugar, etc. we use chemical compounds as ingredients

Chemistry Recipes & Stoichiometry

- Looking at a reaction tells us how much of something you need to react with something else to get a product (like the cookie recipe)
- Be sure you have a balanced reaction before you start!
  - Example: $2\ Na + Cl_2 \rightarrow 2\ NaCl$
  - This reaction tells us that by mixing 2 moles of sodium with 1 mole of chlorine we will get 2 moles of sodium chloride
  - What if we wanted 4 moles of NaCl? 10 moles? 50 moles?

Stoichiometry

- Stoichiometry uses the mass to mass relationships or mole to mole relationships to solve chemistry problems
- Steps to solving stoichiometry problems
  - Write the balanced equation
  - Find mole to mole ration
  - Convert grams to moles
  - Do mole relationship
  - Change moles back to grams

Stoichiometry

- Example: How many grams of silver chloride can be produced from the reaction of 17.0g of silver nitrate with excess sodium Chloride.
  - $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
  - Mole ratio of silver chloride to silver nitrate is 1:1
  - $17.0\text{ g } AgNO_3 \times \frac{1\text{mole } AgNO_3}{170\text{ g } AgNO_3} \times \frac{1\text{mole } AgCl}{1\text{mole } AgNO_3} \times \frac{144\text{ g}AgCl}{1\text{ mole } AgCl}$ $= 14.4\text{g } AgCl$
  - (Grams given) →(Grams to moles) →(mole ratio) →(Moles to grams) →(grams required)

Mole-Mass Conversions

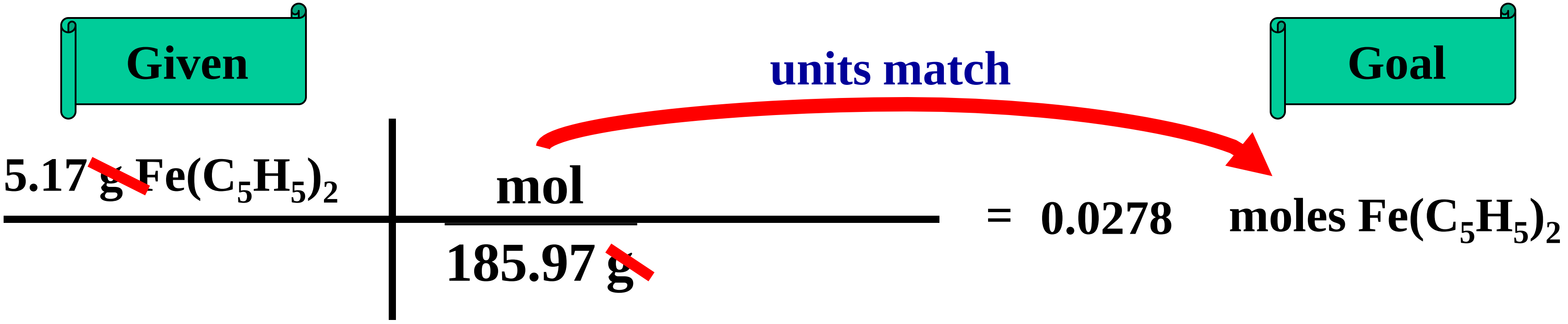
- Most of the time in chemistry, the amounts are given in grams instead of moles
- We still go through moles and use the mole ratio, but now we also use molar mass to get to grams
    - Example: How many grams of chlorine are required to react completely with 5.00 moles of sodium to produce sodium chloride?

$$2\ \text{Na} + \text{Cl}_2 \rightarrow 2\ \text{NaCl}$$

$$\frac{5.00\ \cancel{\text{moles}}\ \text{Na}}{} \bigg| \frac{1\ \cancel{\text{mol Cl}_2}}{2\ \cancel{\text{mol Na}}} \bigg| \frac{70.90\text{g Cl}_2}{1\ \cancel{\text{mol Cl}_2}} = 177\text{g Cl}_2$$

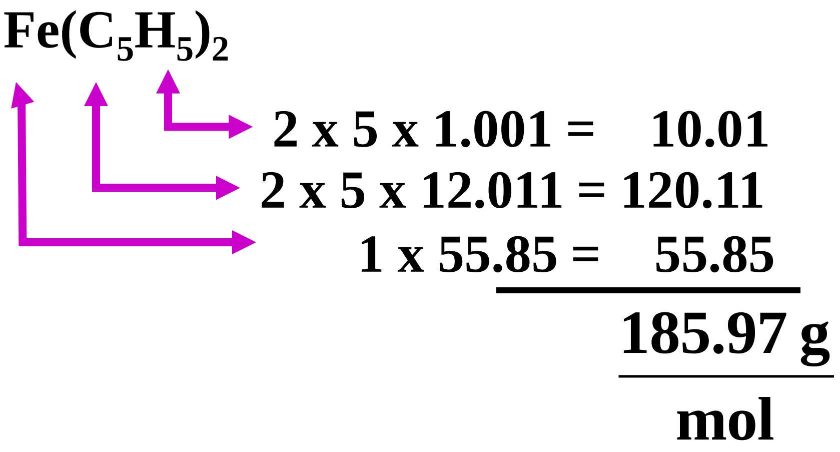
Converting grams to moles.

Determine how many moles there are in 5.17 grams of $Fe(C_5H_5)_2$.



$$\left. \frac{5.17\ \cancel{\text{g}}\ \text{Fe}(\text{C}_5\text{H}_5)_2}{} \right| \frac{\text{mol}}{185.97\ \cancel{\text{g}}} = 0.0278 \quad \text{moles Fe}(\text{C}_5\text{H}_5)_2$$

Use the molar mass to convert grams to moles.



$Fe(C_5H_5)_2$

2 x 5 x 1.001 = 10.01

2 x 5 x 12.011 = 120.11

1 x 55.85 = 55.85

$$\frac{185.97\ \text{g}}{\text{mol}}$$

Stoichiometry (more working with ratios)

Ratios are found within a chemical equation.

$$2HCl + 1Ba(OH)_2 \rightarrow 2H_2O + 1\,BaCl_2$$

coefficients give MOLAR RATIOS

2 moles of HCl react with 1 mole of $Ba(OH)_2$ to form 2 moles of H_2O and 1 mole of $BaCl_2$

Mole – Mole Conversions

When N_2O_5 is heated, it decomposes:

$$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$$

a. How many moles of NO_2 can be produced from 4.3 moles of N_2O_5?

$$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$$

4.3 mol ? mol Units match

$$\begin{array}{c|c} 4.3\ \cancel{\text{mol } N_2O_5} & 4\text{mol } NO_2 \\ \hline & 2\cancel{\text{mol } N_2O_5} \end{array} = 8.6 \text{ moles } NO_2$$

b. How many moles of O_2 can be produced from 4.3 moles of N_2O_5?

$$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$$

4.3 mol ? mol

$$\begin{array}{c|c} 4.3\ \cancel{\text{mol } N_2O_5} & 1\text{mol } O_2 \\ \hline & 2\cancel{\text{mol } N_2O_5} \end{array} = 2.\underline{2} \text{ mole } O_2$$

gram ↔ mole and gram ↔ gram conversions

When N_2O_5 is heated, it decomposes:

$$2N_2O_5(g) \longrightarrow 4NO_2(g) + O_2(g)$$

a. How many moles of N_2O_5 were used if 21$\underline{0}$g of NO_2 were produced?

$$\underset{\text{? moles}}{2N_2O_5(g)} \rightarrow \underset{21\underline{0}\text{g}}{4NO_2(g)} + O_2(g)$$

Units match

$$\frac{21\underline{0}\ \cancel{\text{g NO}_2}}{} \left| \frac{\cancel{\text{mol NO}_2}}{46.0\cancel{\text{g NO}_2}} \right| \frac{2\text{mol N}_2\text{O}_5}{4\cancel{\text{mol NO}_2}} = 2.2\underline{8} \text{ moles N}_2\text{O}_5$$

b. How many grams of N_2O_5 are needed to produce 75.0 grams of O_2?

$$\underset{\text{? grams}}{2N_2O_5(g)} \rightarrow 4NO_2(g) + \underset{75.0\text{ g}}{O_2(g)}$$

$$\frac{75.0\ \cancel{\text{g O}_2}}{} \left| \frac{\cancel{\text{mol O}_2}}{32.0\ \cancel{\text{g O}_2}} \right| \frac{\cancel{2\text{mol N}_2\text{O}_5}}{\cancel{1\text{mol O}_2}} \left| \frac{108\text{g N}_2\text{O}_5}{\cancel{\text{mol N}_2\text{O}_5}} \right. = 50\underline{6} \text{ grams N}_2\text{O}_5$$

Gram to Gram Conversions

Aluminum is an active metal that when placed in hydrochloric acid produces hydrogen gas and aluminum chloride. How many grams of aluminum chloride can be produced when 3.45 grams of aluminum are reacted with an excess of hydrochloric acid?

$$2\,Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

First write a balanced equation.

Gram to Gram Conversions

Aluminum is an active metal that when placed in hydrochloric acid produces hydrogen gas and aluminum chloride. How many grams of aluminum chloride can be produced when 3.45 grams of aluminum are reacted with an excess of hydrochloric acid?

$$2\,Al(s) \; + \; 6HCl(aq) \rightarrow 2AlCl_3(aq) \; + \; 3H_2(g)$$

3.45 g **? grams**

Now let's get organized. Write the information below the substances.

gram to gram conversions

Aluminum is an active metal that when placed in hydrochloric acid produces hydrogen gas and aluminum chloride. How many grams of aluminum chloride can be produced when 3.45 grams of aluminum are reacted with an excess of hydrochloric acid?

$$2\,\text{Al(s)} \ + \ 6\text{HCl(aq)} \rightarrow 2\text{AlCl}_3\text{(aq)} \ + \ 3\text{H}_2\text{(g)}$$

3.45 g **? grams**

Units match

$$3.45\ \cancel{\text{g Al}} \ \Bigg| \ \frac{\cancel{\text{mol Al}}}{27.0\ \cancel{\text{g Al}}} \ \Bigg| \ \frac{2\ \cancel{\text{mol AlCl}_3}}{2\ \cancel{\text{mol Al}}} \ \Bigg| \ \frac{133.3\ \text{g AlCl}_3}{\cancel{\text{mol AlCl}_3}} = 17.0\ \text{g AlCl}_3$$

Now use the molar mass to convert to grams.

Solution Stoichiometry

50.0 mL of 6.0 M H_2SO_4 (battery acid) were spilled and solid $NaHCO_3$ (baking soda) is to be used to neutralize the acid. How many grams of $NaHCO_3$ must be used?

$$H_2SO_4(aq) + 2NaHCO_3 \rightarrow 2H_2O(l) + Na_2SO_4(aq) + 2CO_2(g)$$

Solution Stoichiometry

50.0 mL of 6.0 M H_2SO_4 (battery acid) were spilled and solid $NaHCO_3$ (baking soda) is to be used to neutralize the acid. How many grams of $NaHCO_3$ must be used?

$$H_2SO_4(aq) + 2NaHCO_3 \rightarrow 2H_2O(l) + Na_2SO_4(aq) + 2CO_2(g)$$

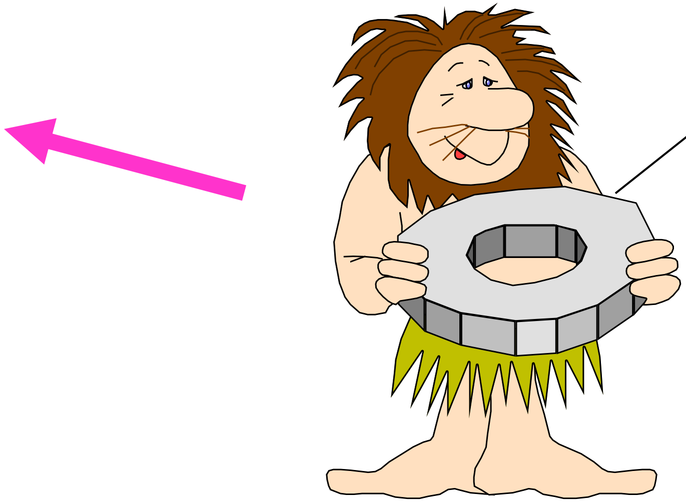
50.0 mL

? g

Our Goal

6.0 M

$$= \frac{6.0\ \text{mol}}{\text{L}}$$



Look!
A conversion factor!

Solution Stoichiometry

50.0 mL of 6.0 M H_2SO_4 (battery acid) were spilled and solid $NaHCO_3$ (baking soda) is to be used to neutralize the acid. How many grams of $NaHCO_3$ must be used?

$$H_2SO_4(aq) + 2NaHCO_3 \rightarrow 2H_2O(l) + Na_2SO_4(aq) + 2CO_2(g)$$

50.0 mL

? g

Our Goal

6.0 M

$\parallel$

$$\frac{6.0\ \text{mol}}{\text{L}}$$

$$\frac{\begin{array}{c}\cancel{H_2SO_4}\\ 50.0\ \cancel{\text{mL}}\end{array}}{\ } \left| \frac{6.0\ \cancel{\text{mol}\ H_2SO_4}}{\begin{array}{c}1000\ \cancel{\text{mL}}\\ \cancel{H_2SO_4}\end{array}} \right| \frac{\begin{array}{c}\cancel{NaHCO_3}\\ 2\ \cancel{\text{mol}}\end{array}}{\begin{array}{c}1\ \cancel{\text{mol}}\\ \cancel{H_2SO_4}\end{array}} \left| \frac{\begin{array}{c}NaHCO_3\\ 84.0\ \text{g}\end{array}}{\begin{array}{c}\cancel{\text{mol}}\\ \cancel{NaHCO_3}\end{array}} \right. = 5\underline{0}.4\ \text{g}\ NaHCO_3$$

Mass-Mass Conversions

- Most often we are given a starting mass and want to find out the mass of a product we will get (called theoretical yield) or how much of another reactant we need to completely react with it (no leftover ingredients!)
- Now we must go from grams to moles, mole ratio, and back to grams of compound we are interested in

Mass-Mass Conversion

- Ex. Calculate how many grams of ammonia are produced when you react 2.00g of nitrogen with excess hydrogen.
- $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

$$\begin{array}{c|c|c|c} 2.00\cancel{\text{g N}}_2 & 1\ \cancel{\text{mol}}\ \text{N}_2 & 2\ \cancel{\text{mol}}\ \text{NH}_3 & 17.06\text{g NH}_3 \\ \hline & 28.02\cancel{\text{g N}}_2 & 1\ \cancel{\text{mol}}\ \text{N}_2 & 1\ \cancel{\text{mol}}\ \text{NH}_3 \end{array}$$

$$= 2.4\ \text{g NH}_3$$

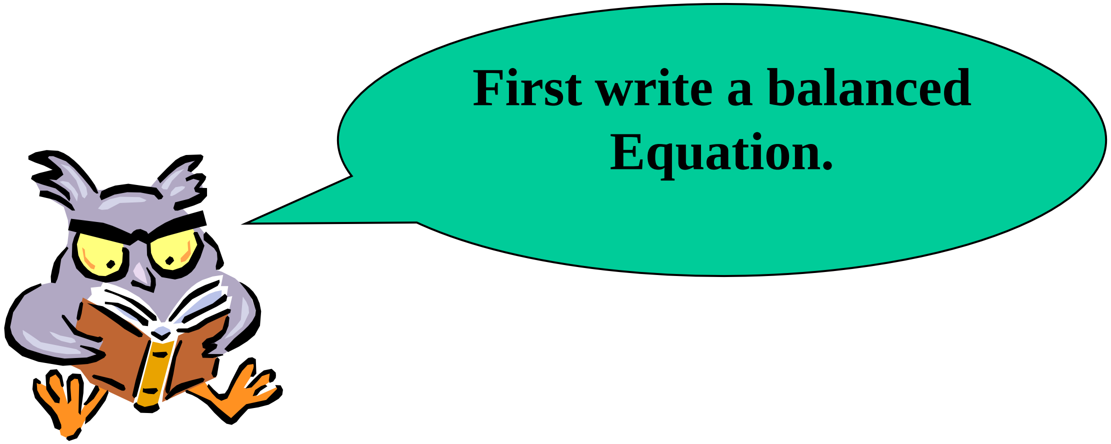
Practice

- How many grams of calcium nitride are produced when 2.00 g of calcium reacts with an excess of nitrogen?

Solution Stoichiometry:

Determine how many mL of 0.102 M NaOH solution are needed to neutralize 35.0 mL of 0.125 M H_2SO_4 solution.

$\underline{\ \ 2\ \ }\ \mathrm{NaOH} + \underline{\ \ 1\ \ }\ \mathrm{H_2SO_4} \rightarrow \underline{\ \ 2\ \ }\ \mathrm{H_2O} + \underline{\ \ 1\ \ }\ \mathrm{Na_2SO_4}$



Solution Stoichiometry:

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$$\underline{\quad 2\ }\,\text{NaOH} + \underline{\quad 1\ }\,\text{H}_2\text{SO}_4 \rightarrow \underline{\quad 2\ }\,\text{H}_2\text{O} + \underline{\quad 1\ }\,\text{Na}_2\text{SO}_4$$

$$\frac{0.102\ \cancel{\text{M}}\ \textit{mol}}{\textit{L}}$$

35.0 mL

$$\frac{0.125\text{ mol}}{\text{L}} = \frac{0.125\text{ mol}}{1000\text{ mL}}$$

? mL

Our Goal

Since 1 L = 1000 mL, we can use this to save on the number of conversions

Now, let's get organized. Place numerical Information and accompanying UNITS below each compound.

Solution Stoichiometry:

Determine how many mL of 0.102 M NaOH solution is needed to neutralize 35.0 mL of 0.125 M H_2SO_4 solution.

$$\underline{\quad 2\ }\ NaOH + \underline{\quad 1\ }\ H_2SO_4 \rightarrow \underline{\quad 2\ }\ H_2O + \underline{\quad 1\ }\ Na_2SO_4$$

$$\frac{0.102\ \cancel{M}\ mol}{L} \qquad 35.0\ mL$$

? mL

$$\frac{0.125\ mol}{L} = \frac{0.125\ mol}{1000mL}$$

shortcut

Units Match

$$\frac{35.0\ \cancel{mL}\ \cancel{H_2SO_4}}{} \left| \frac{0.125\ \cancel{mol}\ \cancel{H_2SO_4}}{1000\ \cancel{mL}\ \cancel{H_2SO_4}} \right| \frac{2\ \cancel{mol}\ \cancel{NaOH}}{1\ \cancel{mol}\ \cancel{H_2SO_4}} \left| \frac{1000\ mL\ NaOH}{0.102\ \cancel{mol\ NaOH}} \right. = 85.8\ mL\ NaOH$$

Now let's get to work converting.

Solution Stoichiometry

What volume of 0.40 M HCl solution is needed to completely neutralize 47.1 mL of 0.75 M $Ba(OH)_2$?

$2HCl(aq) + Ba(OH)_2(aq) \rightarrow 2H_2O(l) + BaCl_2$

0.40 M	47.1 mL
? mL	0.75 M

$$\begin{array}{c|c|c|c} \cancel{\text{Ba(OH)}_2\ 47.1\ \text{mL}} & \cancel{0.75\text{mol}_{\text{Ba(OH)}_2}} & \cancel{\text{HCl}}\ 2\ \cancel{\text{mol}} & \text{HCl}\ 1000\ \text{mL} \\ \hline & \cancel{1000\ \text{mL}_{\text{Ba(OH)}_2}} & 1\ \cancel{\text{mol}\ \text{Ba(OH)}_2} & 0.40\ \cancel{\text{mol}}\ \text{HCl} \end{array} = 1\underline{7}6\ \text{mL}\ \text{HCl}$$

Limiting/Excess/ Reactant and Theoretical Yield Problems :

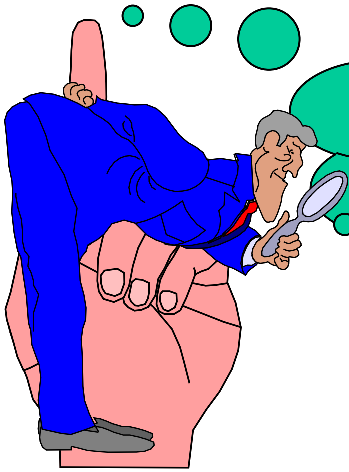
Potassium superoxide, KO_2, is used in rebreathing gas masks to generate oxygen.

$$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$$

a. How many moles of O_2 can be produced from 0.15 mol KO_2 and 0.10 mol H_2O?

b. Determine the limiting reactant.

$$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$$



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Now place
numerical the
information below
the compounds.

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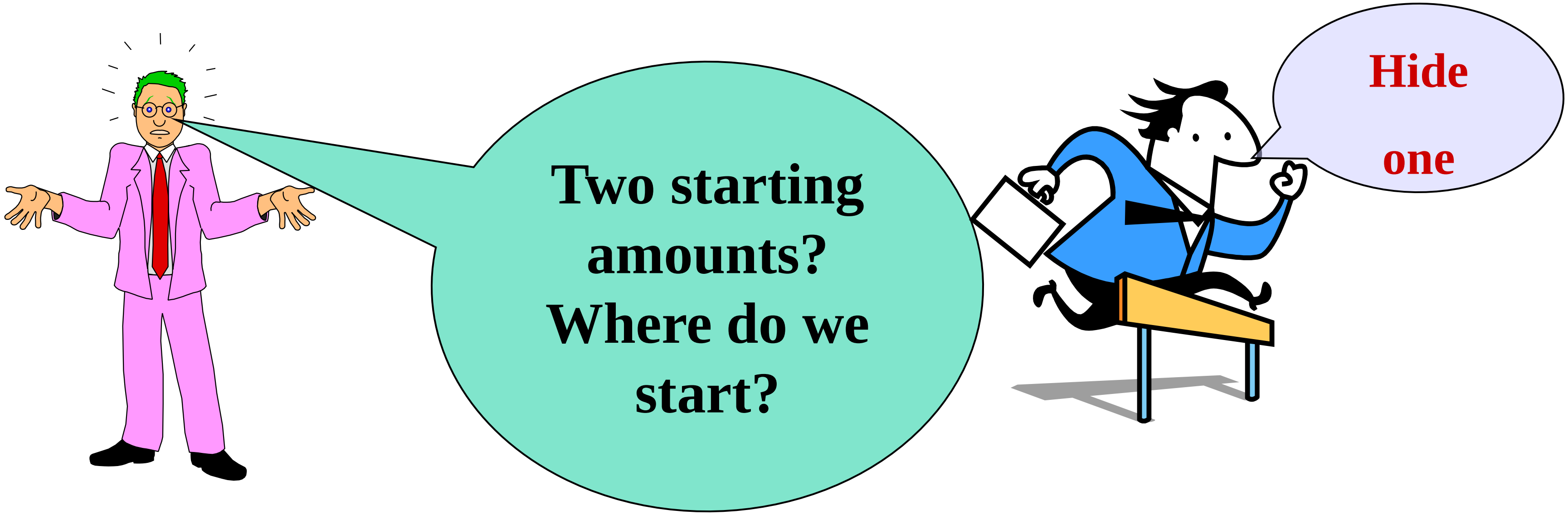
0.15 mol	0.10 mol	? moles



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0.15 mol | **Hide** | **? moles**

Based on: KO_2

$$\frac{0.15\ \cancel{\text{mol KO}_2}}{} \left| \frac{3\text{mol O}_2}{4\cancel{\text{mol KO}_2}} \right. = 0.1\underline{1}25 \quad \text{mol O}_2$$

Limiting/Excess/ Reactant and Theoretical Yield Problems :

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Hide

0.10 mol

? moles

Based on: KO_2

$$\begin{array}{l|l} 0.15\ \cancel{\text{mol } KO_2} & \dfrac{3\text{mol } O_2}{4\cancel{\text{mol } KO_2}} \end{array} = 0.1\underline{1}25 \ \text{ mol } O_2$$

Based on: H_2O

$$\begin{array}{l|l} 0.10\ \cancel{\text{mol } H_2O} & \dfrac{3\text{mol } O_2}{2\cancel{\text{mol } H_2O}} \end{array} = 0.1\underline{5}0 \ \text{ mol } O_2$$

Limiting/Excess/ Reactant and Theoretical Yield Problems :

Potassium superoxide, KO_2, is used in rebreathing gas masks to generate oxygen.

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0.15 mol **0.10 mol** **? moles**

Based on: KO_2

$$0.15\ \cancel{\text{mol } KO_2} \left|\ \frac{3\text{mol } O_2}{4\cancel{\text{mol } KO_2}}\right. = 0.1\underline{1}25 \ \text{mol } O_2$$

It was limited by the amount of KO_2.

Based on: H_2O

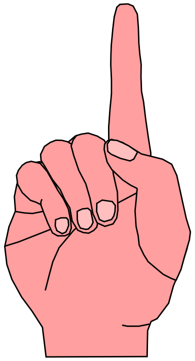
$$0.10\ \cancel{\text{mol } H_2O} \left|\ \frac{3\text{mol } O_2}{2\cancel{\text{mol } H_2O}}\right. = 0.1\underline{5}0 \ \text{mol } O_2$$

H_2O = excess (XS) reactant!

What is the theoretical yield?
Hint: Which is the smallest amount? The is based upon the limiting reactant?

Theoretical yield vs. Actual yield

Suppose the theoretical yield for an experiment was calculated to be 19.5 grams, and the experiment was performed, but only 12.3 grams of product were recovered. Determine the % yield.



Theoretical yield = 19.5 g based on limiting reactant

Actual yield = 12.3 g experimentally recovered

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$\%\ \text{yield} = \frac{12.3}{19.5} \times 100 = 63.1\%\ \text{yield}$$

Limiting/Excess Reactant Problem with % Yield

$$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$$

If a reaction vessel contains 120.0 g of KO_2 and 47.0 g of H_2O, how many grams of O_2 can be produced?

$$4KO_2(s) + \underline{2H_2O(l)} \rightarrow 4KOH(s) + 3O_2(g)$$

120.0 g

Hide one

? g

Based on: KO_2

$$\frac{120.0\ \cancel{\text{g}}\ \cancel{\text{KO}_2}}{}\ \Bigg|\ \frac{\cancel{\text{mol}}}{71.1\ \cancel{\text{g}}}\ \Bigg|\ \frac{3\ \cancel{\text{mol O}_2}}{4\ \cancel{\text{mol KO}_2}}\ \Bigg|\ \frac{32.0\text{g O}_2}{\cancel{\text{mol O}_2}} = 40.\underline{5}1\ \text{g O}_2$$

Limiting/Excess Reactant Problem with % Yield

$$4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$$

If a reaction vessel contains 120.0 g of KO_2 and 47.0 g of H_2O, how many grams of O_2 can be produced?

$$4KO_2(s) + \underset{47.0\ \text{g}}{2H_2O(l)} \rightarrow 4KOH(s) + \underset{?\ \text{g}}{3O_2(g)}$$

Hide

Based on: KO_2

$$\left.\frac{120.0\ \cancel{\text{g}}\ \cancel{KO_2}}{}\right| \frac{\cancel{\text{mol}}}{71.1\ \cancel{\text{g}}} \left| \frac{3\ \cancel{\text{mol}\ O_2}}{4\ \cancel{\text{mol}\ KO_2}} \right| \frac{32.0\text{g}\ O_2}{\cancel{\text{mol}\ O_2}} = 40.\underline{5}1\ \text{g}\ O_2$$

Based on: H_2O

$$\left.\frac{47.0\ \cancel{\text{g}\ H_2O}}{}\right| \frac{\cancel{\text{mol}\ H_2O}}{18.02\ \cancel{\text{g}\ H_2O}} \left| \frac{3\ \cancel{\text{mol}\ O_2}}{2\ \cancel{\text{mol}\ H_2O}} \right| \frac{32.0\text{g}\ O_2}{\cancel{\text{mol}\ O_2}} = 12\underline{5}.3\ \text{g}\ O_2$$

Question if only 35.2 g of O_2 were recovered, what was the percent yield?

$$\frac{\text{actual}}{\text{theoretical}}\ \text{x}\,100 = \frac{35.2}{40.51}\ \text{x}\,100 = 86.9\%\ \text{yield}$$

If a reaction vessel contains 120.0 g of KO_2 and 47.0 g of H_2O, how many grams of O_2 can be produced?

$$\begin{array}{ccccccc} 4KO_2(s) & + & 2H_2O(l) & \rightarrow & 4KOH(s) & + & 3O_2(g) \\ 120.0\text{ g} & & 47.0\text{ g} & & & & ?\text{ g} \end{array}$$

Based on: KO_2

$$\begin{array}{l|l|l|l} 120.0\,\cancel{g}\,\cancel{KO_2} & \cancel{mol} & 3\,\cancel{mol\ O_2} & 32.0g\ O_2 \\ \hline & 71.1\cancel{g} & 4\,\cancel{mol\ KO_2} & \cancel{mol\ O_2} \end{array} = 40.\underline{5}1\ \text{g } O_2$$

Based on: H_2O

$$\begin{array}{l|l|l|l} 47.0\,\cancel{g\ H_2O} & \cancel{mol\ H_2O} & 3\,\cancel{mol\ O_2} & 32.0g\ O_2 \\ \hline & 18.02\,\cancel{g\ H_2O} & 2\,\cancel{mol\ H_2O} & \cancel{mol\ O_2} \end{array} = 12\underline{5}.3\ \text{g } O_2$$

Determine how many grams of Water were left over.

The Difference between the above amounts is directly <u>RELATED</u> to the <u>XS</u> H_2O.

125.3 - 40.51 = 84.79 g of O_2 that could have been formed from the <u>XS</u> water.

$$\begin{array}{l|l|l|l} 84.79\,\cancel{g\ O_2} & \cancel{mol\ O_2} & 2\,\cancel{mol\ H_2O} & 18.02\text{ g } H_2O \\ \hline & 32.0\,\cancel{g\ O_2} & 3\,\cancel{mol\ O_2} & 1\,\cancel{mol\ H_2O} \end{array} = 31.\underline{8}3\ \text{g XS } H_2O$$

Try this problem (then check your answer):

Calculate the molarity of a solution prepared by dissolving 25.6 grams of $Al(NO_3)_3$ in 455 mL of solution.

After you have worked the problem, click here to see setup answer

$$\frac{25.6\ \cancel{\text{g}}}{} \left| \frac{\text{mole}}{213\ \cancel{\text{g}}} \right| \frac{}{455\ \text{x}\ 10^{-3}\ \text{L}} = 0.264\ \frac{\text{mol}}{\text{L}}$$

Limiting Reactants

How Many Cookies Can I Make?

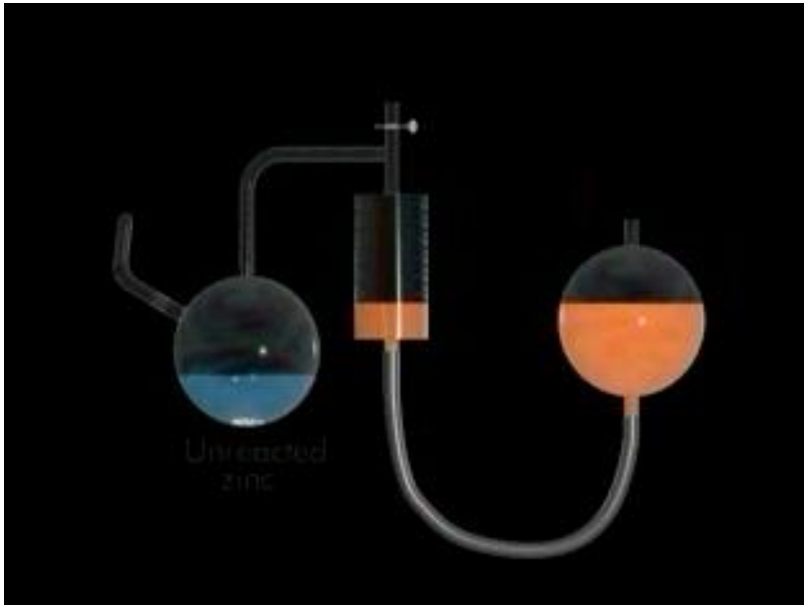
- You can make cookies until you run out of one of the ingredients
- Once this family runs out of sugar, they will stop making cookies (at least any cookies you would want to eat)

How Many Cookies Can I Make?

- In this example the sugar would be the limiting reactant, because it will limit the amount of cookies you can make

Limiting Reactants

The limiting reactant is the reactant present in the smallest stoichiometric amount

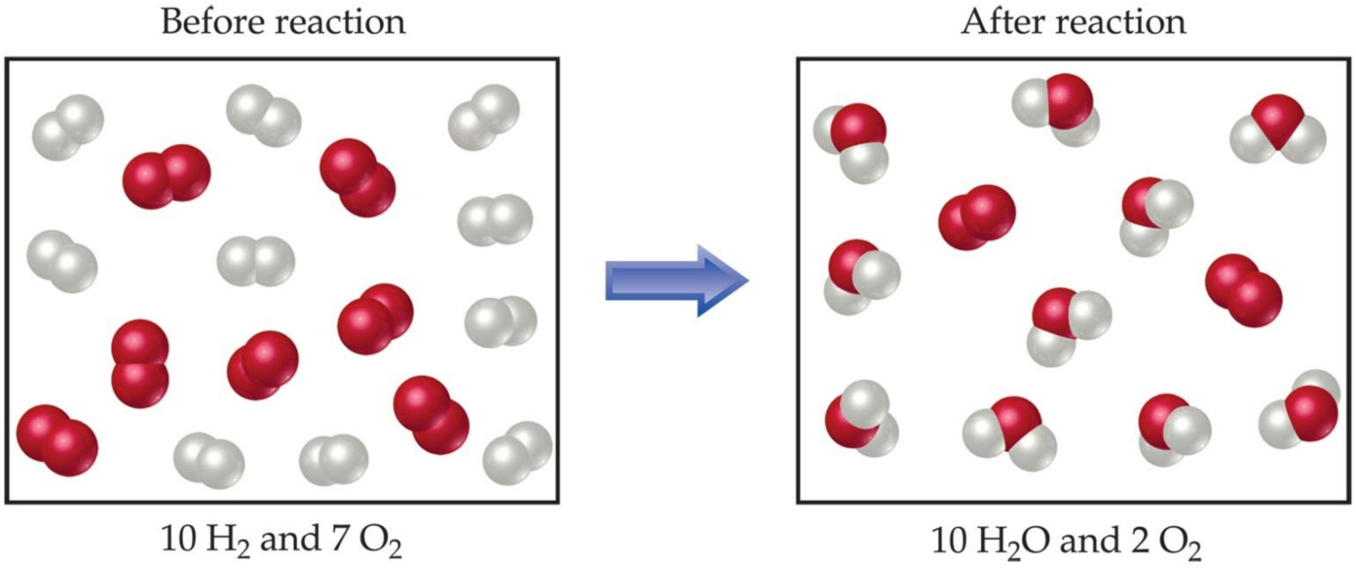


Limiting Reactant

- Most of the time in chemistry we have more of one reactant than we need to completely use up other reactant.
- That reactant is said to be in excess (there is too much).
- The other reactant limits how much product we get. Once it runs out, the reaction  s. This is called the limiting reactant.

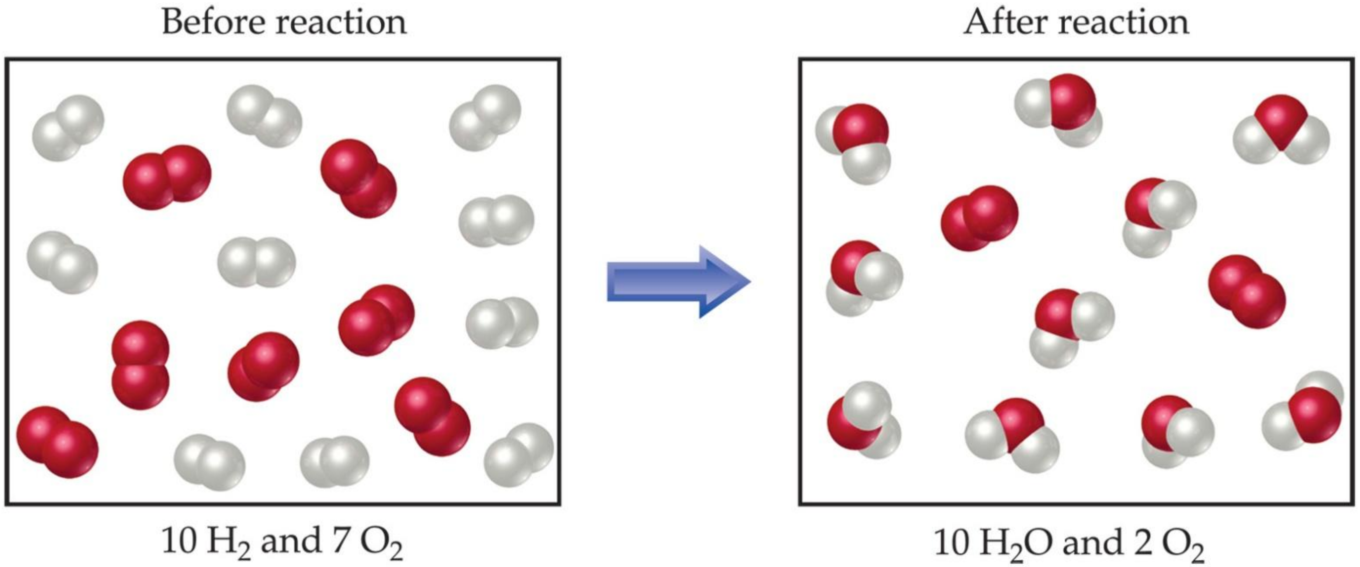
Limiting Reactants

- The limiting reactant is the reactant present in the smallest stoichiometric amount
  - In other words, it’s the reactant you’ll run out of first (in this case, the H_2)



Limiting Reactants

In the example below, the O_2 would be the excess reagent



Limiting Reactant

Limiting Reactant: Example

- 10.0g of aluminum reacts with 35.0 grams of chlorine gas to produce aluminum chloride. Which reactant is limiting, which is in excess, and how much product is produced?

$$2\ \text{Al} + 3\ \text{Cl}_2 \rightarrow 2\ \text{AlCl}_3$$

- Start with Al:

$$\begin{array}{l|l|l|l} 10.0\ \cancel{\text{g Al}} & 1\ \cancel{\text{mol Al}} & 2\ \cancel{\text{mol AlCl}_3} & 133.5\ \text{g AlCl}_3 \\ \hline & 27.0\ \cancel{\text{g Al}} & 2\ \cancel{\text{mol Al}} & 1\ \cancel{\text{mol AlCl}_3} \end{array} \quad \cancel{= 49.4\text{g AlCl}_3}$$

- Now Cl$_2$:

$$\begin{array}{l|l|l|l} 35.0\text{g}\ \cancel{\text{Cl}_2} & 1\ \cancel{\text{mol Cl}_2} & 2\ \cancel{\text{mol AlCl}_3} & 133.5\ \text{g AlCl}_3 \\ \hline & 71.0\ \cancel{\text{g Cl}_2} & 3\ \cancel{\text{mol Cl}_2} & 1\ \cancel{\text{mol AlCl}_3} \end{array} \quad = 43.9\text{g AlCl}_3$$

LR Example Continued

- We get 49.4g of aluminum chloride from the given amount of aluminum, but only 43.9g of aluminum chloride from the given amount of chlorine. Therefore, chlorine is the limiting reactant. Once the 35.0g of chlorine is used up, the reaction comes to a complete
  .

Limiting Reactant Practice

- 15.0 g of potassium reacts with 15.0 g of iodine. Calculate which reactant is limiting and how much product is made.

Finding the Amount of Excess

- By calculating the amount of the excess reactant needed to completely react with the limiting reactant, we can subtract that amount from the given amount to find the amount of excess.
- Can we find the amount of excess potassium in the previous problem?

Finding Excess Practice

- 15.0 g of potassium reacts with 15.0 g of iodine.

  $2\ K + I_2 \rightarrow 2\ KI$

- We found that Iodine is the limiting reactant, and 19.6 g of potassium iodide are produced.

$$\begin{array}{c|c|c|c} 15.0\ \cancel{\text{g I}_2} & 1\ \cancel{\text{mol I}_2} & 2\ \cancel{\text{mol K}} & 39.1\ \text{g K} \\ \hline & 254\ \cancel{\text{g I}_2} & 1\ \cancel{\text{mol I}_2} & 1\ \cancel{\text{mol K}} \end{array} = 4.62\ \text{g K}$$

USED!

15.0 g K – 4.62 g K = 10.38 g K EXCESS

Given amount of excess reactant

Amount of excess reactant actually used

Note that we started with the limiting reactant! Once you determine the LR, you should only start with it!

Limiting Reactant: Recap

1. You can recognize a limiting reactant problem because there is MORE THAN ONE GIVEN AMOUNT.
2. Convert ALL of the reactants to the SAME product (pick any product you choose.)
3. The lowest answer is the correct answer.
4. The reactant that gave you the lowest answer is the LIMITING REACTANT.
5. The other reactant(s) are in EXCESS.
6. To find the amount of excess, subtract the amount used from the given amount.
7. If you have to find more than one product, be sure to start with the limiting reactant. You don’t have to determine which is the LR over and over again!