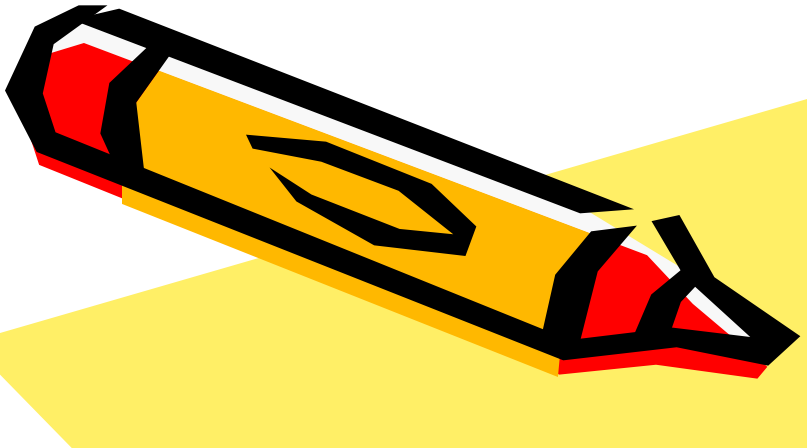
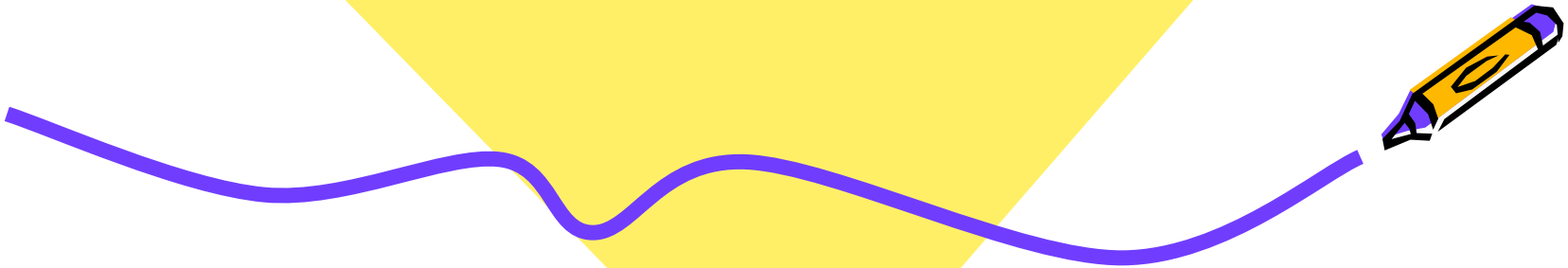
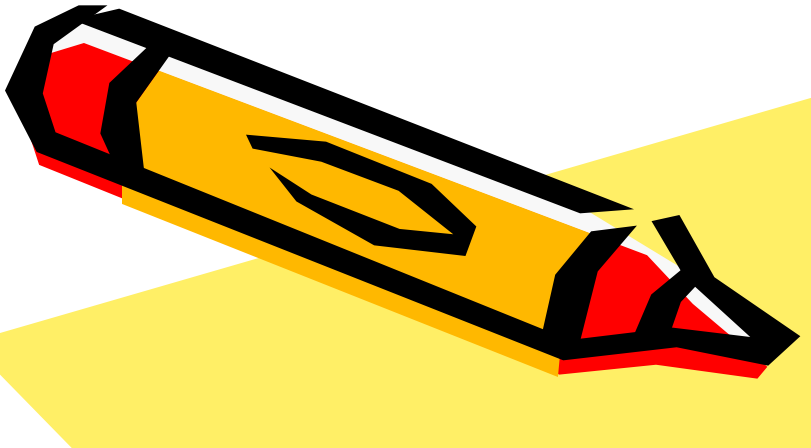
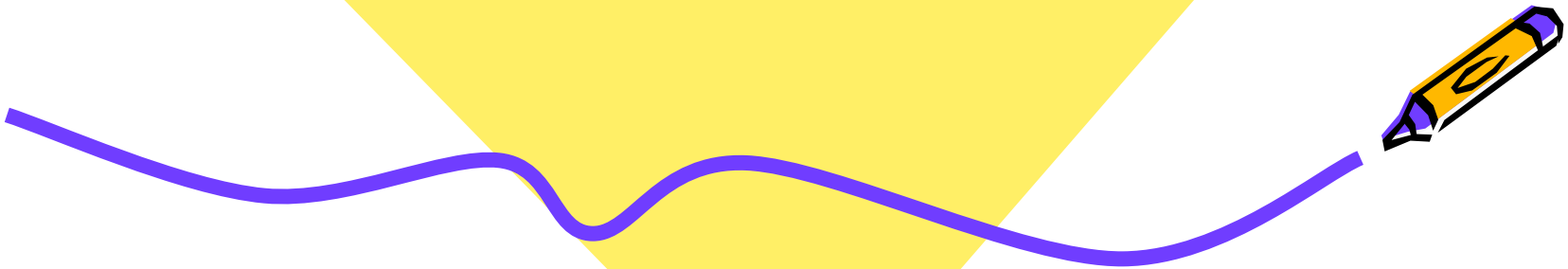


Chapter 3
Chemical Reactions and Reaction Stoichiometry





3.1 Chemical Equations



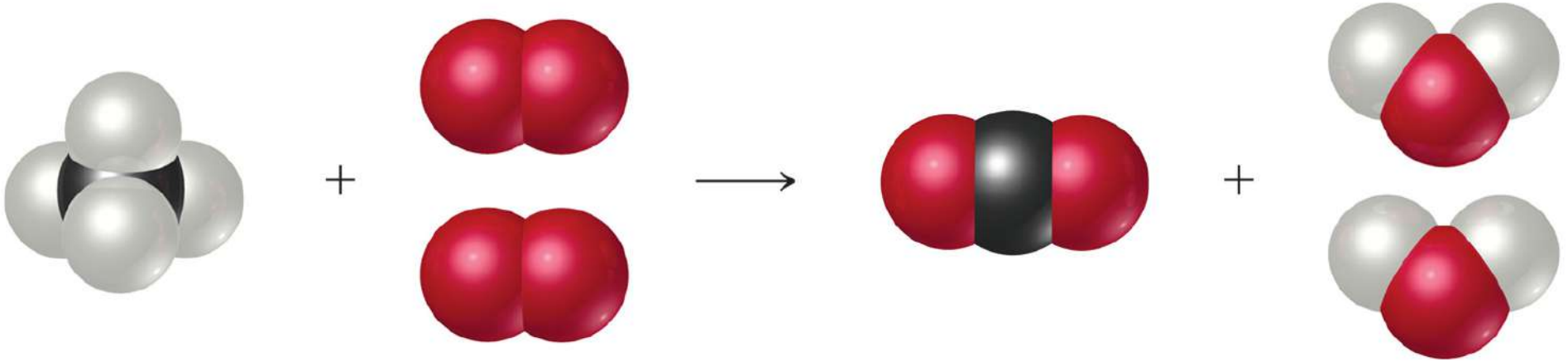
Law of Conservation of Mass

"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

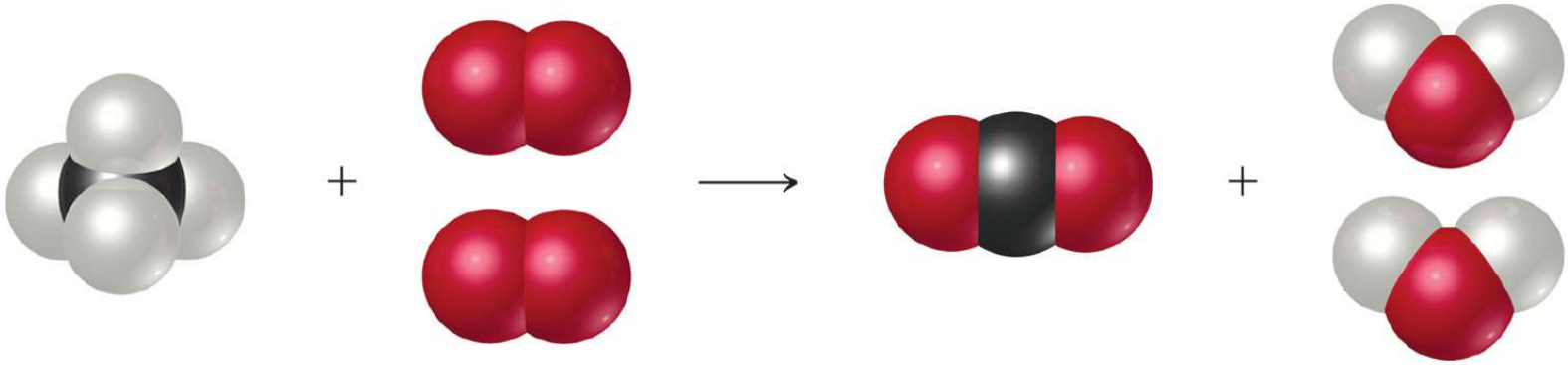
Chemical Equations

Chemical equations are concise representations of chemical reactions.



Anatomy of a Chemical Equation

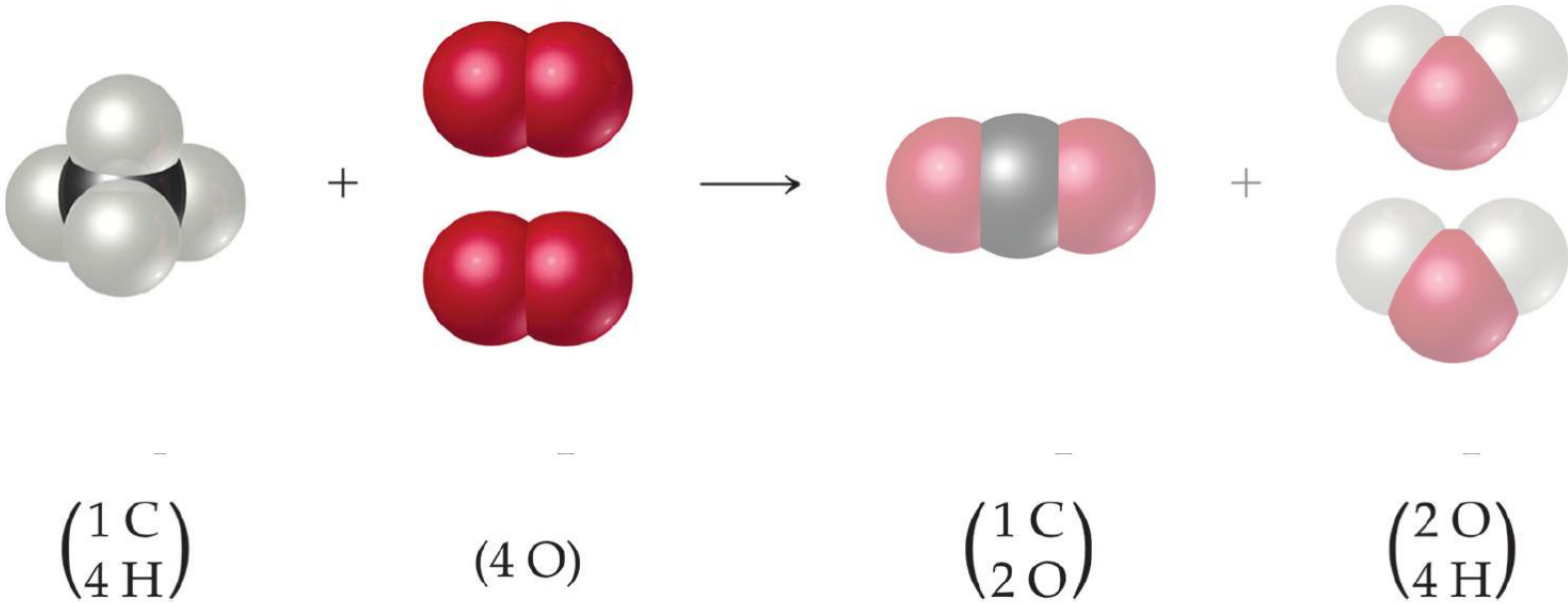
$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$



$\begin{pmatrix}1\,C\\4\,H\end{pmatrix}$ $(4\,O)$ $\begin{pmatrix}1\,C\\2\,O\end{pmatrix}$ $\begin{pmatrix}2\,O\\4\,H\end{pmatrix}$

Anatomy of a Chemical Equation

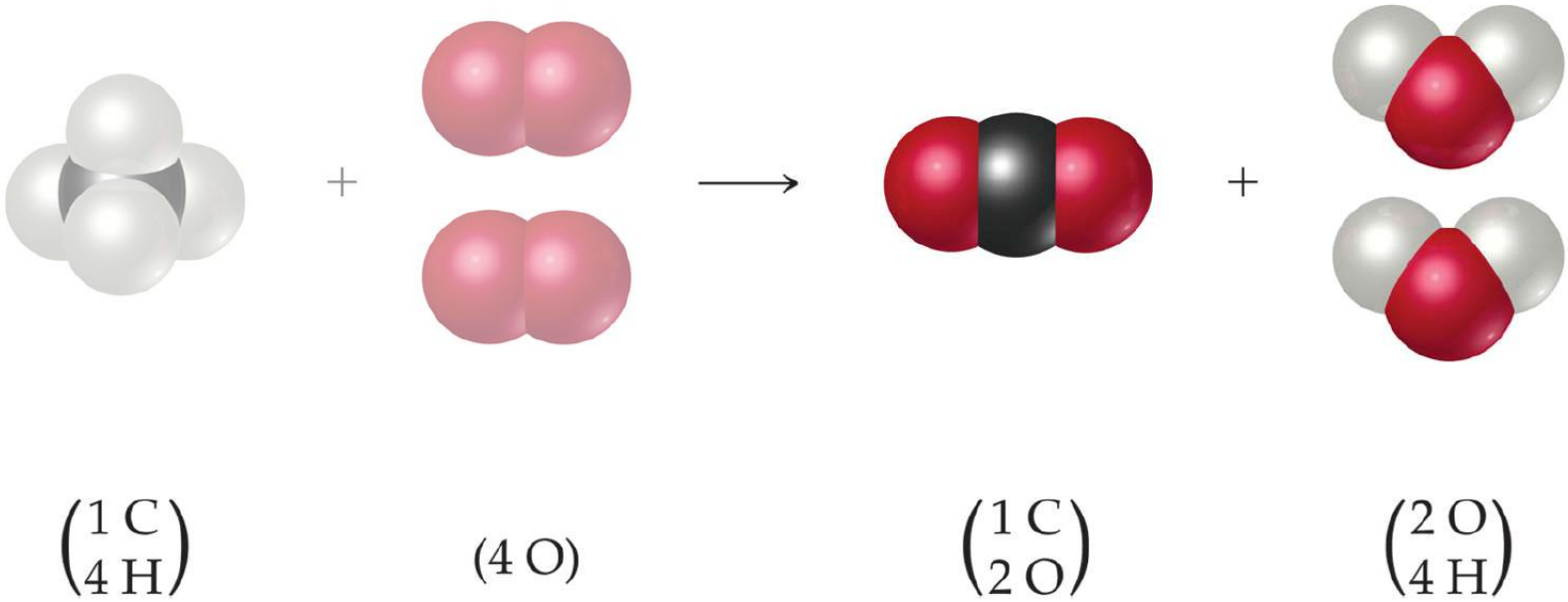
$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{(g)}$$



Reactants appear on the left side of the equation.

Anatomy of a Chemical Equation

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

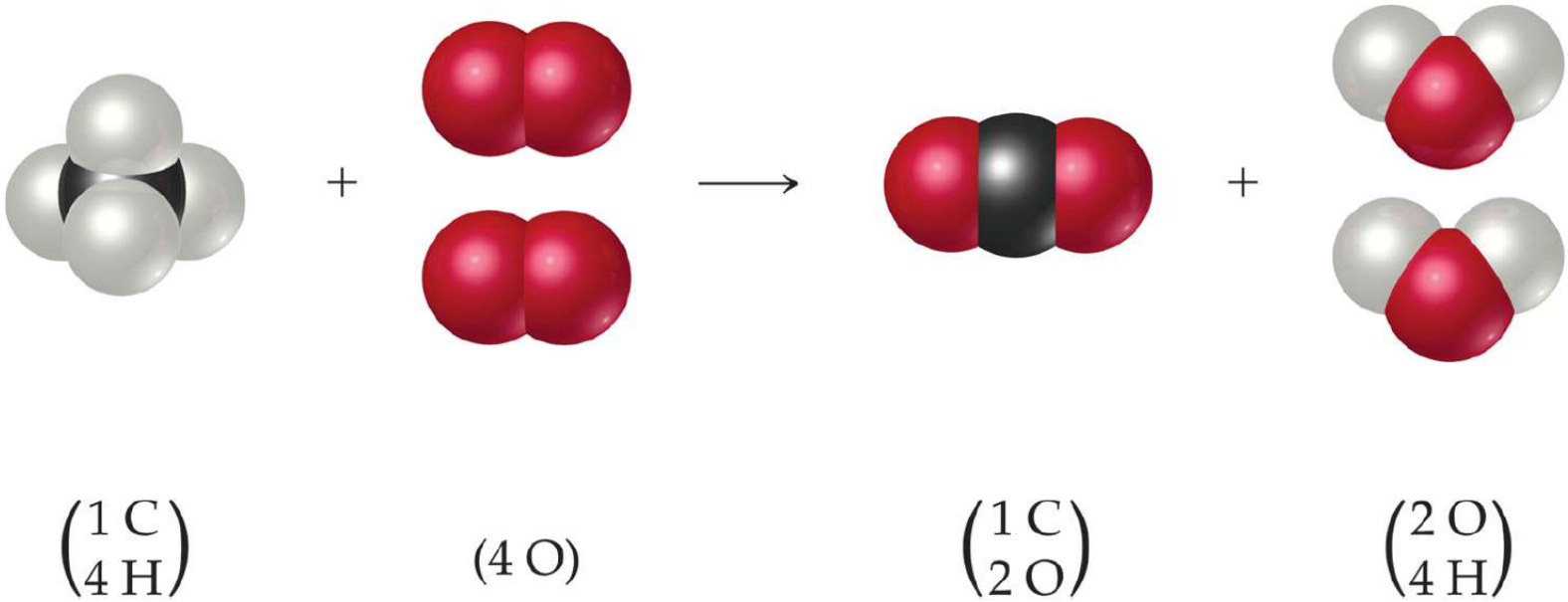


$\begin{pmatrix} 1\,C \\ 4\,H \end{pmatrix}$ $(4\,O)$ $\begin{pmatrix} 1\,C \\ 2\,O \end{pmatrix}$ $\begin{pmatrix} 2\,O \\ 4\,H \end{pmatrix}$

Products appear on the right side of the equation.

Anatomy of a Chemical Equation

$$CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

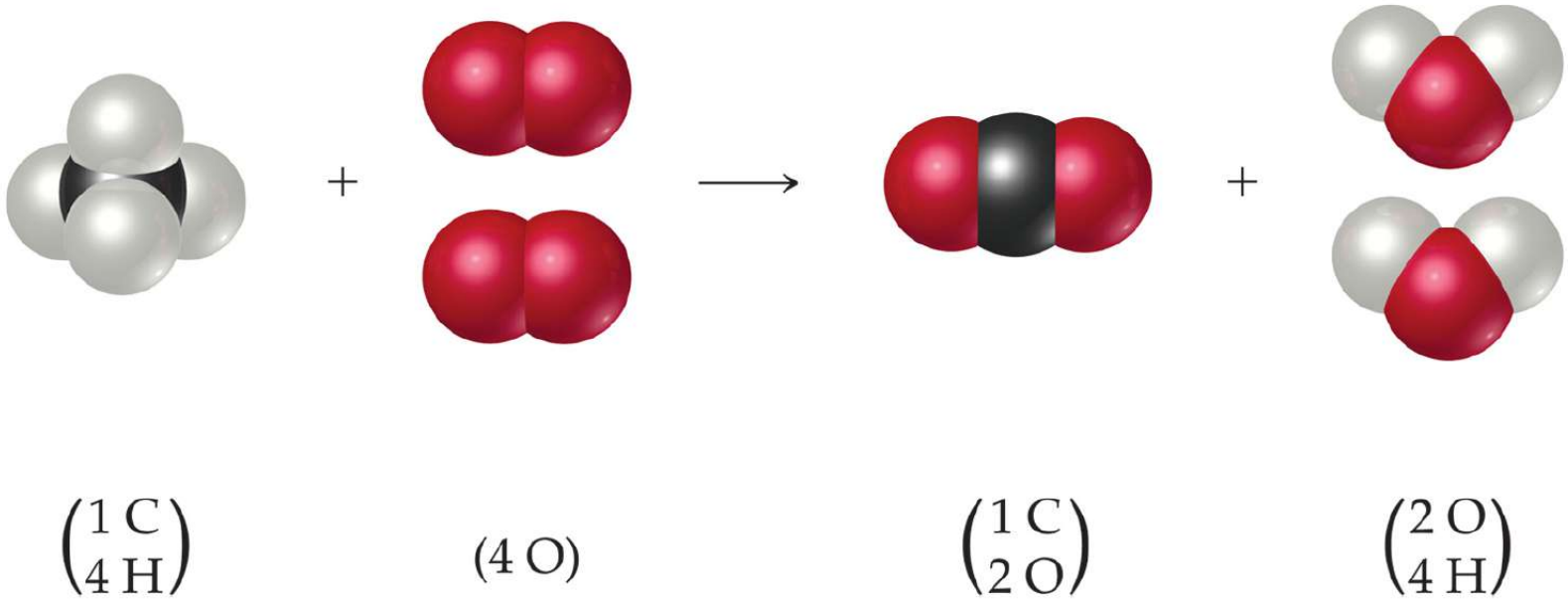


(1 C, 4 H) + (4 O) → (1 C, 2 O) + (2 O, 4 H)

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

Anatomy of a Chemical Equation

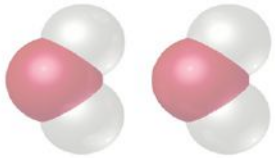
$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{(g)}$$



$\begin{pmatrix} 1\,C \\ 4\,H \end{pmatrix}$ $(4\,O)$ $\begin{pmatrix} 1\,C \\ 2\,O \end{pmatrix}$ $\begin{pmatrix} 2\,O \\ 4\,H \end{pmatrix}$

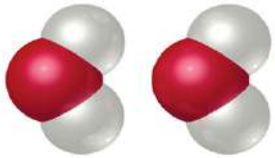
Coefficients are inserted to balance the equation.

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

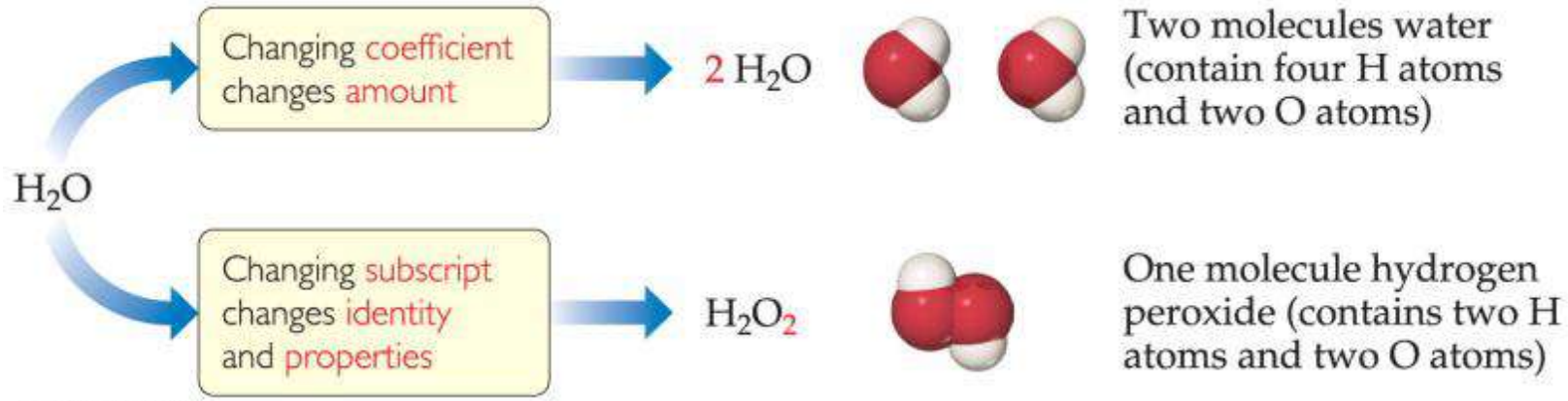
- Subscripts tell the number of atoms of each element in a molecule.

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
- Coefficients tell the number of molecules.

Why Do We Add Coefficients Instead of Changing Subscripts to Balance?



- Hydrogen and oxygen can make water OR hydrogen peroxide:
  - $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$
  - $H_2(g) + O_2(g) \rightarrow H_2O_2(l)$

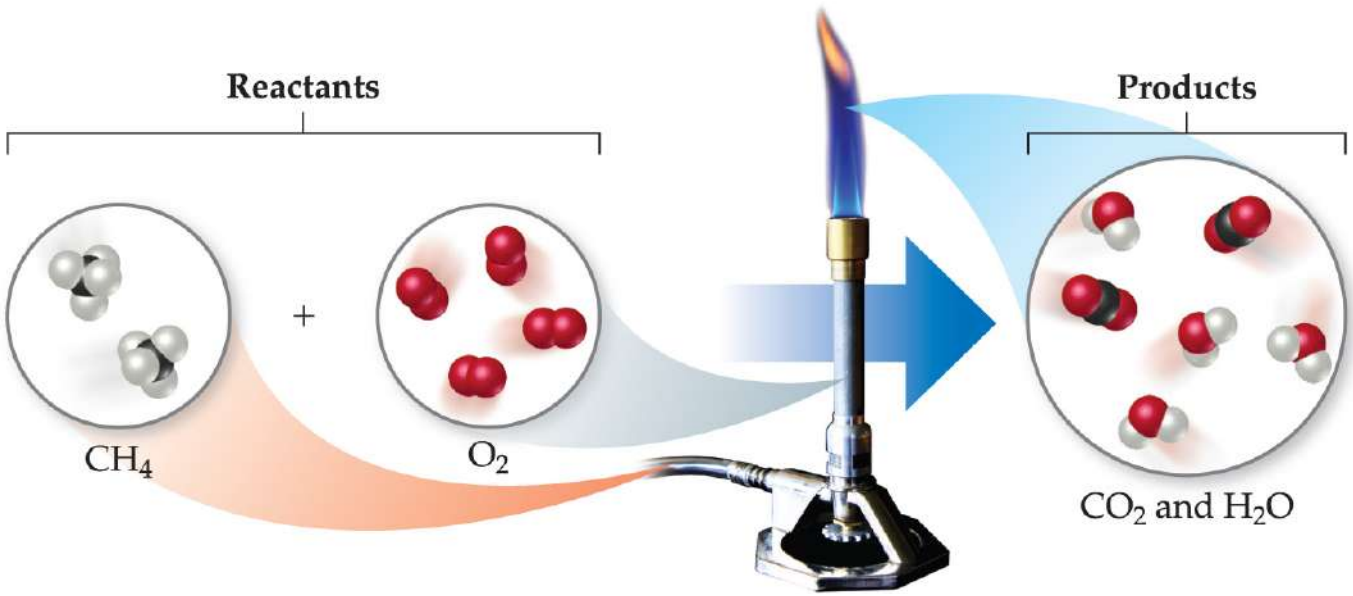
In the molecular level views shown in the figure how many C, H, and O atoms are present on the reactant side? Are the same number of each type of atom present on the product side?

a.18

b.14

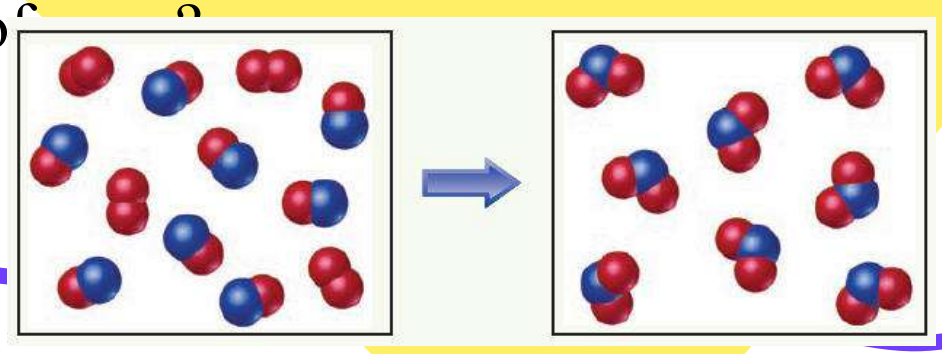
c.6

d.4



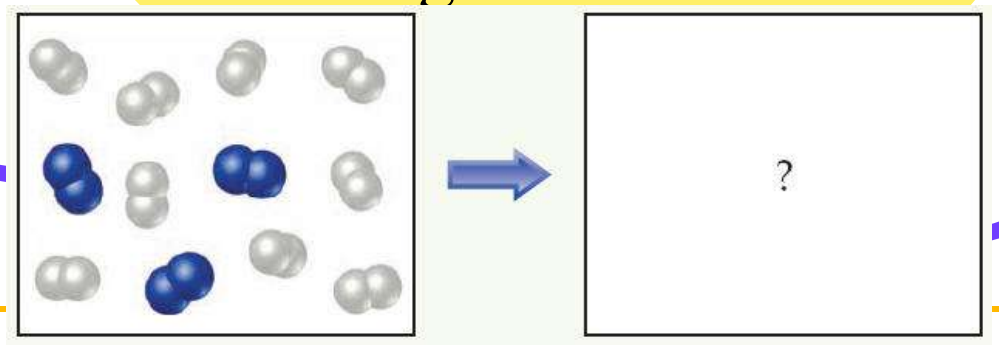
Sample Exercise 3.1 Interpreting and Balancing Chemical Equations

The following diagram represents a chemical reaction in which the red spheres are oxygen atoms and the blue spheres are nitrogen atoms. (a) Write the chemical formulas for the reactants and products. (b) Write a balanced equation for the reaction. (c) Is the diagram consistent with the law of conservation o

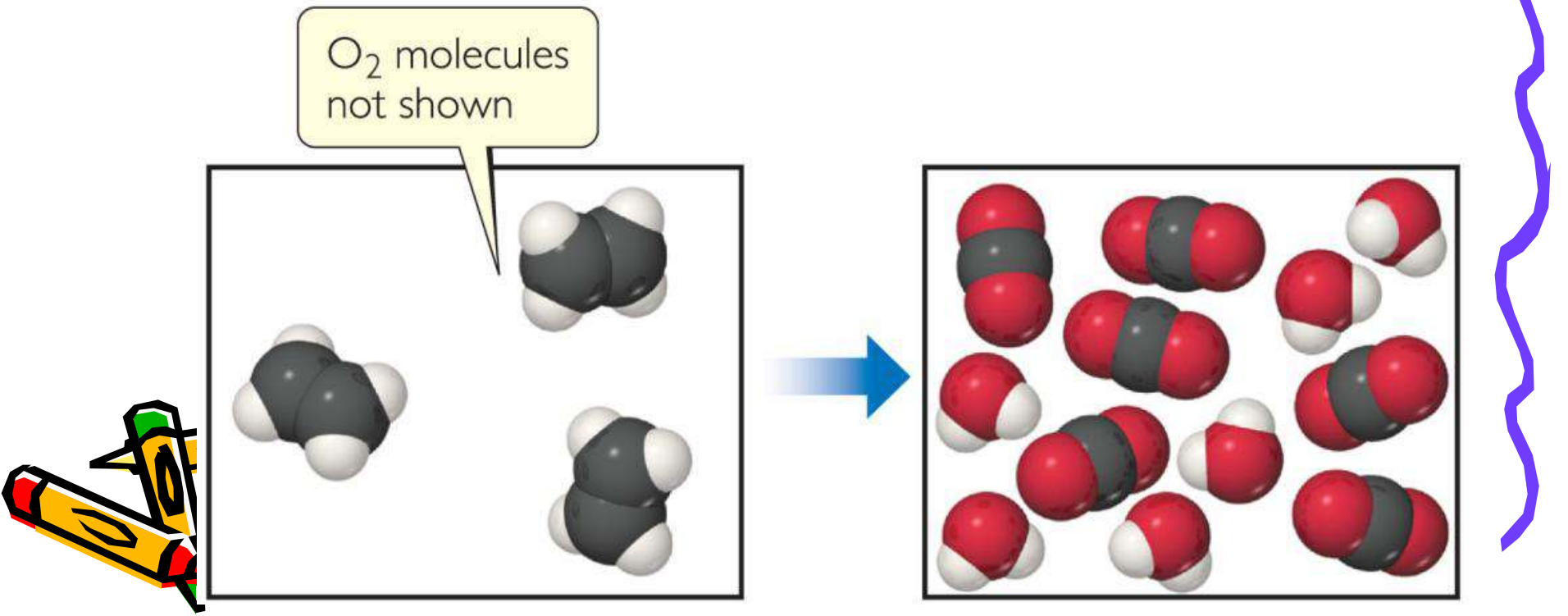


Practice Exercise 1

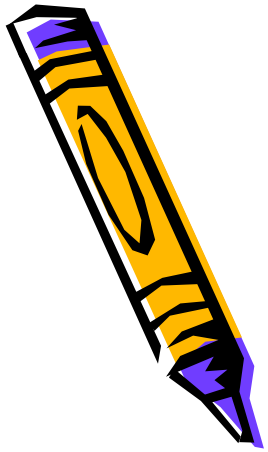
In the following diagram, the white spheres represent hydrogen atoms, and the blue spheres represent nitrogen atoms. To be consistent with the law of conservation of mass, how many NH3 molecules should be shown in the right box?



Practice Exercise 2: White spheres are H, black are C and red are O. Ethylene, C_2H_4 and oxygen, O_2 (not shown) are reactants and CO_2 and H_2O are products. A) Write a balanced chemical equation for the reaction.
B) Determine the number of O_2 molecules that should be shown.



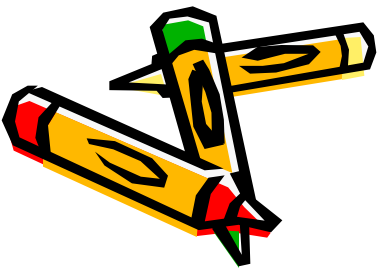
Sample Exercise 3.2 & Practice Exercises

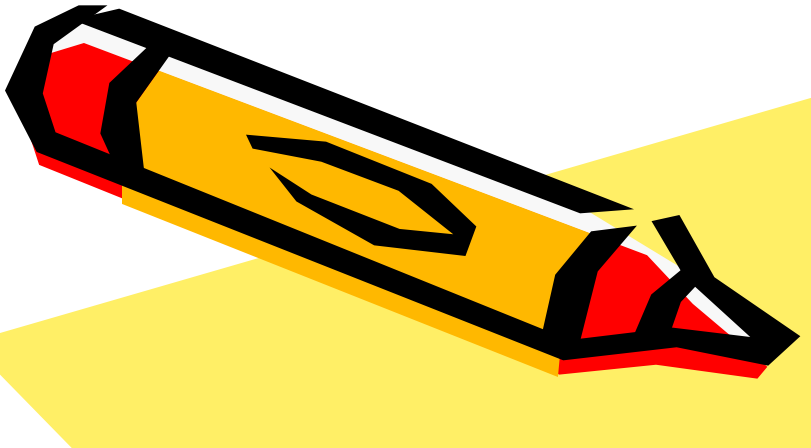


- $Na + H_2O \rightarrow NaOH + H_2$
- $CH_4 + Br_2 \rightarrow CBr_4 + HBr$
- $Fe + O_2 \rightarrow Fe_2O_3$
- $Al + HCl \rightarrow AlCl_3 + H_2$
- $CaCO_3 + HCl \rightarrow CaCl_2 + CO_2 + H_2O$

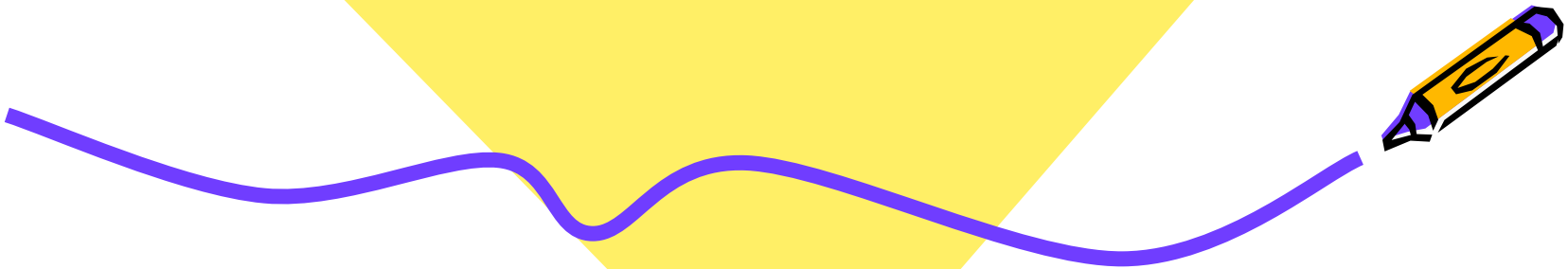
How many atoms of Mg, O, and H are represented by the notation 3 $Mg(OH)_2$?

a. 1 Mg, 2 O, and 2 H
b. 2 Mg, 2 O, and 2 H
c. 6 Mg, 6 O, and 6 H
d. 3 Mg, 6 O, and 6 H

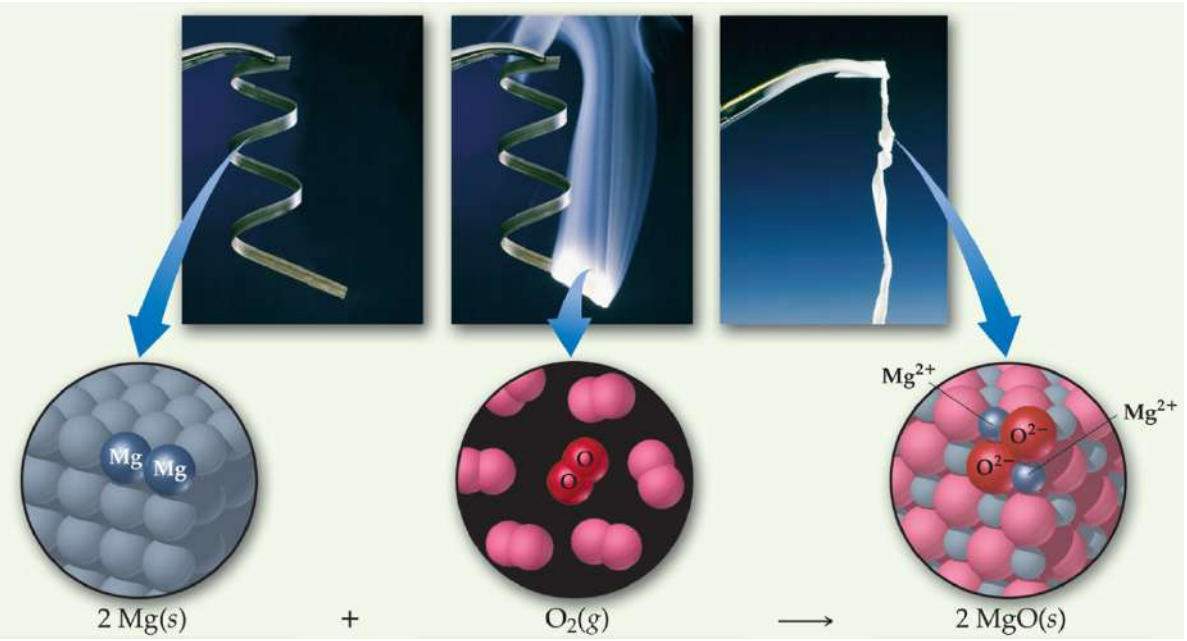




3.2 Simple Patterns of Chemical Reactivity



Combination Reactions



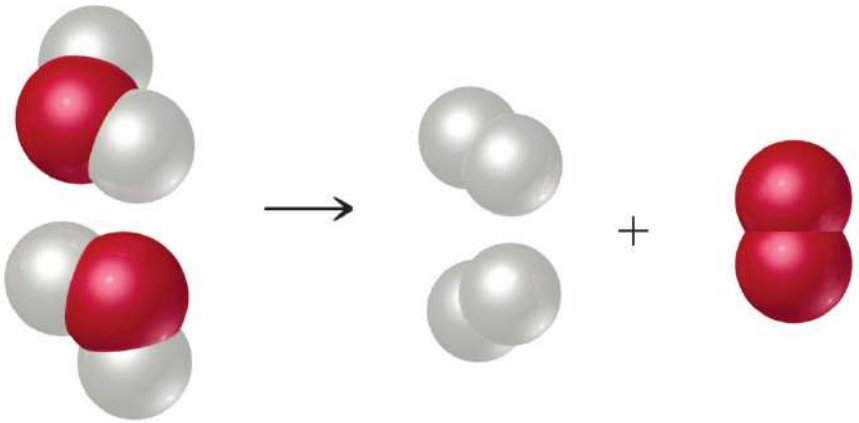
- In this type of reaction two or more substances react to form one product.

- Examples:
  - $2\ Mg_{(s)} + O_{2\ (g)} \longrightarrow 2\ MgO_{(s)}$
  - $N_{2\ (g)} + 3\ H_{2\ (g)} \longrightarrow 2\ NH_{3\ (g)}$
  - $C_3H_{6\ (g)} + Br_{2\ (l)} \longrightarrow C_3H_6Br_{2\ (l)}$

When Na and S undergo a combination reaction, what is the chemical formula of the product?

a. NaS
b. NaS_2
c. Na_2S
d. Na_2S_2

Decomposition Reactions



- In a decomposition one substance breaks down into two or more substances.

- Examples:
  - $CaCO_3\ (s) \longrightarrow CaO\ (s) + CO_2\ (g)$
  - $2\ KClO_3\ (s) \longrightarrow 2\ KCl\ (s) + O_2\ (g)$
  - $2\ NaN_3\ (s) \longrightarrow 2\ Na\ (s) + 3\ N_2\ (g)$

Sample Exercise 3.3

1) Write an equation for lithium metal reacting with fluorine gas.
2) Write an equation for the decomposition of solid barium carbonate (it produces a solid and a gas)
3) Write an equation for the decomposition of mercury (II) oxide.
4) Aluminum metal undergoes a combination reaction with oxygen in the air.

Practice Exercise 1

Which of the following reactions is the balanced equation that represents the decomposition reaction that occurs when silver (I) oxide is heated?

(a) $AgO(s) \rightarrow Ag(s) + O(g)$;

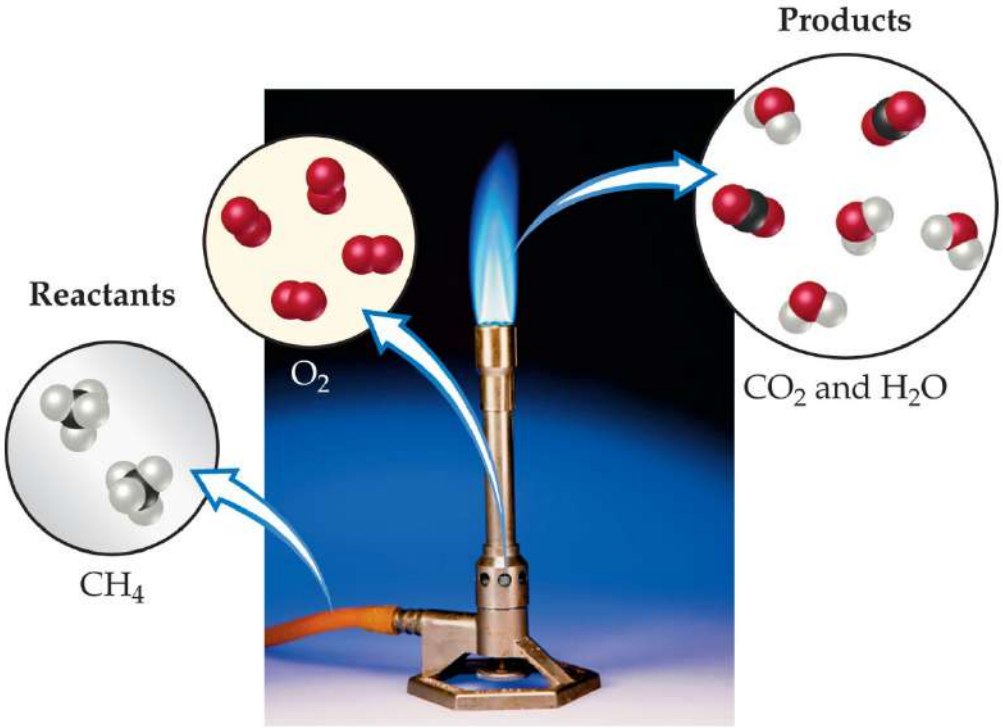
(b) $2\ AgO(s) \rightarrow 2\ Ag(s) + O_2(g)$;

(c) $Ag_2O(s) \rightarrow 2\ Ag(s) + O_2\ (g)$;

(d) $2\ Ag_2O(s) \rightarrow 4\ Ag(s) + O_2\ (g)$;

(e) $Ag_2O(s) \rightarrow 4\ Ag(s) + O_2(g)$;

Combustion Reactions



- These are generally rapid reactions that produce a flame.
- Most often involve hydrocarbons reacting with oxygen in the air.

- Examples:
  - $CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)$
  - $C_3H_8\,(g) + 5\,O_2\,(g) \longrightarrow 3\,CO_2\,(g) + 4\,H_2O\,(g)$

Does this reaction (Figure 3.8) produce or consume thermal energy (heat)?

a. Consumes heat
b. Produces heat

Sample Exercise 3.4

1) Write an equation for the reaction that occurs when methanol (CH_3OH) is burned in air.

2) Write an equation for the reaction that occurs when ethanol (C_2H_5OH) is burned in air.

Practice Exercise 1

Write the balanced equation for the reaction that occurs when ethylene glycol, $C_2H_4(OH)_2$, burns in air.

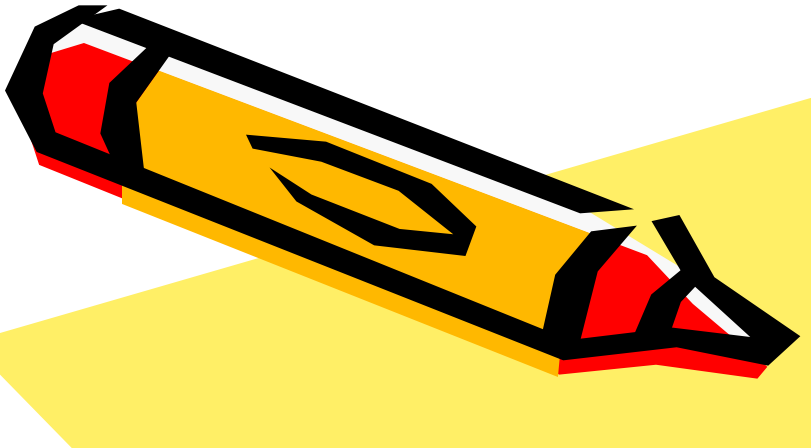
(**a**) $C_2H_4(OH)_2(l) + 5/2\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

(**b**) $2\ C_2H_4(OH)_2(l) + 5\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(g)$

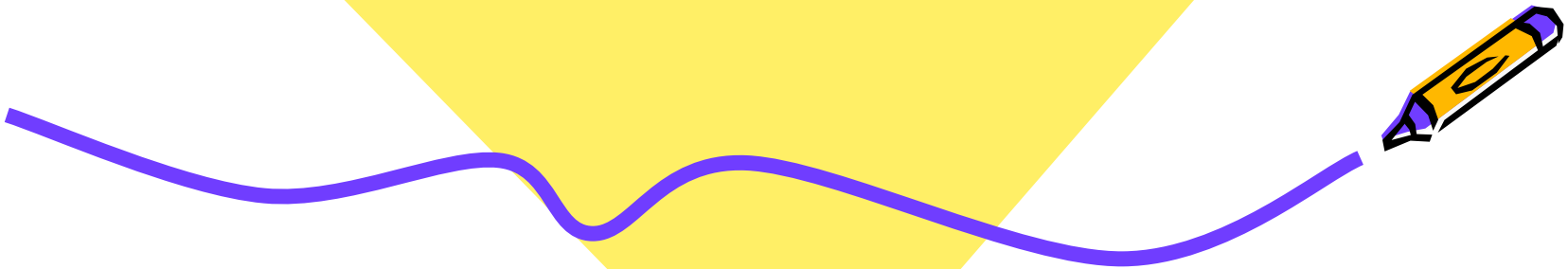
(**c**) $C_2H_4(OH)_2(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

(**d**) $C_2H_4(OH)_2(l) + 5\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$

(**e**) $4\ C_2H_4(OH)_2(l) + 10\ O_2(g) \rightarrow 8\ CO_2(g) + 12\ H_2O(g)$



3.3 Formula Weights



Formula Weight (FW)

- A formula weight is the sum of the atomic weights for the atoms in a chemical formula.
- So, the formula weight of calcium chloride, $CaCl_2$, would be

Ca: 1(40.1 amu)
+ Cl: 2(35.5 amu)
111.1 amu

- Formula weights are generally reported for ionic compounds.

Molecular Weight (MW)

- A molecular weight is the sum of the atomic weights of the atoms in a molecule.
- For the molecule ethane, C_2H_6, the molecular weight would be

C: 2(12.0 amu)
+ H: 6(1.0 amu)
30.0 amu

Sample Exercise 3.5

- Calculate the formula weight of:
  - A) sucrose ($C_{12}H_{22}O_{11}$)
  - B) calcium nitrate, $Ca(NO_3)_2$
  - C) $Al(OH)_3$
  - D) CH_3OH
  - E) TaON

Practice Exercise 1

Which of the following is the correct formula weight for calcium phosphate?

(a) 310.2 amu
(b) 135.1 amu
(c) 182.2 amu
(d) 278.2 amu
(e) 175.1 amu

Percent Composition

One can find the percentage of the mass of a compound that comes from each of the elements in the compound by using this equation:

$$\% \text{ element} = \frac{(\text{number of atoms})(\text{atomic weight})}{(\text{FW of the compound})} \times 100$$

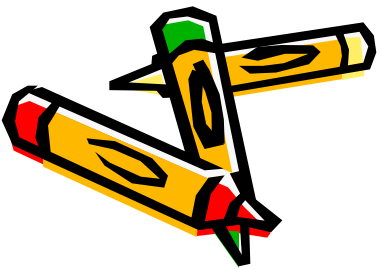
Percent Composition

So the percentage of carbon in ethane (C_2H_6) is…

$$\%C = \frac{(2)(12.0\ \text{amu})}{(30.0\ \text{amu})}$$

$$= \frac{24.0\ \text{amu}}{30.0\ \text{amu}} \times 100$$

$$= 80.0\%$$



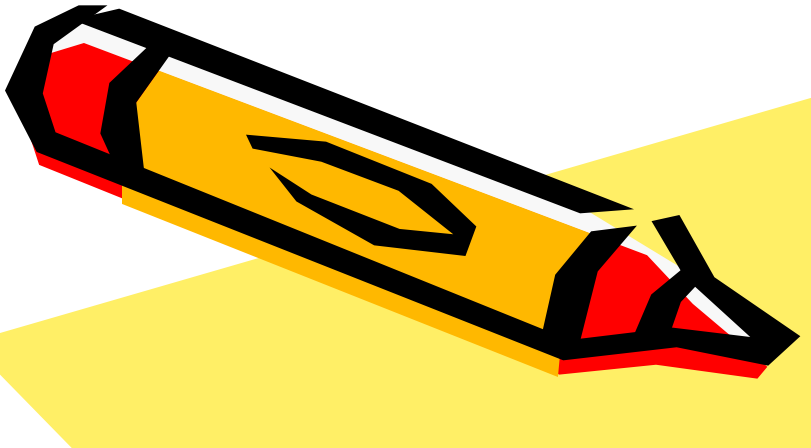
Sample Exercise 3.6

- 1) Calculate the percent by mass of all elements in sucrose ($C_{12}H_{22}O_{11}$).
- 2) Calculate the percent of potassium, by mass, in K_2PtCl_6.

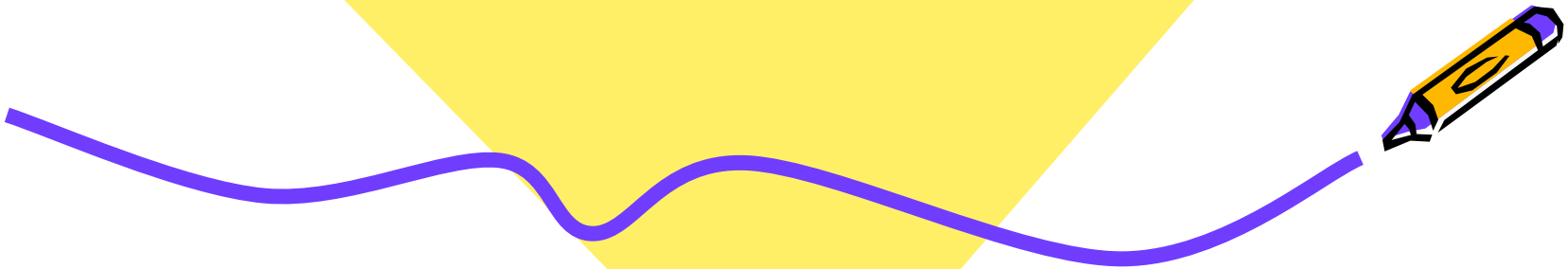
Practice Exercise 1

What is the percentage of nitrogen, by mass, in calcium nitrate?

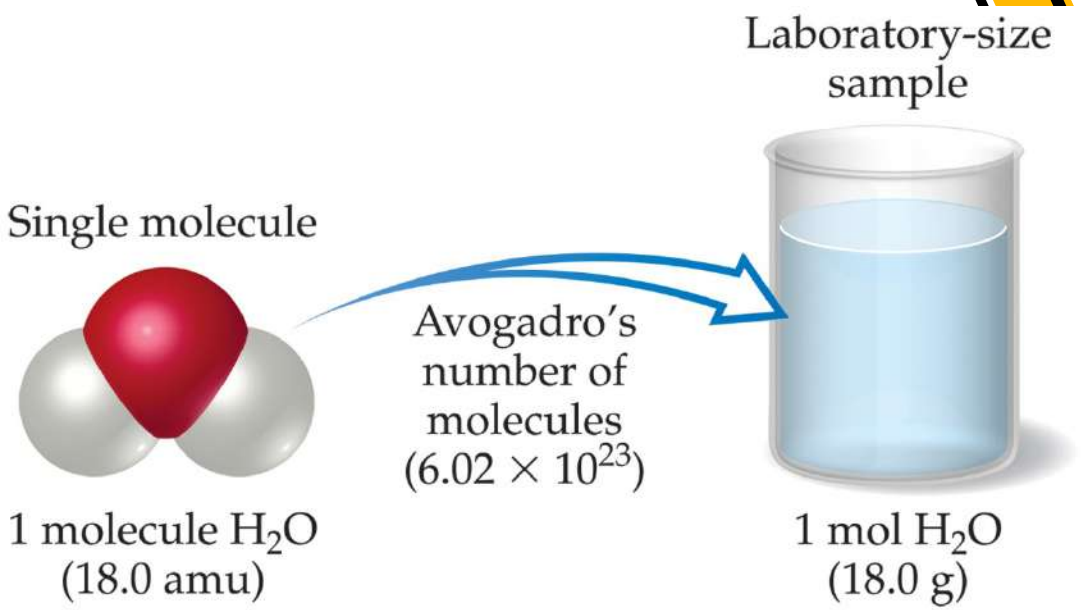
(a) 8.54%
(b) 17.1%
(c) 13.7%
(d) 24.4%
(e) 82.9%.



3.4 Avogadro's Number and the Mole



Avogadro's Number



- 6.02×10^{23}
- 1 mole of ^{12}C has a mass of 12 g.

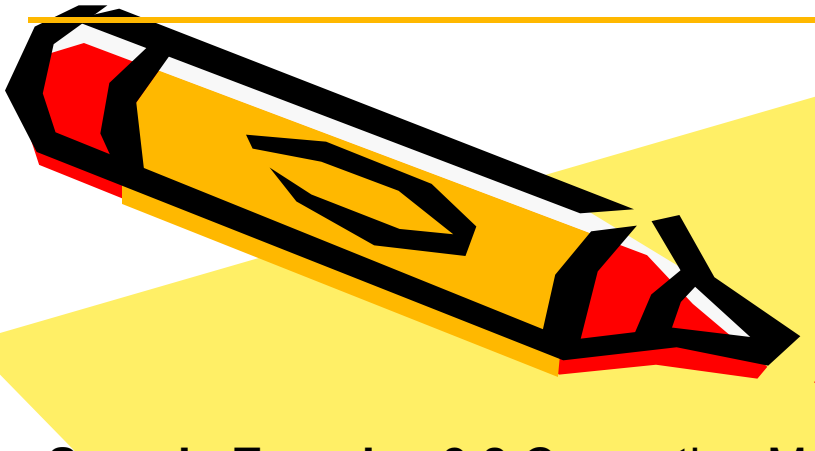
Sample Exercise 3.7

- Without a calculator, arrange the following in order of increasing numbers of carbon atoms: 12 g ^{12}C, 1 mol C_2H_2, 9×10^{23} molecules CO_2
- Arrange in order of increasing number of oxygen atoms: 1 mol H_2O, 1 mol CO_2, 3×10^{23} molecules O_3

Practice Exercise 1

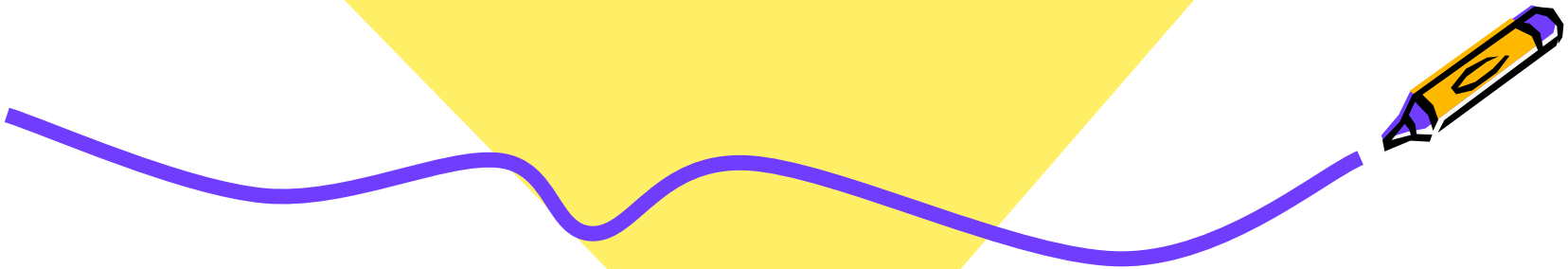
Determine which of the following samples contains the fewest sodium atoms?

(a) 1 mol sodium oxide
(b) 45 g sodium fluoride
(c) 50 g sodium chloride
(d) 1 mol sodium nitrate?



Sample Exercise 3.8 Converting Moles to Atoms

Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$.

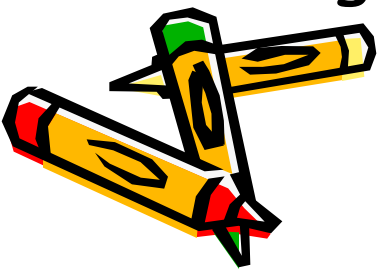


Practice Exercises 1 & 2

- How many sulfur atoms are in:
  - A) 0.45 mol $BaSO_4$
  - B) 1.10 mol aluminum sulfide
- How many oxygen atoms are in:
  - A) 0.25 mol $Ca(NO_3)_2$
  - B) 1.50 mol sodium carbonate

Molar Mass

- By definition, a molar mass is the mass of 1 mol of a substance (i.e., g/mol).
  - The molar mass of an element is the mass number for the element that we find on the periodic table.
  - The formula weight (in amu's) will be the same number as the molar mass (in g/mol).



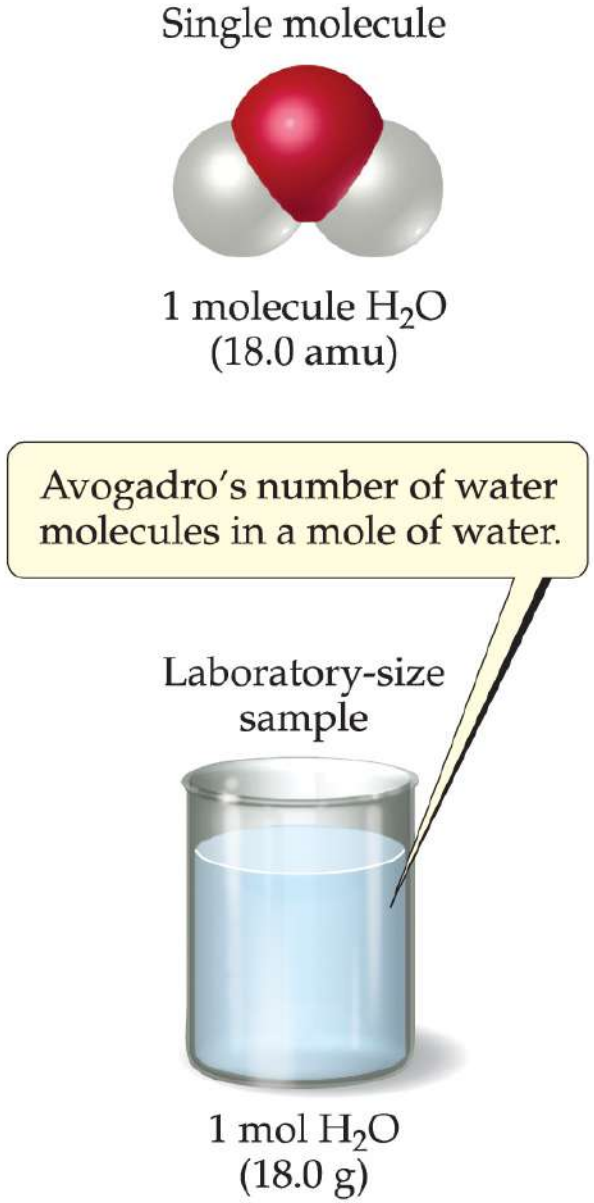
How many H_2O molecules are in a 9.00–g sample of water?

a. 0.500

b. 3.01×10^{23}

c. 2.71×10^{24}

d. 1.08×10^{23}



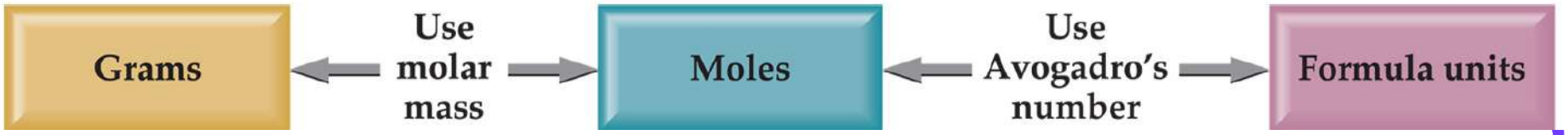
Which has more mass, a mole of water (H_2O) or a mole of glucose ($C_6H_{12}O_6$)?

a. Mole of glucose
b. Mole of water

Which contains more molecules, a mole of water or a mole of glucose?

a. Mole of water
b. Mole of glucose
c. Requires Avogadro's number to answer question
d. They both contain the same number of molecules

Using Moles



Moles provide a bridge from the molecular scale to the real-world scale.

Mole Relationships

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	6.02×10^{23} N atoms
Molecular nitrogen	N_2	28.0	28.0	6.02×10^{23} N_2 molecules $2(6.02 \times 10^{23})$ N atoms
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	6.02×10^{23} $BaCl_2$ units 6.02×10^{23} Ba^{2+} ions $2(6.02 \times 10^{23})$ Cl^- ions

[a] Recall that the electron has negligible mass; thus, ions and atoms have essentially the same mass.

- One mole of atoms, ions, or molecules contains Avogadro's number of those particles.
- One mole of molecules or formula units contains Avogadro's number times the number of atoms or ions of each element in the compound.

Sample Exercise 3.9

- What is the molar mass of glucose, $C_6H_{12}O_6$?
- Calculate the molar mass of $Ca(NO_3)_2$.

Practice Exercise 1

A sample of an ionic compound containing iron and chlorine is analyzed and found to have a molar mass of 126.8 g/mol. What is the charge of the iron in this compound?

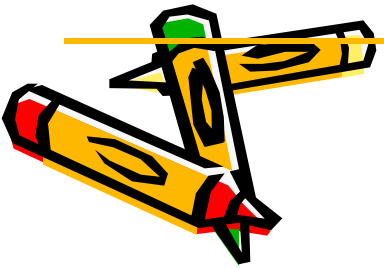
(a) 1+
(b) 2+
(c) 3+
(d) 4+

Sample Exercise 3.10 & Practice Exercises

Calculate the number of moles of glucose ($C_6H_{12}O_6$) in 5.380 g of $C_6H_{12}O_6$.

How many moles of sodium bicarbonate are in 508 g of $NaHCO_3$?

How many moles of water are in 1.00 L of water, whose density is 1.00 g/mL?



Sample Exercise 3.11 & Practice Exercises

Calculate the mass, in grams, of 0.433 mol of calcium nitrate.

What is the mass in grams of 6.33 mol $NaHCO_3$?

What is the mass in grams of 3.0×10^{-5} mol of sulfuric acid?

What is the mass in grams of 0.50 mol diamond (C)?

What is the mass in grams of 0.155 mol of ammonium chloride?

What number would you use to convert (a) moles of CH_4 to grams of CH_4 grams and (b) number of molecules of CH_4 to moles of CH_4?

a. Avogadro's number, 6.02×10^{23} particles/mol
b. Inverse of molar mass of CH_4, 1 mol CH_4/16.0 g CH_4
c. Molar mass of CH_4, 16.0 g CH_4/1 mol CH_4
d. Formula weight of CH_4, 16.0 amu

Sample Exercise 3.12 Calculating the Number of Molecules and Number of Atoms from Mass

How many glucose molecules are in 5.23 g of $C_6H_{12}O_6$? How many oxygen atoms are in this sample?

How many nitric acid molecules are in 4.20 g HNO_3? How many O atoms are in this sample?

Practice Exercise 1

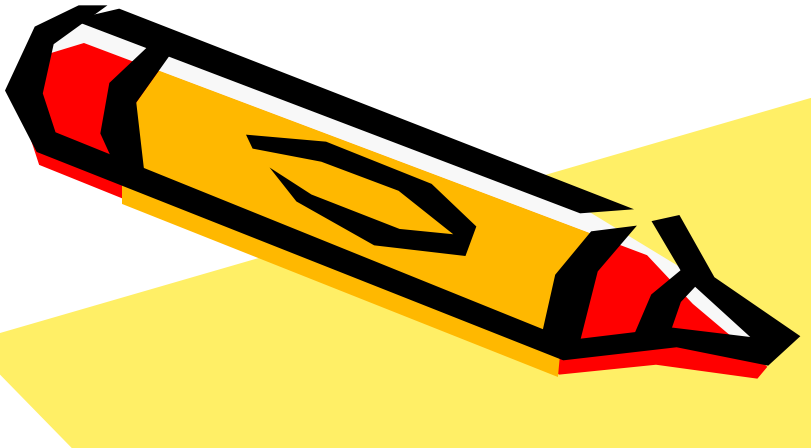
How many chlorine atoms are in 12.2 g of CCl_4?

(a) 4.77×10^{22}

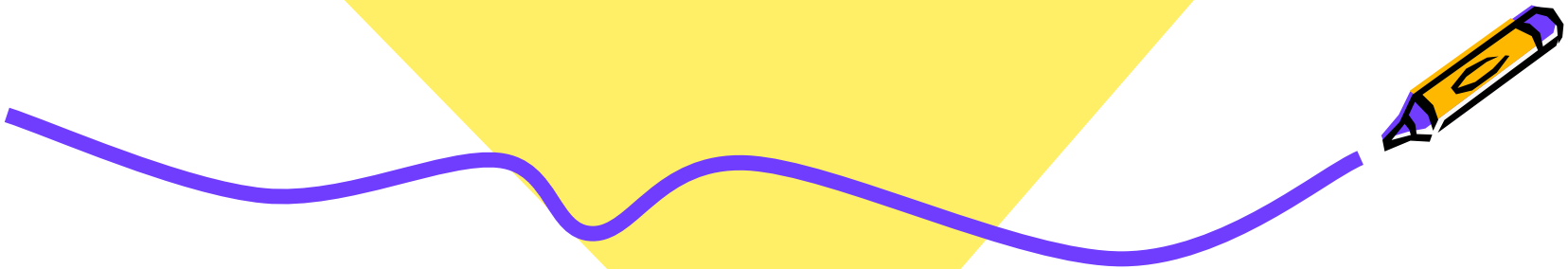
(b) 7.34×10^{24}

(c) 1.91×10^{23}

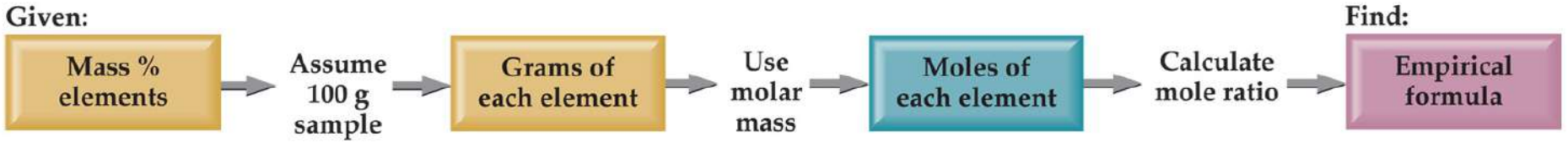
(d) 2.07×10^{23}



3.5 Empirical Formulas from Analyses



Calculating Empirical Formulas

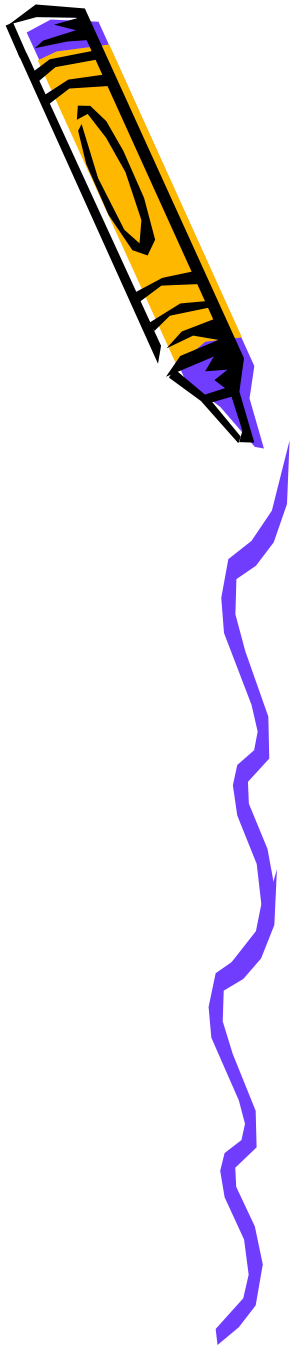


One can calculate the empirical formula from the percent composition.

Calculating Empirical Formulas

The compound *para*-aminobenzoic acid (you may have seen it listed as PABA on your bottle of sunscreen) is composed of carbon (61.31%), hydrogen (5.14%), nitrogen (10.21%), and oxygen (23.33%). Find the empirical formula of PABA.

Calculating Empirical Formulas



Assuming 100.00 g of *para*-aminobenzoic acid,

$$\text{C: } 61.31 \text{ g} \times \frac{1 \text{ mol}}{12.01 \text{ g}} = 5.105 \text{ mol C}$$

$$\text{H: } 5.14 \text{ g} \times \frac{1 \text{ mol}}{1.01 \text{ g}} = 5.09 \text{ mol H}$$

$$\text{N: } 10.21 \text{ g} \times \frac{1 \text{ mol}}{14.01 \text{ g}} = 0.7288 \text{ mol N}$$

$$\text{O: } 23.33 \text{ g} \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 1.456 \text{ mol O}$$

Calculating Empirical Formulas

Calculate the mole ratio by dividing by the smallest number of moles:

$$C := \frac{5.105\ \cancel{mol}}{0.7288\ \cancel{mol}} \approx 7$$

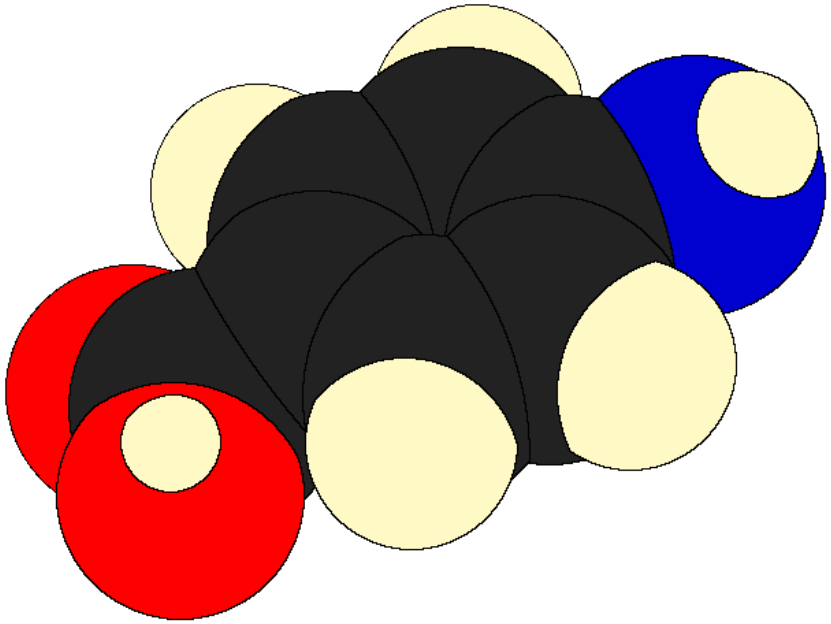
$$H := \frac{5.09\ \cancel{mol}}{0.7288\ \cancel{mol}} \approx 7$$

$$N := \frac{0.7288\ \cancel{mol}}{0.7288\ \cancel{mol}} = 1$$

$$O := \frac{1.458\ \cancel{mol}}{0.7288\ \cancel{mol}} \approx 2$$

Calculating Empirical Formulas

These are the subscripts for the empirical formula: $C_7H_7NO_2$



Could the empirical formula determined from chemical analysis be used to tell the difference between acetylene, C_2H_2, and benzene, C_6H_6?

a. Yes, because the empirical formula is specific to the molecule.

b. Yes, because the empirical formula tells the number of atoms in the molecule.

c. No, because the empirical formula cannot be determined for a liquid.

d. No, because the empirical formulas for C_2H_2 and C_6H_6 will be the same.

Sample Exercise 3.13

- 1) Ascorbic acid (Vitamin C) contains 40.92% C, 4.58% H, and 54.50% O by mass. What is the empirical formula of ascorbic acid?
- 2) A 5.325-g sample of methyl benzoate, a compound used in manufacturing perfumes, is found to contain 3.758 g of carbon, 0.316 g of hydrogen, and 1.251 g of oxygen. What is the empirical formula of this substance?

Practice Exercise 1

A 2.144-g sample of phosgene, a compound used as a chemical warfare agent during World War I, contains 0.260 g of carbon, 0.347 g of oxygen, and 1.537 g of chlorine. What is the empirical formula of this substance?

(a) CO_2Cl_6

(b) $COCl_2$

(c) $C_{0.022}O_{0.022}Cl_{0.044}$

(d) C_2OCl_2

Molecular Formula

- Is a multiple of the empirical formula
- The multiple can be found by comparing the empirical formula weight with the molecular weight.

Sample Exercise 3.14

- 1) Mesitylene, a hydrocarbon in crude oil, has an empirical formula of C_3H_4. The experimentally determined molecular weight is 121 amu. What is the molecular formula?
- 2) Ethylene glycol, used in antifreeze, is 38.7% C, 9.7% H, and 51.6% O by mass. Its molar mass is 62.1 g/mol. What are the empirical and molecular formulas?

Practice Exercise 1

Cyclohexane, a commonly used organic solvent, is 85.6% C and 14.4% H by mass with a molar mass of 84.2 g/mol. What is its molecular formula?

(a) C_6H

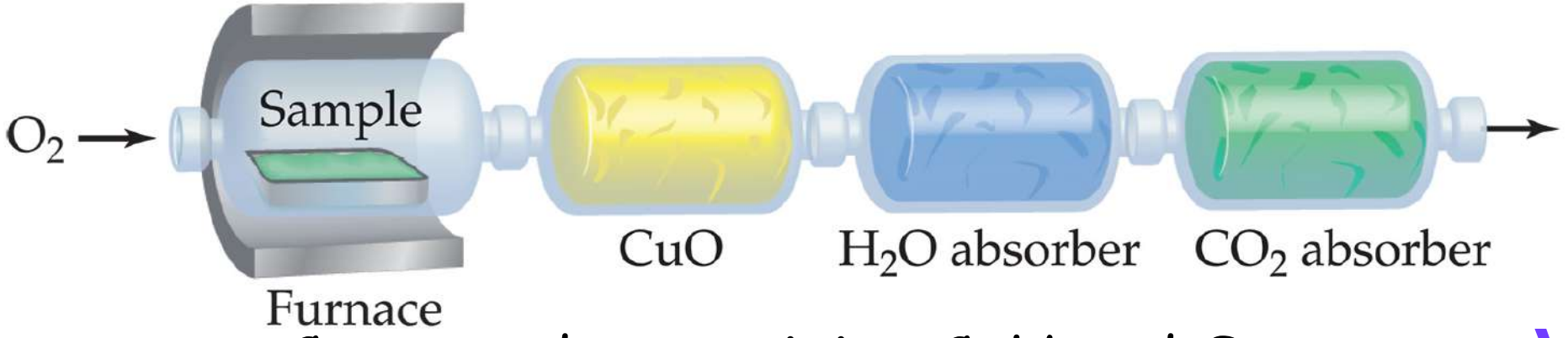
(b) CH_2

(c) C_5H_{24}

(d) C_6H_{12}

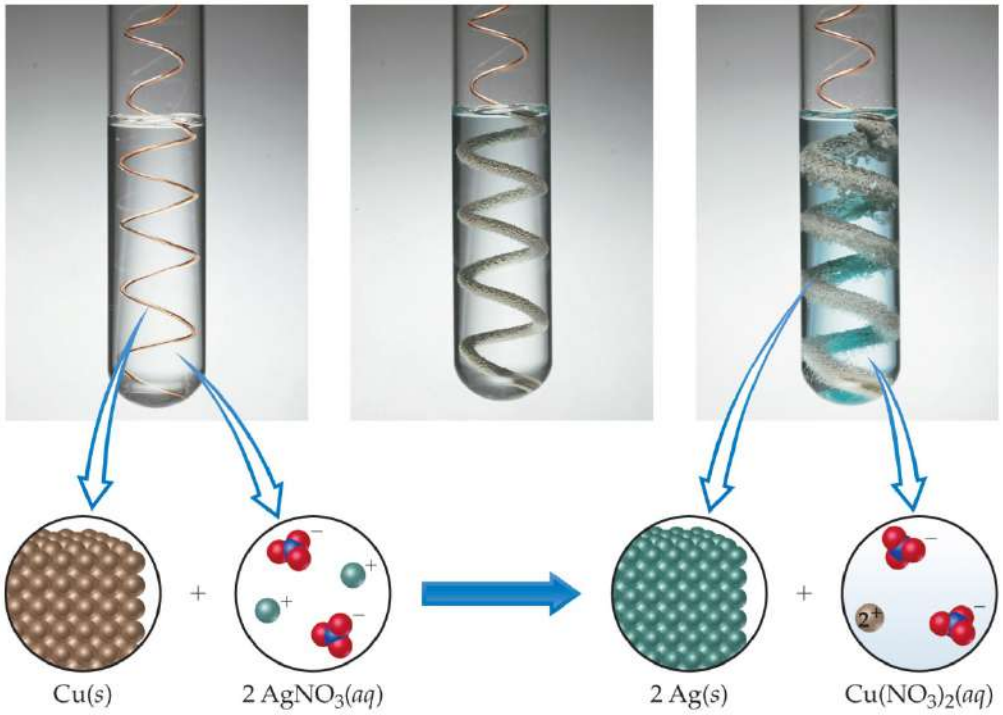
(e) C_4H_8

Combustion Analysis



- Compounds containing *C*, *H* and *O* are routinely analyzed through combustion in a chamber like this.
  - *C* is determined from the mass of CO_2 produced.
  - *H* is determined from the mass of H_2O produced.
  - *O* is determined by difference after the *C* and *H* have been determined.

Elemental Analyses



Compounds containing other elements are analyzed using methods analogous to those used for C, H and O.

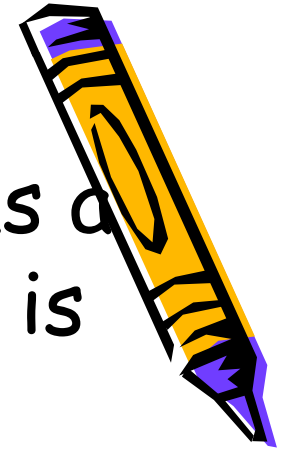
Sample Exercise 3.15 Determining Empirical Formula by Combustion Analysis

Isopropyl alcohol, a substance sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.

In Sample Exercise 3.15, how do you explain that the values in our calculated C:H:O ratio are 3.0:7.9:1.0 rather than exact integers 3:8:1?

a. An incorrect molar mass for carbon is used in the problem.
b. Approximations are used in the problem.
c. An incorrect number of significant figures is used in the problem.
d. Experimental uncertainties in the experimental measurements.

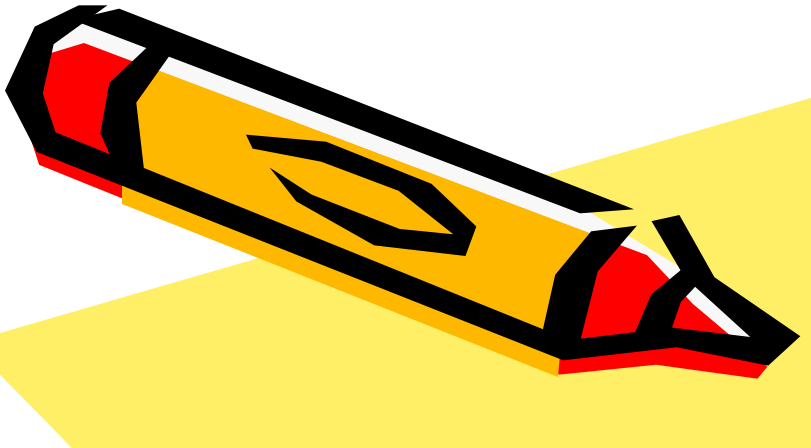
Practice Exercise 1



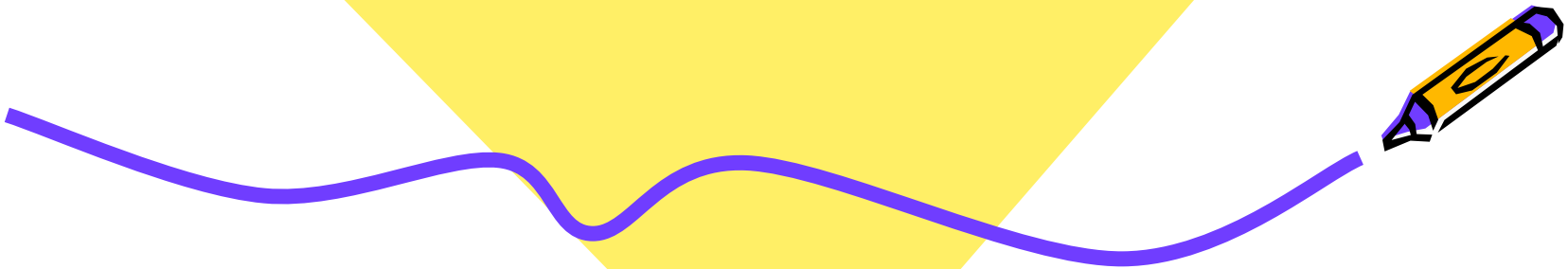
The compound dioxane, which is used as a solvent in various industrial processes, is composed of C, H, and O atoms. Combustion of a 2.203-g sample of this compound produces 4.401 g CO_2 and 1.802 g H_2O. A separate experiment shows that it has a molar mass of 88.1 g/mol. Which of the following is the correct molecular formula for dioxane? **(a)** C_2H_4O, **(b)** $C_4H_4O_2$, **(c)** CH_2, **(d)** $C_4H_8O_2$

Practice Exercise 2

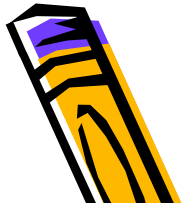
- 1) Caproic acid, which is responsible for the foul odor of dirty socks, is composed of C, H, and O. Combustion of a 0.225 g sample produces 0.512 g CO_2 and 0.209 g H_2O. What is the empirical formula of caproic acid?
- 2) If caproic acid has a molar mass of 116 g/mol, what is its molecular formula?

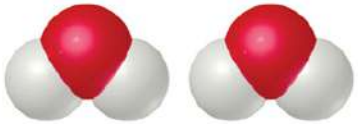


3.6 Quantitative Information from Balanced Equations



Stoichiometric Calculations

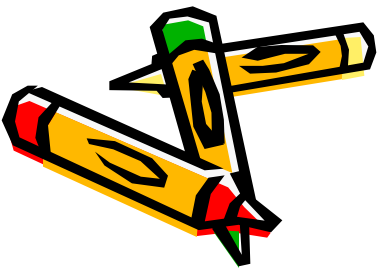


Equation:	$2\ H_2(g)$	+	$O_2(g)$	⟶	$2\ H_2O(l)$
Molecules:	2 molecules H_2	+	1 molecule O_2	⟶	2 molecules H_2O
					
Mass (amu):	4.0 amu H_2	+	32.0 amu O_2	⟶	36.0 amu H_2O
Amount (mol):	2 mol H_2	+	1 mol O_2	⟶	2 mol H_2O
Mass (g):	4.0 g H_2	+	32.0 g O_2	⟶	36.0 g H_2O

The coefficients in the balanced equation give the ratio of *moles* of reactants and products.

When 1.57 mol O_2 reacts with H_2 to form H_2O, how many moles of H_2 are consumed in the process?

a. 1.57 mol
b. 3.14 mol
c. 6.28 mol
d. 9.42 mol



1.00 g of C_4H_{10} reacts with 3.59 g of O_2 to form 3.03 g of CO_2. Using only addition and subtraction, calculate the amount of H_2O produced.

$$2\,C_4H_{10}(l) + 13\,O_2(g) \longrightarrow 8\,CO_2(g) + 10\,H_2O(g)$$

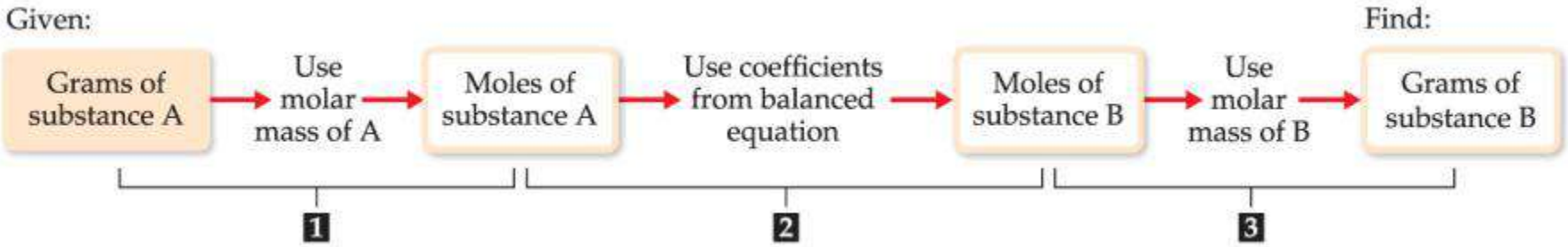
a. 1.56 g H_2O

b. 3.12 g H_2O

c.5.00 g H_2O

d.7.62 g H_2O

Stoichiometric Calculations



We have already seen in this chapter how to convert from grams to moles or moles to grams. The NEW calculation is how to compare two DIFFERENT materials, using the MOLE RATIO from the balanced equation!

Sample Exercise 3.16

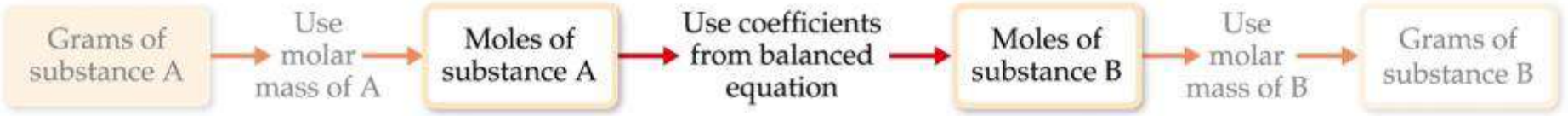
$$\text{Moles } C_6H_{12}O_6 = (1.00 \text{ g } C_6H_{12}O_6)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0 \text{ g } C_6H_{12}O_6}\right)$$

- How many grams of water can be produced from 1.00 g of glucose?

$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$

- There is 1.00 g of glucose to start.
- The first step is to convert it to moles.

Sample Exercise 3.16



$$\text{Moles } H_2O = (1.00\ \text{g } C_6H_{12}O_6)\left(\frac{1\ \text{mol } C_6H_{12}O_6}{180.0\ \text{g } C_6H_{12}O_6}\right)\left(\frac{6\ \text{mol } H_2O}{1\ \text{mol } C_6H_{12}O_6}\right)$$

- The NEW calculation is to convert moles of one substance in the equation to moles of another substance.
- The MOLE RATIO comes from the balanced equation.

Sample Exercise 3.16

$$\text{Grams } H_2O = (1.00 \text{ g } C_6H_{12}O_6)\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0 \text{ g } C_6H_{12}O_6}\right)\left(\frac{6 \text{ mol } H_2O}{1 \text{ mol } C_6H_{12}O_6}\right)\left(\frac{18.0 \text{ g } H_2O}{1 \text{ mol } H_2O}\right)$$

$$= 0.600 \text{ g } H_2O$$

- Then turn the moles of water to grams using molar mass

Practice Exercise 1

Sodium hydroxide reacts with carbon dioxide to form sodium carbonate and water:

$2\ NaOH(s) + CO_2(g) \rightarrow Na_2CO_3(s) + H_2O(l)$

How many grams of Na_2CO_3 can be prepared from 2.40 g of NaOH?

(a) 3.18 g

(b) 6.36 g

(c) 1.20 g

(d) 0.0300 g

Practice Exercise 2

- 1) $2KClO_3 \rightarrow 2KCl + O_2$
  - A) How many grams of O_2 can be prepared from 4.50 g of $KClO_3$?
  - B) How many moles of O_2 can be produced from 2.5 moles of $KClO_3$?
  - C) How many moles of KCl can be produced from 3.67 g of $KClO_3$?

Sample Exercise 3.17

Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide. Lithium hydroxide reacts with gaseous carbon dioxide to form solid lithium carbonate and liquid water. How many grams of carbon dioxide can be absorbed by 1.00 g of lithium hydroxide?

Practice Exercise 1

Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?

a)5.00 g

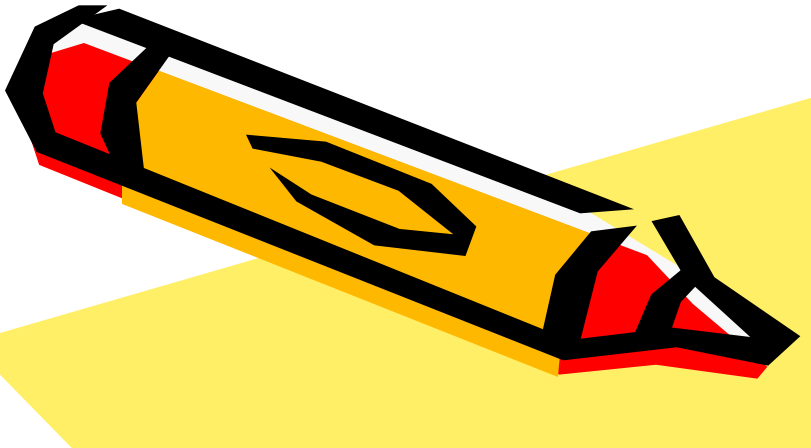
b)0.726 g

c)2.18 g

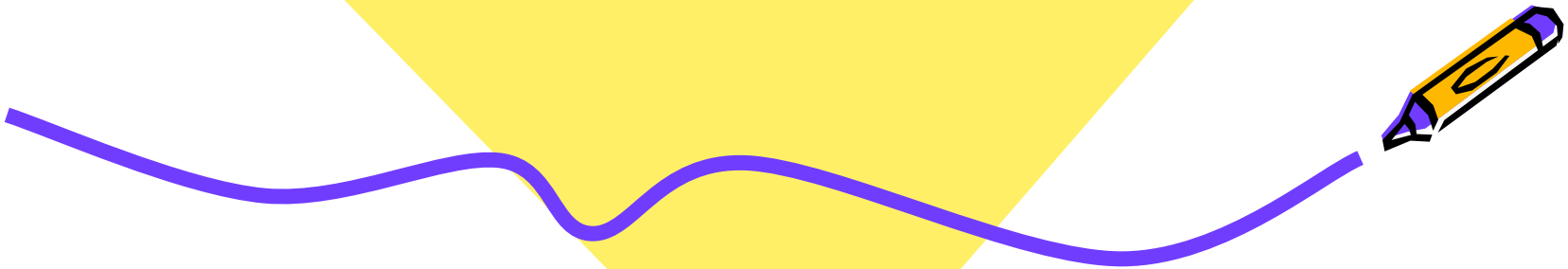
d)3.63 g

Practice Exercise 2

Methanol, CH_3OH, reacts with oxygen from air in a combustion reaction to form water and carbon dioxide. What mass of water is produced in the combustion of 23.6 g of methanol?



3.7 Limiting Reactants



How Many Cookies Can I Make?

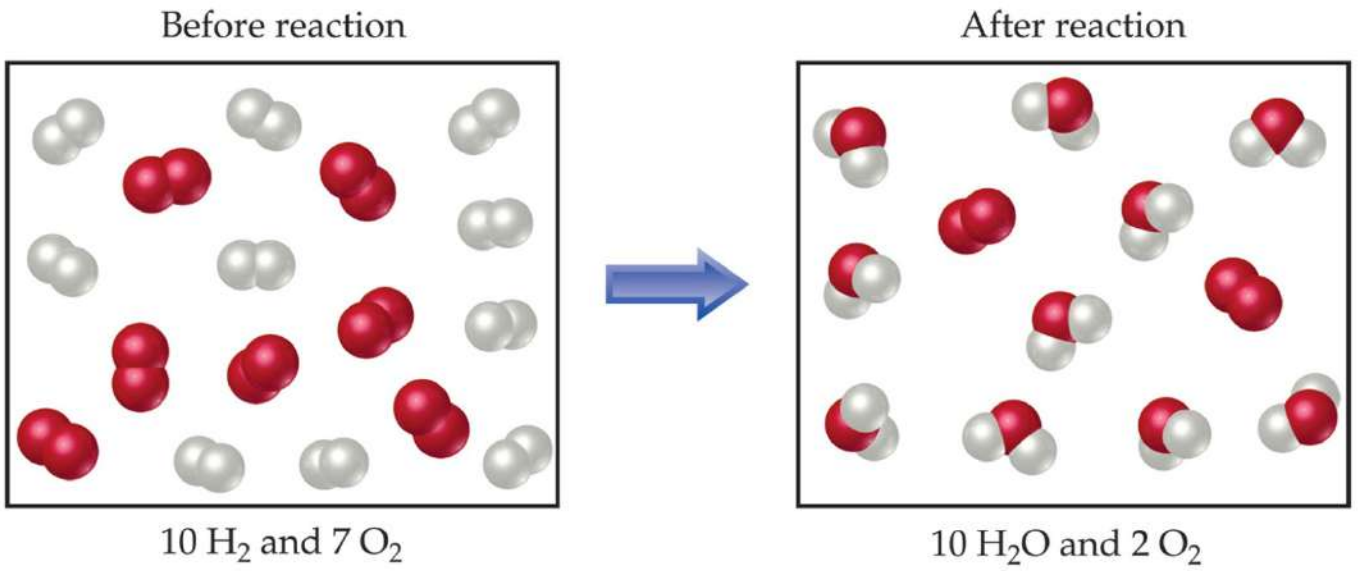
- You can make cookies until you run out of one of the ingredients.
- Once this family runs out of sugar, they will stop making cookies (at least any cookies you would want to eat).

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.

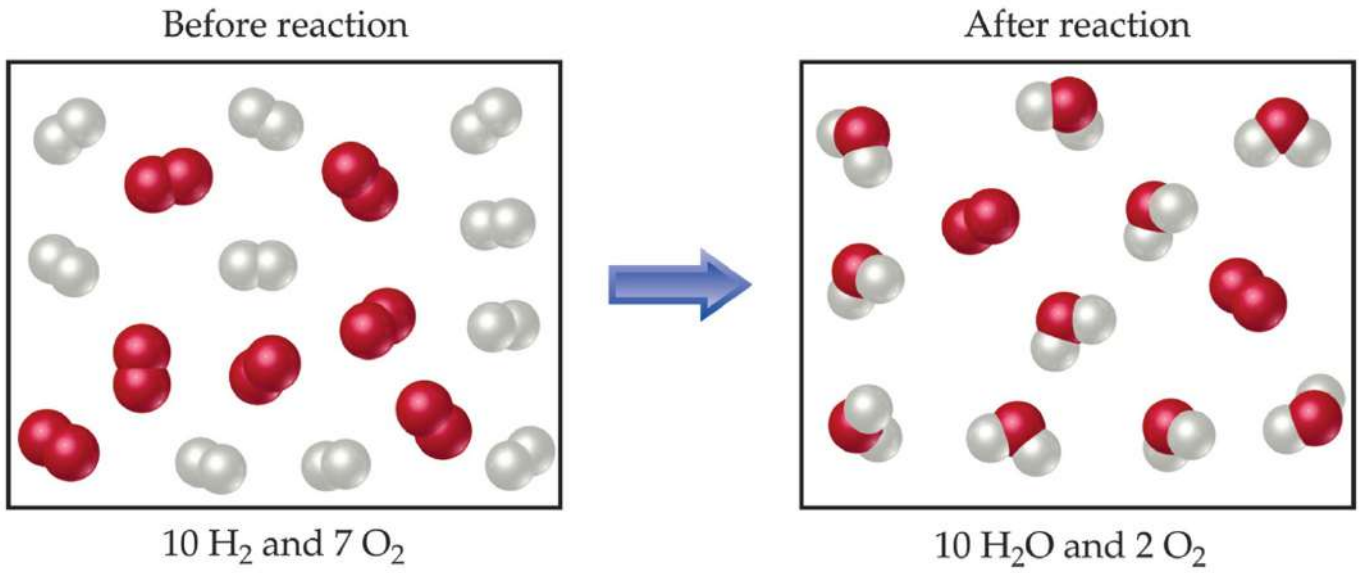
Limiting Reactants

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

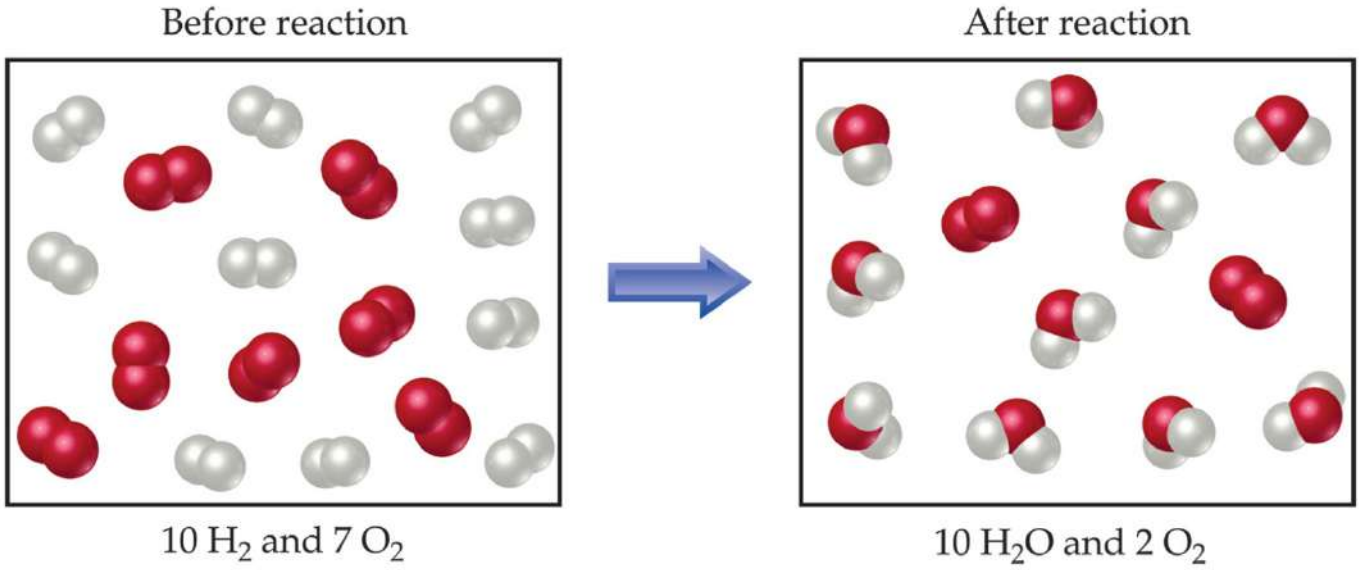


Limiting Reactants

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



If the amount of H_2 is doubled, how many moles of H_2O would have formed?



a. 10 mol H_2O

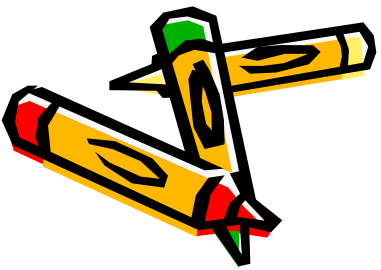
b. 12 mol H_2O

c. 14 mol H_2O

d. 20 mol H_2O

Sample Exercise 3.18

- 1) Given $N_2 + 3H_2 \rightarrow 2NH_3$, how many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?
- 2) In the equation $2Al + 3Cl_2 \rightarrow 2AlCl_3$, 1.50 mol of Al reacts with 3.00 mol Cl_2.
  - A) Which reactant is the limiting reactant?
  - B) How many moles of $AlCl_3$ are formed?
  - C) How many moles of excess reactant are left over at the end of the reaction?



Practice Exercise 1

When 24 mol of methanol and 15 mol of oxygen combine in the combustion reaction $2\ CH_3OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(g)$, what is the excess reactant and how many moles of it remains at the end of the reaction?

(a) 9 mol $CH_3OH(l)$

(b) 10 mol $CO_2(g)$

(c) 10 mol $CH_3OH(l)$

(d) 11 mol $CH_3OH(l)$

(e) 11 mol $O_2(g)$

Sample Exercise 3.19 Calculating the Amount of Product Formed from a Limiting Reactant

Consider the following reaction that occurs in a fuel cell:

$2\ H_2(g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$

This reaction, properly done, produces energy in the form of electricity and water. Suppose a fuel cell is set up with 150 g of hydrogen gas and 1500 grams of oxygen gas (each measurement is given with two significant figures). How many grams of water can be formed?

Practice Exercise 1

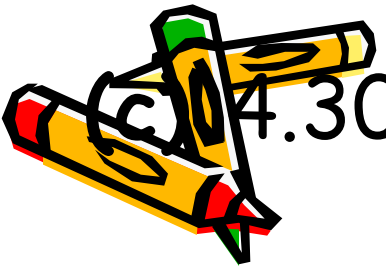
Molten gallium reacts with arsenic to form the semiconductor, gallium arsenide, GaAs, used in light-emitting diodes and solar cells:

$Ga(l) + As(s) \rightarrow GaAs(s)$

If 4.00 g of gallium is reacted with 5.50 g of arsenic, how many grams of the excess reactant are left at the end of the reaction?

(a) 1.20 g As, **(b)** 1.50 g As

(c) 4.30 g As, or **(d)** 8.30 g Ga



Practice Exercise 2

$Zn + 2AgNO_3 \rightarrow 2Ag + Zn(NO_3)_2$

- 2.00 g of zinc is placed in an aqueous solution containing 2.50 g of silver nitrate. Which reactant is limiting?
- How many grams of silver will form?
- How many grams of $Zn(NO_3)_2$ will form?
- How many grams of excess reactant will be left over?

Theoretical Yield

- The **theoretical yield** is the maximum amount of product that can be made.
  - In other words it's the amount of product possible as calculated through the stoichiometry problem.
- This is different from the **actual yield**, which is the amount one actually produces and measures.

Percent Yield

One finds the **percent yield** by comparing the amount actually obtained (actual yield) to the amount it was possible to make (theoretical yield).

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

Sample Exercise 3.20 Calculating Theoretical Yield and Percent Yield

Adipic acid, $H_2C_6H_8O_4$, used to produce nylon, is made commercially by a reaction between cyclohexane (C_6H_{12}) and O_2:

$$2\ C_6H_{12}(l) + 5\ O_2(g) \rightarrow 2\ H_2C_6H_3O_4(l) + 2\ H_2O(g)$$

(a) Assume that you carry out this reaction with 25.0 g of cyclohexane and that it is the limiting reactant. What is the theoretical yield of adipic acid?

(b) If you obtain 33.5 g of adipic acid, what is the percent yield?

Practice Exercise 1

If 3.00 g of titanium metal is reacted with 6.00 g of chlorine gas, Cl_2, to form 7.7 g of titanium (IV) chloride in a combination reaction, what is the percent yield of the product?

(a) 65%
(b) 96%
(c) 48%
(d) 86%

Practice Exercise 2

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

a) If you start with 150 g of Fe_2O_3 as the limiting reagent, what is the theoretical yield of Fe?

b) If the actual yield of Fe in your test was 87.9 g, what is the percent yield?