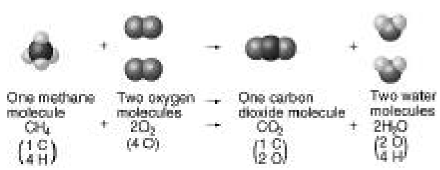
Antoine Lavoisier observed that the total mass of all substances present after a chemical reaction is the same as the total mass before the reaction.

The law of Conservation of Mass

Atoms are neither created nor destroyed.

Chemical reactions are represented in a concise way by chemical equations.

$$CH_4 + 2O_2 \rightarrow CO_2\ 2H_2O$$



Writing Chemical Equations

reactants ⟶ products

A + B ⟶ C + D

$$HC_2H_3O_{2(aq)} + NaHCO_{3(s)} \longrightarrow NaC_2H_3O_{2\,(aq)}\ H_2O_{(l)} + CO_{2(g)}$$

Balancing Chemical Equations

An equation is balanced when the number of atoms of each element is the same on each side of the arrow. The amounts of reactants and products are equal.

$$H_2 + Cl_2 \longrightarrow 2\ HCl$$

$$4\ Al + 3\ O_2 \longrightarrow 2\ Al_2O_3$$

Guidelines for Balancing A Chemical Equation

1. Use Correct formulas
2. Balance each element in the equation by placing a coefficient in front of each formula.
   a) Begin balancing those elements that occur in the fewest chemical formulas on each side of the equation.
   b) Balance polyatomic ions as a single unit. (sometimes they decompose)
   c) The coefficients in an equation must be whole numbers. (hint use ½ for O_2)
3. Check the # of each element on both sides of the equation.

USING THE PERIODIC TABLE TO MAKE CHEMICAL PREDICTIONS (PATTERNS OF CHEMICAL REACTIVITY).

$2K + 2H_2O \rightarrow 2KOH + H_2$

Alkali metal + water → Metal hydroxide + hydrogen

$2M + 2H_2O \rightarrow 2MOH + H_2$

$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$

Alkaline earth metal + water → Metal hydroxide + hydrogen

$M + 2H_2O \rightarrow M(OH)_2 + H_2$

Combustion Reactions

When hydrocarbons are combusted they react with O_2 to form CO_2 and H_2O

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + H_2O$

Combination reaction

Two or more substance combine into a single compound (*synthesis reaction*).

$A + Z \longrightarrow AZ$

Combination reaction

$A + Z \longrightarrow AZ$

metal + oxygen gas ⟶ metal oxide

$4Fe + 3O_2 \longrightarrow 2\,Fe_2O_3$

nonmetal + oxygen gas ⟶ nonmetal oxide

$S + O_2 \longrightarrow SO_2$

metal + nonmetal ⟶ ionic compound

$2Na + Cl_2 \longrightarrow 2\,NaCl$

Decomposition reaction

A single compound is broken down into two or more substances.

$AZ \longrightarrow A + Z$

Decomposition reaction

$AZ \longrightarrow A + Z$

metal hydrogen carbonate ⟶ metal carbonate + water + carbon dioxide

$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2$

Metal carbonate ⟶ metal oxide + carbon dioxide

$CaCO_3 \longrightarrow CaO + CO_2$

oxygen compound ⟶ substance + oxygen gas

$2HgO \longrightarrow 2\,Hg + O_2$

ATOMIC MASS SCALE

A hydrogen(^{1}H) atom has a mass of $1.6735x10^{-24}$g

It is convenient to use a unit called the atomic mass unit (amu).

1 amu = $1.66054x10^{-24}$g and 1g = $6.02214x10^{23}$ amu

The amu is defined by assigning a mass of exactly 12 amu to the ^{12}C isotope of carbon. The relative mass of an atom is compared to a reference atom using an instrument called a mass spectrometer. Carbon-12 has been chosen as a reference. The masses of other atoms are compared relative to carbon-12.

Average Atomic Masses

Most elements occur in nature as mixtures of isotopes. We can determine the average atomic mass of an element by using the masses of its various isotopes and their relative abundance.

Naturally occurring carbon is composed of 98.892% ^{12}C and 1.108% ^{13}C.

(0.98892)(12 amu) + (0.01108)(13.00335 amu)=12.011 amu

The average atomic mass of each element (amu) is also known as its atomic weight. Although the term average atomic mass is more proper, the term atomic weight has become common.

Formula and Molecular Weights

The formula weight of a substance is the sum of the atomic masses, for each element, found in the compound.

Exp: Calculate the formula weight of H_2SO_4

FW = 2(AW of H) + 1(AW of S) + 4(AW of)

2(1.008) + 1(32.0) + 4(15.99) = 98.0 amu

THE MOLE

A sensitive balance can weight to the nearest 0.0001 g, but a typical atom has a mass of only 0.000 000 000 000 000 000 000 01 g

Is it possible to keep track of atoms by counting them? Yes, but not directly. We count them in groups. That is we count atoms in the same way we count eggs by the dozen (12), bottle rockets by the gross (144), beer by the case (24), and sheets of paper by the ream (500).

Experiments have shown that 12.01 g of carbon contains 6.02×10^{23} atoms of carbon.

In fact, the gram atomic mass of each element contains $6.02x10^{23}$ atoms of that element.

MOLE:

Amount of a substance containing $6.02x10^{23}$ particles

AVOGADRO'S #:

Number of atoms in 12.01 grams of carbon

6.02×10^{23} atoms

Molar Mass

The molar mass of an element is equal to its formula weight expressed in grams.

The molar mass of a compound is equal to the sum of the formula weight, for each element, found in the compound.

$CuSO_4$ has a molar mass of 159.5g/mol

63.5 + 32 + 4(16) = 159.5

MOLAR MASS:

Q: What is the mass of 6.02×10^{23} H_2O molecules?

A: 18.015 g

Empirical formula Simplest formula for a compound; it is the smallest whole-number ratio of the atoms present.

Molecular formula Expression of the formula for a compound; it shows the actual number of atoms of each element present in one molecule of the compound.

Percent Composition

The percent composition is a list of the mass percent of each element in a compound.

Calculate the percent composition of proply chloride (C_3H_7Cl)

3(12.01)+7(1.01)+1(35.45)=78.55g/mol

(36.03/78.55) x 100 = 45.87% carbon

(7.07/78.55) x 100 = 9.00% hydrogen

(35.45/78.55) x 100 = 45.13% chlorine

Calculate the empirical formula for a compound composed of 26.6% potassium, 35.4% chromium, and 38.1% oxygen.

Percent means a fraction of 100. Thus, if we assume a100 grams of sample we can then use the percent values as grams of each element.

From grams we find moles of each element in the formula.

$$26.6\ g\ K\left(\frac{1\ mol\ K}{39.10\ g}\right) = 0.6803\ moles\ K$$

$$35.4g\ Cr\left(\frac{1\ molCr}{52.00g}\right) = 0.6808molesCr$$

$$38.1\ g\ O\left(\frac{1\ mol\ O}{16.00\ g}\right) = 2.381\ moles\ O$$

Next, set up a mole ratio that relates the moles of each element to the moles of the element that is least present.

$$K\left(\frac{0.6803}{0.6803}\right)=1 \quad Cr\left(\frac{0.6808}{0.6803}\right)=1 \quad O\left(\frac{2.381}{0.603}\right)=3.5$$

Because a chemical formula must have only whole number we multiply my 2 to yield: $K_2Cr_2O_7$ as the empirical formula.

An oxide of nitrogen gave the following analysis: 3.04g of nitrogen combined with 6.95g of oxygen. Experiments show that the molecular mass of this compound is 91.7 amu. Determine its molecular formula

Find moles nitrogen $3.04\ g\ N\left(\frac{1\ mole\ N}{14.01\ g\ N}\right) = 0.217\ moles\ N$

Find moles oxygen $6.95\ g\ O\left(\frac{1\ mole\ O}{16.00\ g\ O}\right) = 0.4344\ moles\ O$

Set up mole ratio $N\left(\frac{0.217}{0.217}\right)\quad O\left(\frac{0.4344}{0.217}\right)$

$N_{1.00}O_{2\cdot002}$ or NO_2

NO_2 has a molecular mass of, 14+2(16)= 46 amu

We are told that the compound has molecular mass of this compound is 91.7 amu. Given that that the empirical formula is some multiple of the molecular, lets ratio the two molecular masses (91.7/46)=1.99

Thus, the molecular formula is a multiple of two and must be N_2O_4

Consider the following chemical equation:

$$2\ NO + O_2 \rightarrow 2\ NO_2$$

The coefficients in the balanced equation indicate that 2 molecules of NO react with 1 molecule of O_2.

Similarly, this equation also states that 2 moles of NO react with 1 mole of O_2 to yield 2 mole of NO_2 product.

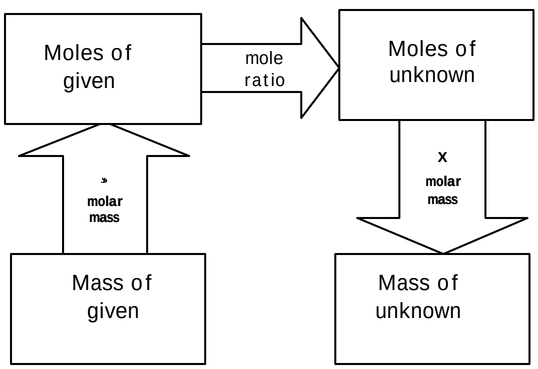
Interpretation of Chemical Equation Coefficients

For the General Equation	2 A + 3 B $\rightarrow$ C + 2 D
The ratio of molecules is	2 : 3 : 1 : 2
The ratio of moles is	2 : 3 : 1 : 2
The ratio of volumes of gas is	2 : 3 : 1 : 2

Stoichiometry 1: a branch of science that deals with the application of the laws of definite proportions and of the conservation of matter. 2 : quantitative chemical properties and composition esp. as a factor in processes of chemical or physical change

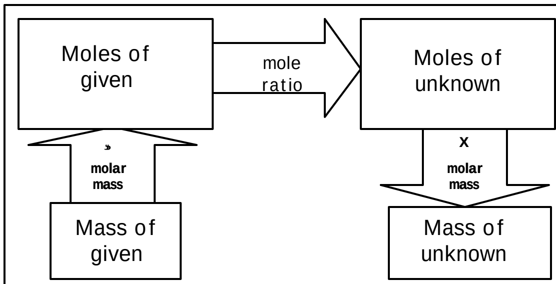
Some stoichiometric calculations relate a given amount of reactant with an unknown amount of product or vice versa. Or, a given amount of reactant can be related to another reactant. Same goes for two products.

The same method of problem solving is used regardless of the quantity sought.



How many grams of H_2 is produced when 0.500 g of Al metal react with hydrochloric acid to give aluminum chloride and hydrogen gas ?

$$2\ Al + 6\ HCl \rightarrow 2AlCl_3 + 3\ H_2$$



How many grams of H_2 is produced when 0.500 g of Al metal react with hydrochloric acid to give aluminum chloride and hydrogen gas ?

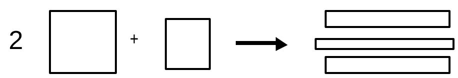
$$2\ Al + 6\ HCl \rightarrow 2AlCl_3 + 3\ H_2$$

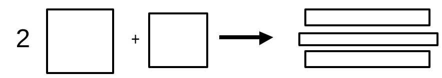
Limiting Reagent

The reactant that is completely used up in a chemical reaction.

The amount of this reactant limits the amount of new compound that can be formed.

If you have six pieces of bread and four pieces of ham, how many sandwiches can you you make?





For each reactant, calculate the amount of product that can be produced.

(6 pieces of bread) $\left(\frac{\text{1 sandwich}}{\text{2 pieces of bread}}\right)$ = 3 sandwiches

(4 pieces of ham) $\left(\frac{\text{1 sandwich}}{\text{1 pieces of ham}}\right)$ = 4 sandwiches

If 0.654 g of hydrogen react with 0.500 g of oxygen how many grams of water are produced? Which is the limiting reactant hydrogen or oxygen?

$$2H_2 + O_2 \rightarrow 2\ H_2O$$

For each reactant, calculate the amount of product that can be produced

First, the mass of water based on hydrogen.

$$0.654\ g\ H_2\left(\frac{1\ mol\ H_2}{2.02\ g\ H_2}\right)\left(\frac{2\ mol\ H_2O}{2\ mol\ H_2}\right)\left(\frac{18.01\ g\ H_2O}{1\ mol\ H_2O}\right) = 5.83\ g\ H_2O$$

Now, the mass of water based on oxygen

$$0.500\ g\ O_2\left(\frac{1\ mol\ O_2}{32.00\ g\ O_2}\right)\left(\frac{2\ mol\ H_2O}{1\ mol\ O_2}\right)\left(\frac{18.01\ g\ H_2O}{1\ mol\ H_2O}\right) = 0.563\ g\ H_2O$$

Now one sees that only 0.563 g of water can be produced and oxygen is the limiting reagent.

Be sure you know what a mole is.

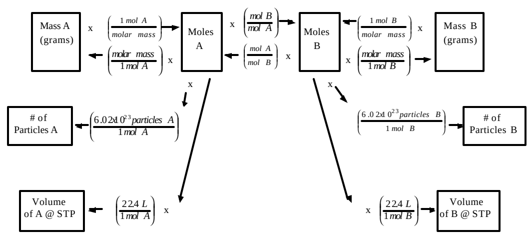
1 mole = $6.022*10^{23}$ particles = Avogadro's number

One atom of Au has a mass of 196.97 amu
6.022×10^{23} atoms of Au has a mass of 196.97 g

Why do we need this tool (the mole)?

Chemists and many others, need to know relative amounts of substances.

MOLE MAP



Relate moles to particles and particles to moles.

How many H_2 molecules are there in 0.25 moles of H_2?

(0.025 mol H_2) $\left(\frac{6.022\text{x}10^{23}\text{ molecules}}{\text{mole } H_2}\right)$ = 1.5 x10^{23} molecules of H_2

7.9 billion carbon atoms represents how many moles of carbon?

(7.9x10^{9} C atoms) $\left(\frac{1\text{ mol C}}{6.022 \text{ x } 10^{23}\text{ C atoms}}\right)$=1.3 x$10^{-14}$ mol C