**STOICHIOMETRY**

Stoichiometry is the calculation of relative quantities of reactants and products in [chemical reactions](http://en.wikipedia.org/wiki/Chemical_reaction). Stoichiometry is founded on the [law of conservation of mass](http://en.wikipedia.org/wiki/Law_of_conservation_of_mass) where the total mass of the reactants equals the total mass of the products leading to the insight that the relations among quantities of reactants and products typically form a ratio of positive integers. This means that if the amounts of the separate reactants are known, then the amount of the product can be calculated. Conversely, if one reactant has a known quantity and the quantity of product can be empirically determined, then the amount of the other reactants can also be calculated.

Stoichiometry can take several forms. Three very common forms of stoichiometry are *reaction stoichiometry*, *composition stoichiometry*, and *gas stoichiometry*. The most common type of stoichiometry studied in introductory chemistry is reaction stoichiometry (more commonly known as balancing equations).

For example, the balanced reaction between methane and oxygen is:

CH4 + 2O2 → CO2 + 2H2O

Here, one molecule of methane reacts with two molecules of [oxygen](http://en.wikipedia.org/wiki/Oxygen) gas to yield one molecule of [carbon dioxide](http://en.wikipedia.org/wiki/Carbon_dioxide) and two molecules [water](http://en.wikipedia.org/wiki/Properties_of_water). Stoichiometry measures these quantitative relationships, and is used to determine the amount of products/reactants that are produced/needed in a given reaction. Describing the quantitative relationships among substances as they participate in chemical reactions is known as ***reaction stoichiometry***. In the example above, reaction stoichiometry measures the relationship between the methane and oxygen as they react to form carbon dioxide and water.

Because of the well-known relationship of moles to atomic weights, the ratios that are arrived at by stoichiometry can be used to determine quantities by weight in a reaction described by a balanced equation. This is called ***composition stoichiometry***. For example, the stoichiometry of hydrogen and oxygen in H2O is 2:1. In stoichiometric compounds, the molar proportions are whole numbers.

***Gas stoichiometry*** deals with reactions involving gases, where the gases are at a known temperature, pressure, and volume, and can be assumed to be [ideal gases](http://en.wikipedia.org/wiki/Ideal_gas). For gases, the volume ratio is ideally the same by the [ideal gas law](http://en.wikipedia.org/wiki/Ideal_gas_law), but the mass ratio of a single reaction has to be calculated from the [molecular masses](http://en.wikipedia.org/wiki/Molecular_mass) of the reactants and products. In practice, due to the existence of [isotopes](http://en.wikipedia.org/wiki/Isotope), [molar masses](http://en.wikipedia.org/wiki/Molar_mass) are used instead when calculating the mass ratio.

**Molecular Mass**

The molecular mass of a substance is *the sum of the atomic masses of all the atoms in a molecule of the substance.* It is, therefore, the average mass of a molecule of that substance, expressed in atomic mass units. For example, the molecular mass of water, H2O, is 18.0 amu (2 × 1.0 amu from two H atoms plus 16.0 amu from one O atom).

**Formula Mass**

The formula mass of a substance is *the sum of the atomic masses of all atoms in a formula unit of the compound,* whether molecular or not. Sodium chloride, with the formula unit NaCl, has a formula mass of 58.44 amu (22.99 amu from Na plus 35.45 amu from Cl). NaCl is ionic, so strictly speaking the expression “molecular mass of NaCl” has no meaning. On the other hand, the molecular mass and the formula mass calculated from the molecular formula of a substance are identical.

**The Mole Concept**

A mole (symbol mol) is defined as *the quantity of a given substance that contains as many molecules or formula units as the number of atoms in exactly 12 g of carbon-12.* One mole of ethanol, for example, contains the same number of ethanol molecules as there are carbon atoms in 12 g of carbon-12. *The number of atoms in a 12-g sample of carbon-12* is called Avogadro’s number (NA). Recent measurements of this number give the value 6.0221367 × 1023, to which three significant figures is 6.02 ×1023.

**Converting grams to moles**

Stoichiometry is not only used to balance chemical equations but also used in conversions, i.e., converting from grams to moles, or from grams to milliliters. For example, to find the number of moles in 2.00 g of NaCl, one would do the following:

\frac{2.00 \mbox{ g NaCl}}{58.44 \mbox{ g NaCl mol}^{-1}} = 0.034 \ \text{mol}

In the above example, when written out in fraction form, the units of grams form a multiplicative identity, which is equivalent to one (g/g=1), with the resulting amount of moles (the unit that was needed), is shown in the following equation,

\left(\frac{2.00 \mbox{ g NaCl}}{1}\right)\left(\frac{1 \mbox{ mol NaCl}}{58.44 \mbox{ g NaCl}}\right) = 0.034\ \text{mol}

**Molar proportions**

Stoichiometry is often used to balance chemical equations (reaction stoichiometry). For example, the two [diatomic](http://en.wikipedia.org/wiki/Diatomic_molecule) gases, [hydrogen](http://en.wikipedia.org/wiki/Hydrogen) and [oxygen](http://en.wikipedia.org/wiki/Oxygen), can combine to form liquid water, in an [exothermic reaction](http://en.wikipedia.org/wiki/Exothermic_reaction), as described by the following equation:

2H2 + O2 → 2H2O

Reaction stoichiometry describes the 2:1:2 ratio of hydrogen, oxygen, and water molecules in the above equation. The molar ratio allows for conversion between moles of one substance and moles of another. For example, in the following reaction the number of moles of water that will be produced by the combustion of 0.27 moles of CH3OH is obtained using the molar ratio between CH3OH and H2O of 2 to 4.

2CH3OH + 3O2 → 2CO2 + 4H2O

\left(\frac{0.27 \mbox{ mol }\mathrm{CH_3OH}}{1}\right)\left(\frac{4 \mbox{ mol }\mathrm{H_2O}}{2 \mbox{ mol } \mathrm{CH_3OH}}\right) = 0.54\ \text{mol}

**Determining amount of product**

Stoichiometry can also be used to find the quantity of a product yielded by a reaction. If a piece of solid [copper](http://en.wikipedia.org/wiki/Copper) (Cu) were added to an aqueous solution of [silver nitrate](http://en.wikipedia.org/wiki/Silver_nitrate), AgNO3, the [silver](http://en.wikipedia.org/wiki/Silver) (Ag) would be replaced in a [single displacement reaction](http://en.wikipedia.org/wiki/Single_displacement_reaction) forming aqueous [copper(II) nitrate](http://en.wikipedia.org/wiki/Copper(II)_nitrate), Cu(NO3)2 and solid silver.

How much silver is produced if 16.00 grams of Cu is added to the solution of excess silver nitrate? The following steps would be used:

1. Write and Balance the Equation
2. Mass to Mole: Convert g Cu to moles Cu
3. Mole Ratio: Convert moles of Cu to moles of Ag produced
4. Mole to Mass: Convert moles Ag to grams of Ag produced

The complete balanced equation would be:

Cu + 2AgNO3 → Cu(NO3)2 + 2Ag

For the mass to mole step, the amount of copper (16.00 g) would be converted to moles of copper by dividing the mass of copper by its [molecular mass](http://en.wikipedia.org/wiki/Molecular_mass): 63.55 g/mol.

\left(\frac{16.00 \mbox{ g Cu}}{1}\right)\left(\frac{1 \mbox{ mol Cu}}{63.55 \mbox{ g Cu}}\right) = 0.2518\ \text{mol Cu}

Now that the amount of Cu in moles (0.2518) is found, we can set up the mole ratio. This is found by looking at the coefficients in the balanced equation: Cu and Ag are in a 1:2 ratio.

\left(\frac{0.2518 \mbox{ mol Cu}}{1}\right)\left(\frac{2 \mbox{ mol Ag}}{1 \mbox{ mol Cu}}\right) = 0.5036\ \text{mol Ag}

Now that the moles of Ag produced is known to be 0.5036 mol, we convert this amount to grams of Ag produced to come to the final answer:

\left(\frac{0.5036 \mbox{ mol Ag}}{1}\right)\left(\frac{107.87  \mbox{ g Ag}}{1 \mbox{ mol Ag}}\right) = 54.32 \ \text{g Ag}

This set of calculations can be further condensed into a single step:

m_\mathrm{Ag} = \left(\frac{16.00 \mbox{ g }\mathrm{Cu}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{Cu}}{63.55 \mbox{ g }\mathrm{Cu}}\right)\left(\frac{2 \mbox{ mol }\mathrm{Ag}}{1 \mbox{ mol }\mathrm{Cu}}\right)\left(\frac{107.87 \mbox{ g }\mathrm{Ag}}{1 \mbox{ mol Ag}}\right) = 54.32 \mbox{ g}

**Further examples**

For propane (C3H8) reacting with oxygen gas (O2), the balanced chemical equation is:

C3H8 + 5O2 → 3CO2 + 4H2O

The mass of water formed if 120 g of propane (C3H8) is burned in excess oxygen is then

m_\mathrm{H_2O} = \left(\frac{120. \mbox{ g }\mathrm{C_3H_8}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{C_3H_8}}{44.09 \mbox{ g }\mathrm{C_3H_8}}\right)\left(\frac{4 \mbox{ mol }\mathrm{H_2O}}{1 \mbox{ mol }\mathrm{C_3H_8}}\right)\left(\frac{18.02 \mbox{ g }\mathrm{H_2O}}{1 \mbox{ mol }\mathrm{H_2O}}\right) = 196 \mbox{ g}

**Stoichiometric ratio**

Stoichiometry is also used to find the right amount of one reactant to "completely" react with the other reactant in a chemical reaction - that is, the stoichiometric amounts that would result in no leftover reactants when the reaction takes place. An example is shown below using the [reaction](http://en.wikipedia.org/wiki/Thermite_reaction):

Fe2O3 + 2Al → Al2O3 + 2Fe

This equation shows that 1 mole of [iron(III) oxide](http://en.wikipedia.org/wiki/Iron(III)_oxide) and 2 moles of [aluminum](http://en.wikipedia.org/wiki/Aluminum) will produce 1 mole of [aluminium oxide](http://en.wikipedia.org/wiki/Aluminium_oxide" \o "Aluminium oxide) and 2 moles of [iron](http://en.wikipedia.org/wiki/Iron). So, to completely react with 85.0 g of iron(III) oxide (0.532 mol), 28.7 g (1.06 mol) of aluminium are needed.

m_\mathrm{Al} = \left(\frac{85.0 \mbox{ g }\mathrm{Fe_2O_3}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{Fe_2 O_3}}{159.7 \mbox{ g }\mathrm{Fe_2 O_3}}\right)\left(\frac{2 \mbox{ mol Al}}{1 \mbox{ mol }\mathrm{Fe_2 O_3}}\right)\left(\frac{26.98 \mbox{ g Al}}{1 \mbox{ mol Al}}\right) = 28.7 \mbox{ g}

**Limiting reagent and percent yield**

The **limiting reagent** is the reagent that limits the amount of product that can be formed and is completely consumed during the reaction. The **excess reactant** is the reactant that is left over once the reaction has stopped due to the limiting reactant.

Consider the equation of roasting [lead(II) sulfide](http://en.wikipedia.org/wiki/Lead(II)_sulfide) (PbS) in oxygen (O2) to produce [lead(II) oxide](http://en.wikipedia.org/wiki/Lead(II)_oxide) (PbO) and [sulfur dioxide](http://en.wikipedia.org/wiki/Sulfur_dioxide) (SO2):

2 PbS + 3 O2 → 2 PbO + 2 SO2

To determine the theoretical yield of lead(II) oxide if 200.0 g of lead(II) sulfide and 200.0 grams of oxygen are heated in an open container:

m_\mathrm{PbO} = \left(\frac{200.0 \mbox{ g }\mathrm{PbS}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{PbS}}{239.27 \mbox{ g }\mathrm{PbS}}\right)\left(\frac{2 \mbox{ mol }\mathrm{PbO}}{2 \mbox{ mol }\mathrm{PbS}}\right)\left(\frac{223.2 \mbox{ g }\mathrm{PbO}}{1 \mbox{ mol }\mathrm{PbO}}\right) = 186.6 \mbox{ g}

m_\mathrm{PbO} = \left(\frac{200.0 \mbox{ g }\mathrm{O_2}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{O_2}}{32.00 \mbox{ g }\mathrm{O_2}}\right)\left(\frac{2 \mbox{ mol }\mathrm{PbO}}{3 \mbox{ mol }\mathrm{O_2}}\right)\left(\frac{223.2 \mbox{ g }\mathrm{PbO}}{1 \mbox{ mol }\mathrm{PbO}}\right) = 930.0 \mbox{ g}

Because a lesser amount of PbO is produced for the 200.0 g of PbS, it is clear that PbS is the limiting reagent.

In reality, the actual yield is not the same as the stoichiometrically-calculated theoretical yield. **Percent yield**, then, is expressed in the following equation:

\mbox{percent yield} = \frac{\mbox{actual yield}}{\mbox{theoretical yield}} \times \!\, 100

If 170.0 g of lead(II) oxide is obtained, then the percent yield would be calculated as follows:

\mbox{percent yield} = \frac{\mbox{170.0 g PbO}}{\mbox{186.6 g PbO}} \times \!\, 100 = 91.12\%

**Example**

Consider the following reaction, in which [iron(III) chloride](http://en.wikipedia.org/wiki/Iron(III)_chloride) reacts with [hydrogen sulfide](http://en.wikipedia.org/wiki/Hydrogen_sulfide) to produce [iron(III) sulfide](http://en.wikipedia.org/wiki/Iron(III)_sulfide) and [hydrogen chloride](http://en.wikipedia.org/wiki/Hydrogen_chloride):

2 FeCl3 + 3 H2S → Fe2S3 + 6 HCl

Suppose 90.0 g of FeCl3 reacts with 52.0 g of H2 S. To find the limiting reagent and the mass of HCl produced by the reaction, we could set up the following equations:

m_\mathrm{HCl} = \left(\frac{90.0 \mbox{ g }\mathrm{FeCl_3}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{FeCl_3}}{162 \mbox{ g }\mathrm{FeCl_3}}\right)\left(\frac{6 \mbox{ mol }\mathrm{HCl}}{2 \mbox{ mol }\mathrm{FeCl_3}}\right)\left(\frac{36.5 \mbox{ g }\mathrm{HCl}}{1 \mbox{ mol }\mathrm{HCl}}\right) = 60.8 \mbox{ g}

m_\mathrm{HCl} = \left(\frac{52.0 \mbox{ g }\mathrm{H_2S}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{H_2S}}{34.1 \mbox{ g }\mathrm{H_2S}}\right)\left(\frac{6 \mbox{ mol }\mathrm{HCl}}{3 \mbox{ mol }\mathrm{H_2S}}\right)\left(\frac{36.5 \mbox{ g }\mathrm{HCl}}{1 \mbox{ mol }\mathrm{HCl}}\right) = 111 \mbox{ g}

Thus, the limiting reagent is FeCl3 and the amount of HCl produced is 60.8 g.

To find what mass of excess reagent (H2S) remains after the reaction, we would set up the calculation to find out how much H2S reacts completely with the 90.0 g FeCl3:

m_\mathrm{H_2S} = \left(\frac{90.0 \mbox{ g }\mathrm{FeCl_3}}{1}\right)\left(\frac{1 \mbox{ mol }\mathrm{FeCl_3}}{162 \mbox{ g }\mathrm{FeCl_3}}\right)\left(\frac{3 \mbox{ mol }\mathrm{H_2S}}{2 \mbox{ mol }\mathrm{FeCl_3}}\right)\left(\frac{34.1 \mbox{ g }\mathrm{H_2S}}{1 \mbox{ mol }\mathrm{H_2S}}\right) = {28.4 \mbox{ g }\mathrm{reacted}}

By subtracting this amount from the original amount of H2S, we can come to the answer:

{\mathrm{excess}}{52.0 \mbox{ g }\mathrm{H_2S}} - {28.4 \mbox{ g }\mathrm{H_2S}} = {23.6 \mbox{ g }\mathrm{H_2S}}