Stoichiometry Formula Guide

Mole Constants:

1 Mole = 6.022 x 10^{23} items (atoms, molecules, formula units, etc.); called **Avogadro's**
1 Mole= 22.4 Liters of any gas @ STP (according to ideal gas law)
1 Mole = molar mass (total # of grams) from the periodic table

Definitions:

Molecule = two or more covalently bonded atoms (non-metal to non-metal)
Compound (also called **Formula Unit)** = two or more ionically bonded atoms (metal to non-metal)
Molar Mass = combined mass of each element from the periodic table

Ex. Molar Mass of $Na_{\mathbf{2}}SO_{\mathbf{4}}$ = (**2** x 23 amu) + (**1** x 32 amu) + (**4** x 16 amu) = **142 g/mole**

Mole Ratio = compares coefficients of different substances based on a balanced chemical equation; can be written as a fraction or ratio using a colon

Ex. **2** C_4H_{10} + **13** O_2 → **8** CO_2 + **10** H_2O

The ratio of butane to carbon dioxide can be written in the following ways:

$$\frac{\textbf{2 moles } \mathbf{C_4H_{10}}}{\textbf{8 moles } \mathbf{CO_2}} \quad \text{or} \quad \frac{\textbf{8 moles } \mathbf{CO_2}}{\textbf{2 moles } \mathbf{C_4H_{10}}}$$

One Step Equations:

1. Atoms*-to-Moles

$$\frac{\#\text{atom}}{1}\left|\frac{1\text{ mole}}{6.022 \times 10^{23}\text{ atoms}}\right| = \textbf{\# moles}$$

Orange* = **Avogadro's #; 6.022 x 10^{23}** refers to the # of ***atoms or molecules*** in one mole of any substance.

2. Moles-to-Atoms*

$$\frac{\#\text{ moles}}{1}\left|\frac{6.022 \times 10^{23}\text{ atoms}}{1\text{ mole}}\right| = \textbf{\# atoms}$$

3. Grams-to-Moles

$$\frac{\#\text{ grams}}{1}\left|\frac{1\text{ mole}}{\#\text{ grams from Periodic Table}}\right| = \textbf{\# moles}$$

Blue = **Molar Mass**; refers to the ***combined grams/mole*** of all the atoms in a given substance; grams are found by using the atomic masses from the periodic table.

4. Moles-to-Grams

$$\frac{\#\text{ moles}}{1}\left|\frac{\#\text{ grams from Periodic Table}}{1\text{ mole}}\right| = \textbf{\# grams}$$

5. Moles-to-Liters

$$\frac{\#\ \text{moles}}{1}\left|\frac{22.4\ \text{Liters}}{1\ \text{mole}}\right| = \#\ \textbf{Liters}$$

6. Liters-to-Moles

$$\frac{\#\ \text{Liters}}{1}\left|\frac{1\ \text{mole}}{22.4\ \text{Liters}}\right| = \#\ \textbf{moles}$$

Purple = 22.4 Liters; the volume of any gas @ STP, which means at standard temperature and pressure units.

NOTE: Remember 6.022 x 10^{23} refers to the number of ***atoms or molecules*** in 1 mole of any substance.

Two Step Equations:

7. Atoms* – to – Grams

$$\frac{\#\ \text{atoms}}{1}\left|\frac{1\ \text{mole}}{6.022\ \times 10^{23}\ \text{atoms}}\right|\frac{\#\ \text{grams from Periodic Table}}{1\ \text{mole}}\right| = \#\ \textbf{grams}$$

8. Grams – to – Atoms*

$$\frac{\#\ \text{grams}}{1}\left|\frac{1\ \text{mole}}{\#\ \text{grams from Periodic Table}}\right|\frac{6.022\ \times 10^{23}\ \text{atoms}}{1\ \text{mole}}\right| = \#\ \textbf{atoms}$$

9. Atoms* – to – Liters

$$\frac{\#\ \text{atoms}}{1}\left|\frac{1\ \text{mole}}{6.022\ \times 10^{23}\ \text{atoms}}\right|\frac{22.4\ \text{Liters}}{1\ \text{mole}}\right| = \#\ \textbf{Liters}$$

10. Liters – to – Atoms*

$$\frac{\#\ \text{Liters}}{1}\left|\frac{1\ \text{mole}}{22.4\ \text{Liters}}\right|\frac{6.022\ \times 10^{23}\ \text{atoms}}{1\ \text{mole}}\right| = \#\ \textbf{atoms}$$

11. Grams – to – Liters

$$\frac{\#\ \text{grams}}{1}\left|\frac{1\ \text{mole}}{\#\ \text{grams from Periodic Table}}\right|\frac{22.4\ \text{Liters}}{1\ \text{mole}}\right| = \#\ \textbf{Liters}$$

12. Liters – to – Grams

$$\frac{\#\ \text{Liters}}{1}\left|\frac{1\ \text{mole}}{22.4\ \text{Liters}}\right|\frac{\#\ \text{grams from Periodic Table}}{1\ \text{mole}}\right| = \#\ \textbf{grams}$$

Conversion Requiring Mole Ratios:

- Mole Ratios are used to compare **one reactant to another reactant** or **a reactant to a product**
- Mole Ratios use the **coefficients** from balanced chemical equations.
- Coefficients are the **BIG** numbers in front used to balance the chemical equation.

Green = Mole Ratios; compares the ***coefficients*** from different substances based on the balanced chemical equation; coefficients are the big # in front of each substance in the balanced chemical equation.

One Step Mole Ratio:

12. Moles A– to – Moles B

$$\frac{\#\text{ moles A}}{1}\left|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right| = \#\textbf{ moles B}$$

Two Step Mole Ratio:

13. Moles A – to – Liters B

$$\frac{\#\text{ moles A}}{1}\left|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right|\frac{\text{22.4 Liters B}}{\text{1 mole B}}\right| = \#\textbf{ Liters B}$$

14. Liters A – to – Moles B

$$\frac{\#\text{ Liters A}}{1}\left|\frac{\text{1 mole A}}{\text{22.4 Liters A}}\right|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right| = \#\textbf{ moles B}$$

15. Moles A – to – Grams B

$$\frac{\#\text{ moles A}}{1}\left|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right|\frac{\#\text{ grams B from Periodic Table}}{\text{1 mole B}}\right| = \#\textbf{ grams B}$$

16. Grams A – to – Moles B

$$\frac{\#\text{ grams A}}{1}\left|\frac{\text{1 mole A}}{\#\text{ grams A from Periodic Table}}\right|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right| = \#\textbf{ moles B}$$

Three Step Mole Ratio:

17. Grams A – to – Grams B

$$\frac{\#\text{ grams A}}{1}\left|\frac{\text{1 mole A}}{\#\text{ grams A from Periodic Table}}\right|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right|\frac{\#\text{ grams B from Periodic Table}}{\text{1 mole B}}\right| = \#\textbf{ grams B}$$

18. Grams A – to – Liters B

$$\frac{\#\text{ grams A}}{1}\left|\frac{\text{1 mole A}}{\#\text{ grams A from Periodic Table}}\right|\frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}}\right|\frac{\text{22.4 Liters B}}{\text{1 mole B}}\right| = \#\textbf{ Liters B}$$

Limiting Reactant, Excess Reactant, & Theoretical Yield

- When a chemical reaction occurs, sometimes one reactant will be in **excess** and the other will **run out** so that the reaction cannot happen anymore. When this happens the reactant that ran out is called the "**limiting reactant**".

- In order to find the limiting reactant, you must <u>calculate how much product can be made</u> from the starting amount of each reactant. The "limiting reactant" will be the one that makes **LESS** product.

- The **theoretical yield** is the amount of **product** that can be made once the limiting reactant runs out.

- Use **Equation #18** or **Equation #19** to find the limiting reactant and theoretical yield.

Example:

a) If 12.0 grams of iron (II) chloride reacts with 23.0 grams of sodium hydroxide, determine the limiting reagent and the theoretical yield?

$$1\ FeCl_2\ (aq) + 2\ NaOH\ (aq) \rightarrow 1\ Fe(OH)_2\ (s) + 2\ NaCl\ (aq)$$

$$\frac{12.0 \text{ gram } \mathbf{FeCl_2}}{1}\left|\frac{1 \text{ mole } FeCl_2}{126.753 \text{ gram } FeCl_2}\right|\frac{1 \text{ mole } Fe(OH)_2}{1 \text{ mole } FeCl_2}\left|\frac{89.861 \text{ grams } Fe(OH)_2}{1 \text{ mole } Fe(OH)_2}\right| = \mathbf{8.51 \text{ grams } Fe(OH)_2}$$

$$\frac{23.0 \text{ gram NaOH}}{1}\left|\frac{1 \text{ mole NaOH}}{39.997 \text{ gram NaOH}}\right|\frac{1 \text{ mole } Fe(OH)_2}{2 \text{ mole NaOH}}\left|\frac{89.861 \text{ grams } Fe(OH)_2}{1 \text{ mole } Fe(OH)_2}\right| = \mathbf{25.8 \text{ grams } Fe(OH)_2}$$

- **$FeCl_2$** (iron II chloride) is the **limiting reactant** because it produces **LESS** iron II hydroxide, $Fe(OH)_2$.

- **8.51 grams $Fe(OH)_2$** is the theoretical yield of **product** because it is based on the limiting reactant.

b) Determine the amount of excess reactant remaining once the reaction stops?

First you must determine how much excess reactant was used by comparing the limiting reactant to the excess reactant. Then you can subtract the amount used from the starting amount to determine the amount leftover.

$$\frac{12.0 \text{ gram } \mathbf{FeCl_2}}{1}\left|\frac{1 \text{ mole } FeCl_2}{126.753 \text{ gram } FeCl_2}\right|\frac{2 \text{ mole NaOH}}{1 \text{ mole } FeCl_2}\left|\frac{39.997 \text{ grams NaOH}}{1 \text{ mole NaOH}}\right| = \mathbf{7.57 \text{ grams NaOH used}}$$

23.0 grams NaOH starting – 7.57 grams NaOH used = 15.4 grams NaOH leftover

Percent Yield

- When chemists actually perform experiments in the lab, they rarely make 100% of the product. Factors that can cause a reaction not to happen completely are:

  1) The limiting reactant can **stick to the walls** of the container causing less of it to react.
  2) The temperature and/or pressure may be less than ideal causing **ineffective particle collisions**.

- In order to determine the **percent yield**, a chemist must compare the actual amount made in the lab to the **theoretical yield**.

Example:

a) A chemist reacts 12.0 grams of iron (II) chloride with 23.0 grams of sodium hydroxide. Determine the percent yield, if 7.95 grams of iron II hydroxide are obtained.

$$1\ \mathbf{FeCl_2\ (aq)} + 2\ \mathbf{NaOH\ (aq)} \rightarrow 1\ \mathbf{Fe(OH)_2\ (s)} + 2\ \mathbf{NaCl\ (aq)}$$

$$\frac{12.0\text{ gram }\mathbf{FeCl_2}}{1}\left|\frac{1\text{ mole }\mathrm{FeCl_2}}{126.753\text{ gram }\mathrm{FeCl_2}}\right|\frac{1\text{ mole }\mathrm{Fe(OH)_2}}{1\text{ mole }\mathrm{FeCl_2}}\left|\frac{89.861\text{ grams }\mathrm{Fe(OH)_2}}{1\text{ mole }\mathrm{Fe(OH)_2}}\right| = \mathbf{8.51\ grams\ Fe(OH)_2}$$

- ➢ As determined earlier, **$FeCl_2$** is the **limiting reactant**. The total amount of product it can possibly produce is **8.51 grams $Fe(OH)_2$**. This is the called the ***theoretical yield*** of product.

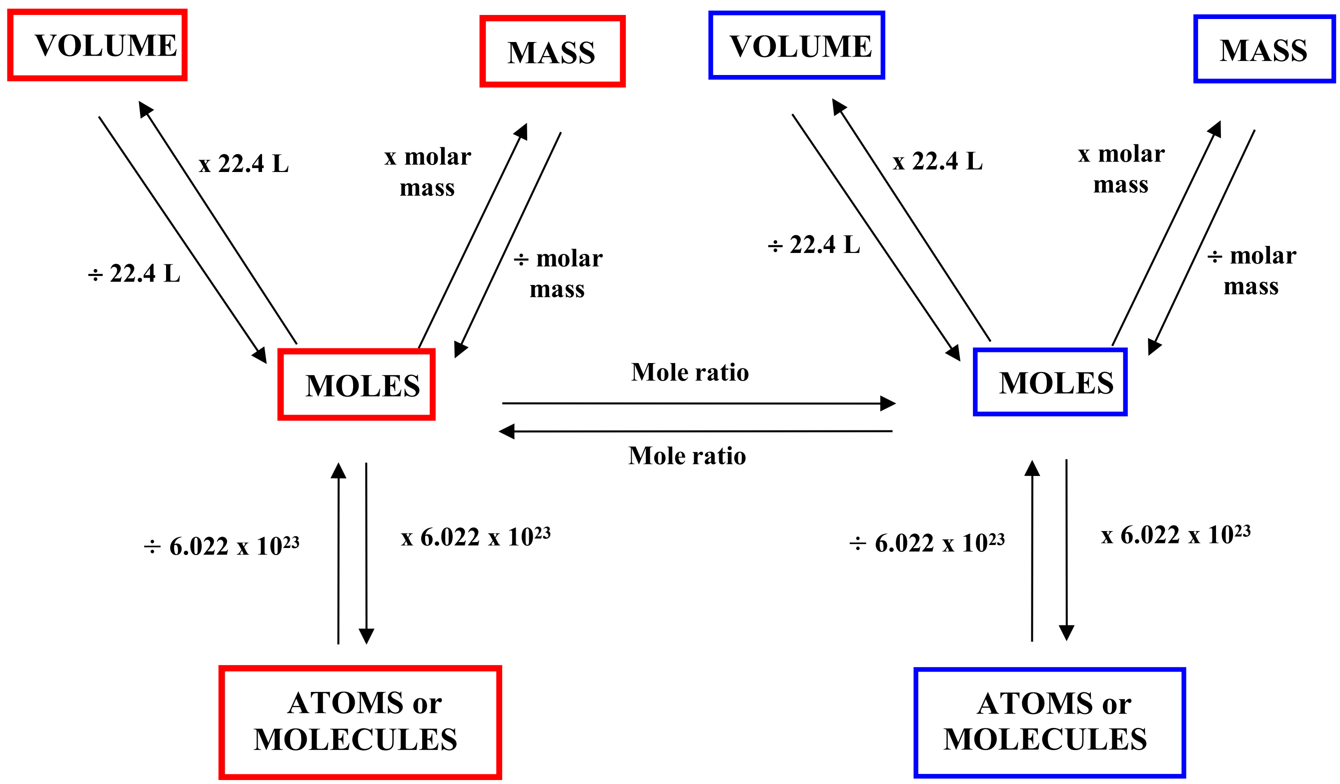
- ➢ To determine the ***percent yield*** for the situation given above, use the following equation:

$$\frac{\textbf{Amount Obtained in the Lab}}{\textbf{Theoretical Yield}} \times \mathbf{100} = \textbf{Percent Yield}$$

- ➢ So for this situation the ***percent yield*** is

$$\frac{\mathbf{7.95\ grams\ Fe(OH)_2}}{\mathbf{8.51\ grams\ Fe(OH)_2}} \times \mathbf{100} = \mathbf{93.4\ \%}$$

The Roadmap to Stoichiometry

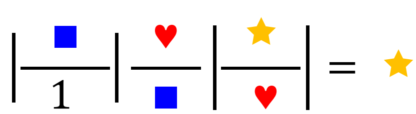


Remember:

The direction of stoichiometry depends on where you start and where you want to go.

Always start with the unit given in the word problem and follow the path to the unknown unit.

The pattern looks like this:



If you are only using **1 substance**, you only use the *left side* of the roadmap.

If you are comparing **one substance** to the **other**, you must use *both sides* of the roadmap.