

Stoichiometry Formula Guide

Mole Constants:

1 Mole = 6.022×10^{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_2SO_4 = $(2 \times 23 \text{ amu}) + (1 \times 32 \text{ amu}) + (4 \times 16 \text{ amu}) = 142 \text{ g/mole}$

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. $2 \text{ C}_4\text{H}_{10} + 13 \text{ O}_2 \rightarrow 8 \text{ CO}_2 + 10 \text{ H}_2\text{O}$

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

$$\frac{2 \text{ moles C}_4\text{H}_{10}}{8 \text{ moles CO}_2} \quad \text{or} \quad \frac{8 \text{ moles CO}_2}{2 \text{ moles C}_4\text{H}_{10}}$$

One Step Equations:

1. Atoms*-to-Moles

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

2. Moles-to-Atoms*

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

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

3. Grams-to-Moles

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

4. Moles-to-Grams

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

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.

5. Moles-to-Liters

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

6. Liters-to-Moles

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

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

Two Step Equations:

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

7. Atoms* – to – Grams

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

8. Grams – to – Atoms*

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

9. Atoms* – to – Liters

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

10. Liters – to – Atoms*

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

11. Grams – to – Liters

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

12. Liters – to – Grams

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

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.

One Step Mole Ratio:

12. Moles A – to – Moles B

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

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.

Two Step Mole Ratio:

13. Moles A – to – Liters B

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

14. Liters A – to – Moles B

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

15. Moles A – to – Grams B

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

16. Grams A – to – Moles B

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

Three Step Mole Ratio:

17. Grams A – to – Grams B

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

18. Grams A – to – Liters B

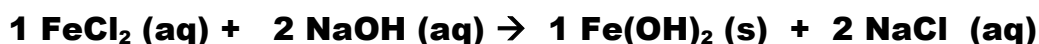
$$\frac{\# \text{ grams A}}{1} \left| \frac{1 \text{ mole A}}{\# \text{ grams A from Periodic Table}} \right| \left| \frac{\text{Mole Coefficient of B}}{\text{Mole Coefficient of A}} \right| \left| \frac{22.4 \text{ Liters B}}{1 \text{ mole B}} \right| = \# \text{ Liters B}$$

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 calculate how much product can be made 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?



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

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

- **FeCl₂** (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.

$$\frac{12.0 \text{ gram FeCl}_2}{1} \left| \frac{1 \text{ mole FeCl}_2}{126.753 \text{ gram FeCl}_2} \right| \left| \frac{2 \text{ mole NaOH}}{1 \text{ mole FeCl}_2} \right| \left| \frac{39.997 \text{ grams NaOH}}{1 \text{ mole NaOH}} \right| = 7.57 \text{ grams NaOH used}$$

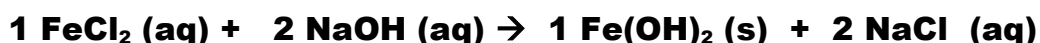
23.0 grams NaOH starting – 7.57 grams NaOH used = **15.4 grams NaOH leftover**

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:
 - The limiting reactant can **stick to the walls** of the container causing less of it to react.
 - 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.



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

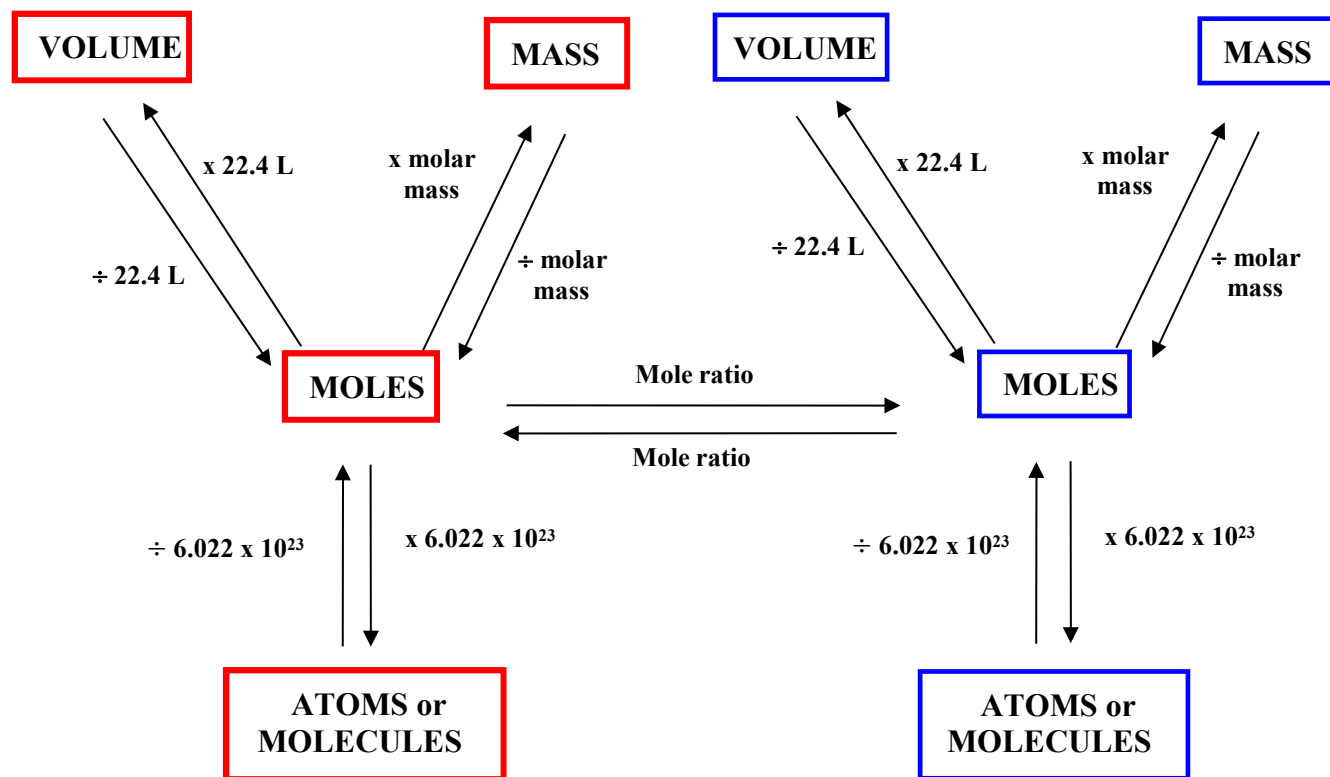
- As determined earlier, FeCl_2 is the **limiting reactant**. The total amount of product it can possibly produce is **8.51 grams Fe(OH)_2** . This is called the **theoretical yield** of product.
- To determine the **percent yield** for the situation given above, use the following equation:

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

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

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

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:

$$\left| \frac{\text{blue square}}{1} \right| \left| \frac{\text{red heart}}{\text{blue square}} \right| \left| \frac{\text{yellow star}}{\text{red heart}} \right| = \text{yellow star}$$

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.