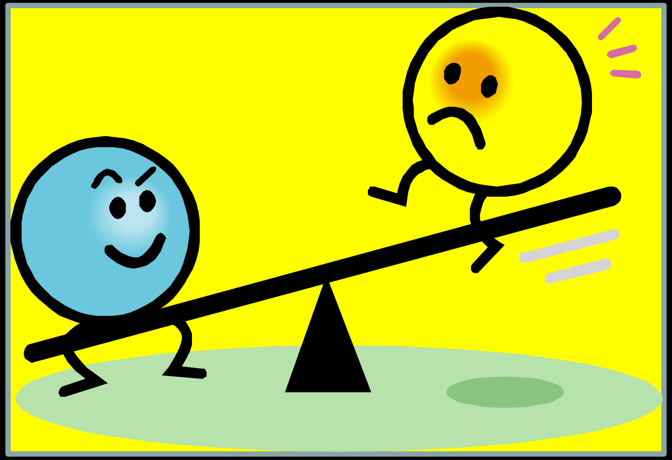

Stoichiometry

Chemistry

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By end of this unit I can…

MS10: identify the molar ratios of a balanced chemical equation.

MS11: calculate molar relationships of a chemical equation.

MS12: determine the limiting reactant in a chemical equation.

MS13: calculate the percent yield of a chemical equation.

Stoichiometry

- Unit changes: Three Molar ratios.
- Relationships between product and reactant set up new truths.
- Being able to determine the amounts and study the quantitative relationships of equations is called stoichiometry.
- The most important idea in this chapter is to look at the coefficients of the formula.

$$N_2H_4 + 2H_2O_2 \rightarrow N_2 + 4H_2O$$

Using Coefficients

$$N_2H_4 + 2H_2O_2 \rightarrow N_2 + 4H_2O$$

- This reaction needs the following:
  - 1 mole of N_2H_4 and 2 moles of H_2O_2
  - To produce…
  - 1 mole of N_2 and 4 moles of H_2O
- How many moles of H_2O are produced if you have 3.55 moles of N_2H_4? (assume you have enough Hydrogen Peroxide)

Same as it ever was

$N_2H_4 + 2H_2O_2 \rightarrow N_2 + 4H_2O$

- Use the *same* ideas from before:
  - Circle what you have and what you want…
  - Write what you know…
  - What you have on bottom (including the coefficient).
  - What you want on top (including the coefficient).

$$3.55 \text{ moles } N_2H_4 \times \frac{4 \text{ moles } H_2O}{1 \text{ mole } N_2H_4}$$

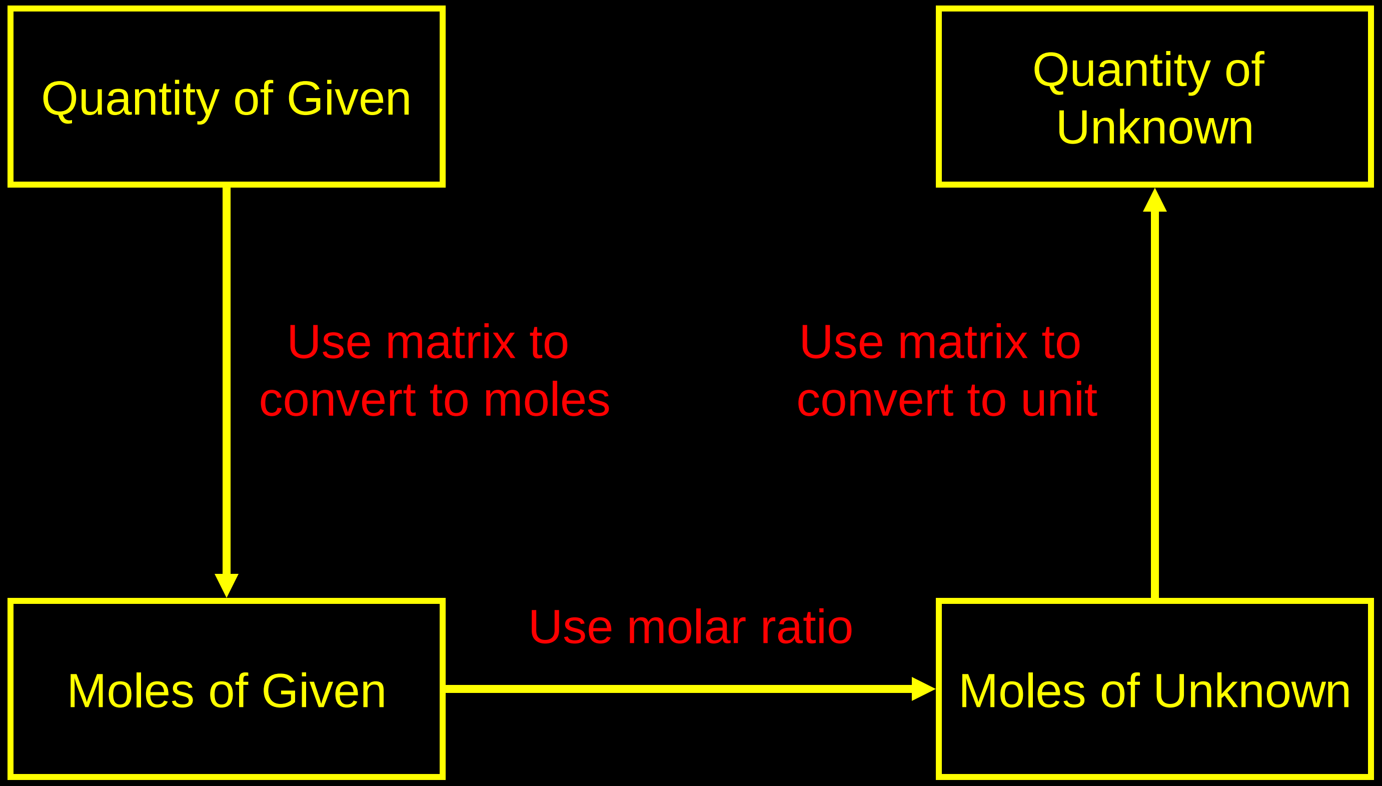
$$14.2 \text{ moles } H_2O$$

Upping the Ante

- The last problem was a mole to mole problem.
- If given grams, Liters (of gas) or units, it is still possible to solve with one new thought
  - After Writing what you know…
  - Ask ‘Is it in moles?’… if no, you must convert to moles (Chapter 10).

Big Idea

- Here is a flow chart that may help.



Thermite Reaction

- In a Thermite reaction Aluminum powder reacts with Iron (III) Oxide to form Aluminum Oxide and Molten Iron… If in a complete reaction 2.3-g of Aluminum are needed, how many grams of Aluminum Oxide are produced?
- Write what you know (and circle).

$$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$$

2.3-g ?-g

Thermite Reaction

$$\overset{2.3\text{-g}}{2Al} + Fe_2O_3 \rightarrow \overset{?\text{-g}}{Al_2O_3} + 2Fe$$

- Convert Grams to Moles…

$$2.3\text{-g Al} \times \frac{1 \text{ mole Al}}{27.0\text{-g Al}} = .0851 \text{ moles Al}$$

- Use Molar Ratio…

$$.0851 \text{ moles Al} \times \frac{1 \text{ mole } Al_2O_3}{2 \text{ moles Al}} = .04 \text{ moles } Al_2O_3$$

- Convert Back to Grams…

$$.04 \text{ moles } Al_2O_3 \times \frac{102.0\text{-g } Al_2O_3}{1 \text{ mole } Al_2O_3} = 4.4\text{-g } Al_2O_3$$

Limiting Reactants

- There are two ways to mix chemicals together.
  - Measure exactly how much of each chemical is needed to produce a reaction.
  - Put random amounts of each chemical together and see how much is produced.
- In the second case one of the chemicals will be a limiting reactant.
- Think about this… A Table requires four legs and one top. If you had 22 legs and 6 tops how many tables could you make? What part is limiting the production?

$$4Lg + 1Tt \rightarrow 1Tb$$

Visual Approach

x 6

6 Tops
=
6 Tables

x 6

x 22

22 Legs
=
5.5 Tables

x 5

The legs limit the construction of more tables.

Solving Limiting Reactants

- Write what you know
  - If not given the balanced formula
  - Circle both given amounts
- Find out how much product each reactant will produce. (double the work)
  - If not told specifically solve for any one product.
- The one that produces the least amount is the limiting reactant.

Limiting Reactant Example

- What is the limiting reactant when 1.7-g of Sodium and 2.6-L of Chlorine gas at STP produce Table Salt?

$$2Na + Cl_2 \rightarrow 2NaCl$$

1.7-g 2.6-L ?-g

$$1.7\text{-g Na} \times \frac{1\text{-mol}}{23.0\text{-g}} \times \frac{2\ NaCl}{2\ Na} \times \frac{58.5\text{-g}}{1\text{-mol}}$$

$1.7\text{-g Na} \rightarrow 4.3\text{-g NaCl}$

$$2.6\text{-L } Cl_2 \times \frac{1\text{-mol}}{22.4\text{-L}} \times \frac{2\ NaCl}{1\ Cl_2} \times \frac{58.5\text{-g}}{1\text{-mol}}$$

$2.6\text{-L } Cl_2 \rightarrow 13.6\text{-g NaCl}$

Percent Yield

- Despite best plans in any reaction the amount of products expected, is more than what is actually produced.
- It can be useful to calculate the efficiency of the formula. (much like % error in a lab).

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Expected Yield}}$$

Percent Yield Example

- A piece of Copper (5.0-g) is placed in a solution of excess Silver Nitrate. When separated 7.2-g of Silver is produced. What is the Percent Yield?

$$Cu + AgNO_3 \rightarrow Ag + CuNO_3$$

5.0-g ?-g

$$5.0\text{-g Cu} \times \frac{1\text{-mol}}{63.5\text{-g}} \times \frac{1\ Ag}{1\ Cu} \times \frac{107.9\text{-g}}{1\text{-mol}}$$

5.0-g Cu should produce 8.5-g Ag

Percent Yield Example

- So in the reaction the 5.0-g of Copper could have yielded a maximum of 8.5-g of Silver, but the problem stated only 7.2-g was produced.

$$\frac{7.2\text{-g Ag}}{8.5\text{-g Ag}} \times 100 = \boxed{84.7\%}$$

- Note the Percent Yield can never be more than 100%.