Chapter 12
"Stoichiometry"

Mr. Mole

Let’s make some Cookies!

- When baking cookies, a recipe is usually used, telling the *exact amount of each ingredient*.
  - If you need more, you can double or triple the amount
- Thus, a **recipe** is much like a **balanced equation.**

Stoichiometry is…

- Greek for "measuring elements"

Pronounced "stoy kee ahm uh tree"

- Defined as: calculations of the *quantities* in chemical reactions, based on a balanced equation.

- There are **4 ways** to *interpret* a balanced chemical equation

#1. In terms of **Particles**

- An Element is made of **atoms**
- A Molecular compound (made of only nonmetals) is made up of **molecules** (Don’t forget the diatomic elements)
- Ionic Compounds (made of a metal and nonmetal parts) are made of **formula units**

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

- Two *molecules* of hydrogen and one *molecule* of oxygen form two *molecules* of water.
- Another example: $2Al_2O_3 \rightarrow 4Al + 3O_2$

$$2 \textit{ formula units } Al_2O_3 \text{ form } 4 \textit{ atoms } Al$$

$$\text{and } 3 \textit{ molecules } O_2$$

Now read this: $2Na + 2H_2O \rightarrow 2NaOH + H_2$

#2. In terms of **<u>Moles</u>**

- The <u>coefficients</u> tell us how many moles of each substance

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

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

- Remember: A balanced equation is a ***<u>Molar Ratio</u>***

#3. In terms of **Mass**

- The Law of Conservation of Mass applies
- We can check mass by using moles.

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

$$2 \text{ moles } H_2 \left(\frac{2.02 \text{ g } H_2}{1 \text{ mole } H_2} \right) = 4.04 \text{ g } H_2$$

$$+$$

$$1 \text{ mole } O_2 \left(\frac{32.00 \text{ g } O_2}{1 \text{ mole } O_2} \right) = 32.00 \text{ g } O_2$$

36.04 g $H_2 + O_2$
reactants

In terms of Mass (for products)

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

$$2 \text{ moles } H_2O \left(\frac{18.02 \text{ g } H_2O}{1 \text{ mole } H_2O} \right) = 36.04 \text{ g } H_2O$$

$$\mathbf{36.04} \text{ g } H_2 + O_2 = \mathbf{36.04} \text{ g } H_2O$$

36.04 grams reactant = 36.04 grams product

The mass of the reactants must <u>equal</u> the mass of the products.

#4. In terms of **Volume**

- At STP, 1 mol of any gas = 22.4 L

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

$(2 \times 22.4\ L\ H_2) + (1 \times 22.4\ L\ O_2) \rightarrow (2 \times 22.4\ L\ H_2O)$

67.2 Liters of reactant **≠** 44.8 Liters of product!

NOTE: **mass and atoms** are **ALWAYS** conserved - however, molecules, formula units, moles, and volumes will not necessarily be conserved!

Practice:

- Show that the following equation follows the Law of Conservation of Mass (show the atoms balance, and the mass on both sides is equal)

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

Section 12.2
Chemical Calculations

- **OBJECTIVES:**
  - *Construct* “mole ratios” from balanced chemical equations, and *apply* these ratios in mole-mole stoichiometric calculations.

Chemical Calculations

- **OBJECTIVES:**
  - *Calculate* stoichiometric quantities from balanced chemical equations using units of moles, mass, representative particles, and volumes of gases at STP.

Mole to Mole conversions

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

- each time we use 2 moles of Al_2O_3 we will also make 3 moles of O_2

$$\left(\frac{2 \text{ moles } Al_2O_3}{3 \text{ mole } O_2}\right) \quad \text{or} \quad \left(\frac{3 \text{ mole } O_2}{2 \text{ moles } Al_2O_3}\right)$$

These are the two possible conversion factors to use in the solution of the problem.

Mole to Mole conversions

- How many moles of O_2 are produced when 3.34 moles of Al_2O_3 decompose?

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

$$3.34 \text{ mol } Al_2O_3 \left(\frac{3 \text{ mol } O_2}{2 \text{ mol } Al_2O_3} \right) = 5.01 \text{ mol } O_2$$

Conversion factor from balanced equation

If you know the amount of **ANY** chemical in the reaction, you can find the amount of *ALL* the other chemicals!

Practice:

$$2C_2H_2 + 5\ O_2 \rightarrow 4CO_2 + 2\ H_2O$$

- If 3.84 moles of C_2H_2 are burned, how many moles of O_2 are needed?(9.6 mol)

•How many moles of C_2H_2 are needed to produce 8.95 mole of H_2O? (8.95 mol)

•If 2.47 moles of C_2H_2 are burned, how many moles of CO_2 are formed? (4.94 mol)

How do you get good at this?

Practice!!

Steps to Calculate Stoichiometric Problems

1. Correctly balance the equation.
2. Convert the given amount into moles.
3. Set up mole ratios.
4. Use mole ratios to calculate moles of desired chemical.
5. Convert moles back into final unit.

Mass-Mass Problem:

6.50 grams of aluminum reacts with an excess of oxygen. How many grams of aluminum oxide are formed?

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

6.50 g Al	1 mol Al	2 mol Al_2O_3	101.96 g Al_2O_3	= ? g Al_2O_3
	26.98 g Al	4 mol Al	1 mol Al_2O_3	

(6.50 x 1 x 2 x 101.96) ÷ (26.98 x 4 x 1) = **12.3 g Al_2O_3** are formed

Another example:

- If 10.1 g of Fe are added to a solution of Copper (II) Sulfate, how many grams of solid copper would form?

$$2Fe + 3CuSO_4 \rightarrow Fe_2(SO_4)_3 + 3Cu$$

Answer = 17.2 g Cu

Volume-Volume Calculations:

- How many liters of CH_4 at STP are required to completely react with 17.5 L of O_2 ?

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

$$17.5\ \text{L}\ O_2 \left(\frac{1\ \text{mol}\ O_2}{22.4\ \text{L}\ O_2}\right)\left(\frac{1\ \text{mol}\ CH_4}{2\ \text{mol}\ O_2}\right)\left(\frac{22.4\ \text{L}\ CH_4}{1\ \text{mol}\ CH_4}\right)$$

$$= 8.75\ \text{L}\ CH_4$$

Notice anything relating these two steps?

Avogadro told us:

- Equal volumes of gas, at the same temperature and pressure contain the same number of particles.
- Moles are numbers of particles
- You can treat reactions as if they happen <u>liters</u> at a time, as long as you keep the temperature and pressure the same.

1 mole = 22.4 L @ STP

Shortcut for Volume-Volume?

- How many liters of CH_4 at STP are required to completely react with 17.5 L of O_2?

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

$$17.5 \text{ L } O_2 \left(\frac{1 \text{ L } CH_4}{2 \text{ L } O_2} \right) = 8.75 \text{ L } CH_4$$

Note: This only works for Volume-Volume problems.

Section 12.3
Limiting Reagent & Percent Yield

- **OBJECTIVES:**
  - *Identify* the limiting reagent in a reaction.

Section 12.3
Limiting Reagent & Percent Yield

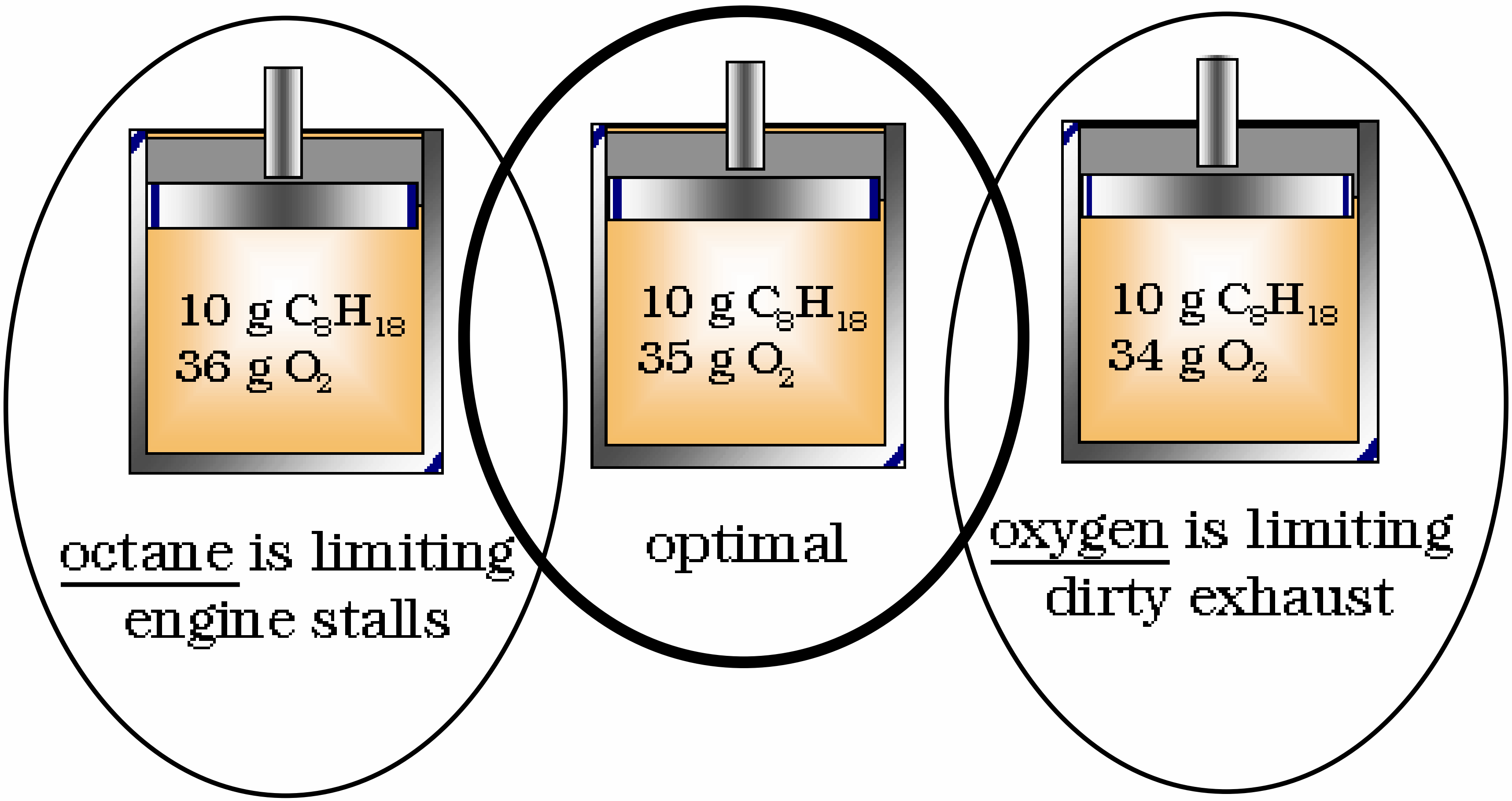
- **OBJECTIVES:**
  - *Calculate* theoretical yield, percent yield, and the amount of excess reagent that remains unreacted given appropriate information.

“Limiting” Reagent

- If you are given one dozen loaves of bread, a gallon of mustard, and three pieces of salami, how many salami sandwiches can you make?
- The *limiting reagent* is the reactant you run out of first.
- The *excess reagent* is the one you have left over.
- The limiting reagent determines how much product you can make

Limiting Reagents

$$2\ C_8H_{18}(l) + 25\ O_2(g) = 16\ CO_2(g) + 18\ H_2O(l)$$



How do you find out which is limited?

- The chemical that makes the *least amount of product* is the "limiting reagent".
- You can recognize limiting reagent problems because they will give you 2 amounts of chemical
- Do *two* stoichiometry problems, one for each reagent you are given.

- If 10.6 g of copp... r... 3.83 g sulfur, ho... he product (coppe... ...med?

2...

$$10.6 \text{ g Cu} \left(\frac{1 \text{ mol} \ldots}{63.55\text{g Cu} \ldots}\right) \ldots \left(\frac{\ldots \text{g } Cu_2S}{\ldots\text{ol } Cu_2S}\right)$$

= 13.3 g Cu_2S

$$3.83 \text{ g S} \left(\frac{1 \text{ mol S}}{32.06\text{g S}}\right)\left(\frac{1 \text{ mol } Cu_2S}{1 \text{ mol S}}\right)\left(\frac{159.16 \text{ g } Cu_2S}{1 \text{ mol } Cu_2S}\right)$$

= 19.0 g Cu_2S

Cu is the Limiting Reagent, since it produced less product.

Another example:

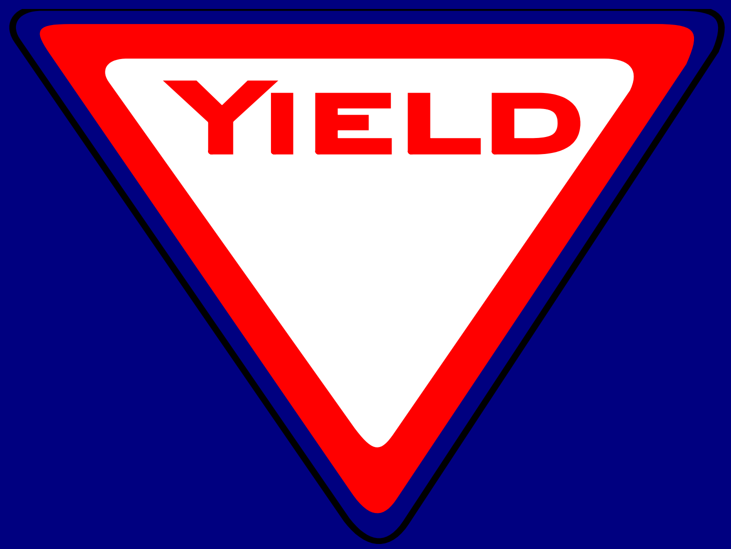
- If 10.3 g of aluminum are reacted with 51.7 g of $CuSO_4$ how much copper (grams) will be produced?

$$2Al + 3CuSO_4 \rightarrow 3Cu + Al_2(SO_4)_3$$

the $CuSO_4$ is limited, so Cu = 20.6 g

- How much excess reagent will remain? Excess = 4.47 grams

The Concept of:



A little different type of yield than you had in Driver’s Education class.

What is Yield?

- Yield is the amount of product made in a chemical reaction.
- There are three types:

1. Actual yield- what you actually get in the lab when the chemicals are mixed
2. Theoretical yield- what the balanced equation tells ***should*** be made
3. Percent yield = $\frac{\text{Actual}}{\text{Theoretical}}$ **x 100**

Example:

- 6.78 g of copper *is produced* when 3.92 g of Al are reacted with excess copper (II) sulfate.

$$2Al + 3\ CuSO_4 \rightarrow Al_2(SO_4)_3 + 3Cu$$

- What is the actual yield? = 6.78 g Cu
- What is the theoretical yield? = 13.8 g Cu
- What is the percent yield? = 49.1 %

Details on Yield

- Percent yield tells us how “efficient” a reaction is.
- Percent yield **can not** be bigger than 100 %.
- Theoretical yield will always be larger than actual yield!
  - **Why?** Due to impure reactants; competing side reactions; loss of product in filtering or transferring between containers; measuring

End of Chapter 12