Chemical Reactions (equations)

Reactants → Products

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

Balanced chemical equations DO tell you:

1. Identity of reactants and products;
2. Amounts (# of moles) of each reactant relative to other reactants and to products; ***(Stoichiometry)***
3. Physical state (phase) of the reactants and products*
   *often, not always

Balanced chemical equations DO NOT tell you:

1. If energy is consumed or released in the process
   Thermodynamics
2. How fast the reaction proceeds (rate)
   Kinetics

1

Balancing Chemical Reactions

A reaction isn't very useful unless it is <u>balanced</u>.

- Balanced reaction: # of atoms of element A must be the same on both the *reactant* and *product* sides of the equation.
- No magic formula for balancing reactions – trial and error process.
  - Only adjust coefficients, not formulas for compounds!
  - Don't introduce species that aren't present in the reaction

<u>Example</u>: Write a balanced chemical equation for the complete combustion of tetraethylead, $Pb(C_2H_5)_4$.

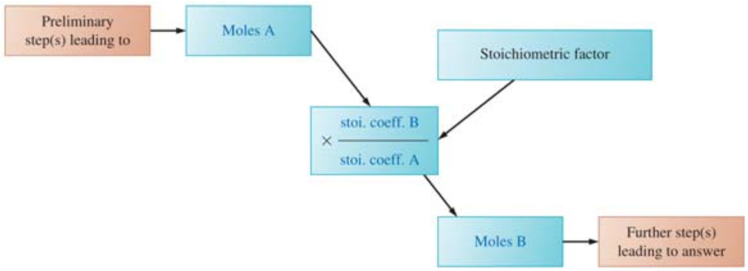
$\_\_Pb(C_2H_5)_4\ (l) + \_\_O_2\ (g) \rightarrow \_\_PbO\ (s) + \_\_H_2O\ (l) + \_\_CO_2(g)$

2

Using Balanced Reactions

Most common: Predicting quantities of material produced or consumed in a reaction.

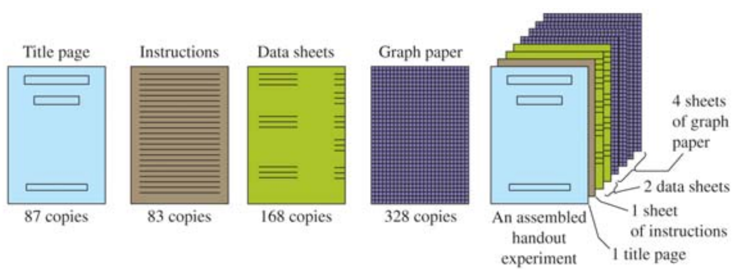
- Generally do mass to mole (or moles to mass) conversion
- *Number of atoms (moles) is most important!!*



Example: What mass of water is produced by the combustion of 12.6 g octane (C_8H_{18})?

Using Balanced Reactions: Practical Considerations

The quantity of products formed in a reaction is determined by the reactant that is completely consumed first – **limiting reagent (reactant)**.



Example: What mass of water will be produced if 12.6 g C_8H_{18} is allowed to react with 12.6 g O_2?

Using Balanced Reactions:
Practical Considerations

In real life, it is very rare to produce 100% of the material you would expect based on reaction stoichiometry – **theoretical yield**

- may be some loss (waste) during the reaction
- some reactions simply don't go to completion!
- quantify by calculating percent yield

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Using Stoichiometry

Example: You have 2.357 g of a mixture of $BaCl_2$ and $BaCl_2 \bullet 2\ H_2O$. If experiment shows that the mixture has a mass of only 2.108 g after heating to drive of all the waters of hydration in $BaCl_2 \bullet 2\ H_2O$, what is the weight percent of $BaCl_2 \bullet 2\ H_2O$ in the original mixture?

Example: A compound contains only C, H, and O. Combustion of 10.68 mg of the compound yields 16.01 mg CO_2 and 4.37 mg H_2O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of the compound?

Using Balanced Reactions:
Considerations of reactions in solution

Many (most) reactions occur when reactants are dissolved in solution. How does this impact our approach?

- We need a measure of the concentration of the reactant (solute) in the solvent.
- Many ways to represent concentration
  - Molarity, molality, %, ppm…

Molarity is most convenient for stoichiometry

$$\text{molarity (M)} = \frac{\text{moles solute}}{\text{volume solution (L)}}$$

Examples:

What is the molarity of NaCl when 12.6 g NaCl is dissolved in 250 mL solution?

How many grams of glucose ($C_6H_{12}O_6$) must be dissolved in water to produce 75.0 mL of 0.350 M glucose?

Using Balanced Reactions:
Considerations of reactions in solution

Dilution: Number of moles is always constant!

$M_1V_1 = M_2V_2$ (or $M_{conc}V_{conc} = M_{dil}V_{dil}$)

Example: How many mL of concentrated sulfuric acid must be diluted to 100 mL to prepare a 1.00 M solution? Concentrated sulfuric acid is 18.0 M.

Example: How many mL of 0.650 M K_2CrO_4 are needed to precipitate all of the silver in 415 mL of 0.186 M $AgNO_3$ as $AgCrO_4$(s)?

$$2\ AgNO_3\ (aq) + K_2CrO_4\ (aq) \rightarrow Ag_2CrO_4(s) + 2\ KNO_3\ (aq)$$