Chemistry 1B
Winter Quarter 2012

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Balancing Chemical Equations

Chemical equations are usually written in a balanced way so as to conserve the total number of all types of atoms, and therefore the total mass.

Unbalanced (incorrect)
\[
\text{CH}_4(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)
\]

1 C, 4 H, 2 O atoms ≠ 1 C, 2 H, 3 O atoms

Balanced (correct)
\[
\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)
\]

1 C, 4 H, 4 O atoms = 1 C, 4 H, 4 O atoms

Sample Problem

Solid ammonium dichromate can be ignited to form nitrogen gas, water vapor, and solid chromium (III) oxide. Write a balanced chemical equation for this reaction.

1. Write down the formulas and physical states for all the reactants and products

Reactants: \((\text{NH}_4)_2\text{Cr}_2\text{O}_7\) (s)
Products: \(\text{N}_2\) (g), H\(_2\)O (g), Cr\(_2\)O\(_3\) (s)

3.7 – Reactions and Equations

Chemical Reaction: a process that changes one set of chemical substances into another set.

Gas flare: methane, oxygen change into water, carbon dioxide

Chemical Equation: uses chemical symbols to represent a chemical reaction.

\[
\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)
\]

Reactants (left side) Products (right side)

Information in Chemical Equations

\[
\text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g)
\]

It does NOT mean: 1 g methane + 2 g oxygen → 1 g carbon dioxide + 2 g water

The reaction equations may also specify the physical states of the reactants and products:

\(g\) = gaseous \(s\) = solid
\(l\) = liquid \(aq\) = aqueous

Sample Problem

1. Write an unbalanced equation with reactants on the left side and products on the right side.

\[
(\text{NH}_4)_2\text{Cr}_2\text{O}_7\) (s) → \text{N}_2\) (g) + H\(_2\)O (g) + Cr\(_2\)O\(_3\) (s)
\]

2. Write a balanced chemical equation for this reaction.

\[
(\text{NH}_4)_2\text{Cr}_2\text{O}_7\) (s) → \text{N}_2\) (g) + 4 \text{H}_2\text{O}(g) + Cr\(_2\)O\(_3\) (s)
\]

3. Balance the equation, starting with the most complicated molecule.

\[
\begin{align*}
(\text{NH}_4)_2\text{Cr}_2\text{O}_7\) (s) & \rightarrow \text{N}_2\) (g) + 4 \text{H}_2\text{O}(g) + Cr\(_2\)O\(_3\) (s) \\
\end{align*}
\]
Sample Problems (Board)

Aqueous solutions of sodium sulfate and calcium nitrate are mixed, forming a calcium sulfate precipitate. Write a balanced chemical equation for this precipitation reaction.

Aqueous perchloric acid reacts with solid magnesium hydroxide to form water and aqueous magnesium perchlorate. Write a balanced chemical equation for this acid-base reaction.

The Hindenburg exploded in 1927 during a landing attempt. Hydrogen gas reacted with oxygen in the atmosphere to produce water vapor. Write a balanced chemical equation for this combustion reaction.

Stoichiometry is the study of the quantitative relationships between the reactants and products in chemical equations.

3.8 – Stoichiometry

From Greek:
Stoikheion = element metria = measure

Stoichiometric Calculations

Reaction:  \(A \rightarrow B\)

- Write a balanced chemical equation
- Convert the given masses of reactants into moles
- Use coefficients in balanced equation to calculate the number of moles of the product of interest
- Convert molar amounts of products into their masses
- Pay attention to the number of significant digits

Sample Problem

For the ammonium chromate volcano reaction, calculate the masses of products formed from 10.8 g of \((\text{NH}_4)_2\text{Cr}_2\text{O}_7\).

1. Balance the equation for the reaction.
 \[(\text{NH}_4)_2\text{Cr}_2\text{O}_7 (s) \rightarrow \text{N}_2 (g) + 4 \text{H}_2\text{O} (g) + \text{Cr}_2\text{O}_3 (s)\]

2. Convert the known mass of the reactant or product to moles of that material.

\[
\frac{10.8 \text{ g (NH}_4)_2\text{Cr}_2\text{O}_7}{252.1 \text{ g (NH}_4)_2\text{Cr}_2\text{O}_7} = 0.0428 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7
\]

3. Using the balanced equation, get the correct mole ratios.

1 mole of \((\text{NH}_4)_2\text{Cr}_2\text{O}_7 (s)\) makes:

- 1 mole of \(\text{N}_2 (g)\)
- 4 moles of \(\text{H}_2\text{O} (g)\)
- 1 mole of \(\text{Cr}_2\text{O}_3 (s)\)

4. Use these molar ratios to calculate the number of moles of the reactant or product of interest.

\[
\begin{align*}
0.0428 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7 & \times \frac{1 \text{ mol N}_2}{1 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7} = 0.0428 \text{ mol N}_2 \\
0.0428 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7 & \times \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7} = 0.171 \text{ mol H}_2\text{O} \\
0.0428 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7 & \times \frac{1 \text{ mol Cr}_2\text{O}_3}{1 \text{ mol (NH}_4)_2\text{Cr}_2\text{O}_7} = 0.0428 \text{ mol Cr}_2\text{O}_3
\end{align*}
\]

5. Convert from molar amounts back into grams.

\[
\begin{align*}
0.0428 \text{ mol N}_2 & \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 1.20 \text{ g N}_2 \\
0.171 \text{ mol H}_2\text{O} & \times \frac{18.01 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 3.08 \text{ g H}_2\text{O} \\
0.0428 \text{ mol Cr}_2\text{O}_3 & \times \frac{152.0 \text{ g Cr}_2\text{O}_3}{1 \text{ mol Cr}_2\text{O}_3} = 6.51 \text{ g Cr}_2\text{O}_3
\end{align*}
\]

6. Report the answers the same number of significant digits as in the initial values (3 digits in this case).
Sample Problem

Methanol burns in air according to the equation

\[ 2 \text{CH}_3\text{OH} + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 4 \text{H}_2\text{O} \]

If 209 g of methanol are used up in the combustion, what mass of water is produced?

\[
\begin{align*}
\text{grams CH}_3\text{OH} & \quad \text{moles CH}_3\text{OH} & \quad \text{moles H}_2\text{O} & \quad \text{grams H}_2\text{O} \\
209 \text{g CH}_3\text{OH} & \times 1 \text{ mol CH}_3\text{OH} & \times \frac{4 \text{ mol H}_2\text{O}}{2 \text{ mol CH}_3\text{OH}} & \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \\
& = 235 \text{ g H}_2\text{O}
\end{align*}
\]

3.9 – Limiting and Excess Reactants

\[ \text{H}_2 (\text{g}) + \text{Cl}_2 (\text{g}) \rightarrow 2 \text{HCl (g)} \]

Before the reaction:
5 molecules of \( \text{H}_2 \)
2 molecules of \( \text{Cl}_2 \)

Reaction goes to completion:
3 unused molecules of \( \text{H}_2 \)
2 molecules of \( \text{HCl} \)

\( \text{Cl}_2 \) is the\textbf{ limiting reactant}
\( \text{H}_2 \) is the\textbf{ excess reactant}

3.10 – Reaction Yield

Very few reactions actually go to completion (some reagents, both limiting and excess, are always left)

\textbf{Theoretical Yield} is the amount of product that would result if all the limiting reagent reacted.
\textbf{Actual Yield} is the amount of product actually obtained from a reaction.

\[
\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100
\]

Sample Problem

In one process, 124 g of \( \text{Al} \) are reacted with 601 g of \( \text{Fe}_2\text{O}_3 \)

\[ 2 \text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2 \text{Fe} \]

Calculate the maximal mass of \( \text{Al}_2\text{O}_3 \) that can be formed. If 200 g of \( \text{Al}_2\text{O}_3 \) is actually formed what is the reaction yield?

1. Figure out which reactant is limiting by calculating their molar amounts.

\[
\begin{align*}
n_{\text{Al}} &= \frac{m_{\text{Al}}}{MW_{\text{Al}}} = \frac{124 \text{ g}}{26.98 \text{ g/mol}} = 4.60 \text{ mol} \\
n_{\text{Fe}_2\text{O}_3} &= \frac{m_{\text{Fe}_2\text{O}_3}}{MW_{\text{Fe}_2\text{O}_3}} = \frac{601 \text{ g}}{159.7 \text{ g/mol}} = 3.76 \text{ mol}
\end{align*}
\]

To have all of \( \text{Fe}_2\text{O}_3 \) react we need \( 2 \times 3.76 \text{ mol} = 7.52 \text{ mol Al} \)
We only have 4.60 mol Al, so \textbf{Al is the limiting reactant}

2. Calculate the maximal amount of product (\( \text{Al}_2\text{O}_3 \)) that can be produced from the limiting reactant (\( \text{Al} \))

\[ 2 \text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2 \text{Fe} \]

2 molecules of \( \text{Al} \) → 1 molecule of \( \text{Al}_2\text{O}_3 \)

4.60 moles of \( \text{Al} \) → 2.30 mole of \( \text{Al}_2\text{O}_3 \)

\[
m_{\text{Al}_2\text{O}_3} = n_{\text{Al}_2\text{O}_3} \times MW_{\text{Al}_2\text{O}_3} = 2.30 \text{ mol} \times 101.96 \frac{g}{\text{mol}} = 234.5 \text{ g}
\]

3. Calculate the percent yield from the actual yield of the product (200 g of \( \text{Al}_2\text{O}_3 \) actually produced)

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{maximum yield}} \times 100\% = \frac{200 \text{ g}}{234.5 \text{ g}} \times 100\% = 85.2\%
\]

2. Calculate the maximal amount of product (\( \text{Al}_2\text{O}_3 \)) that can be produced from the limiting reactant (\( \text{Al} \))

\[ 2 \text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Al}_2\text{O}_3 + 2 \text{Fe} \]

2 molecules of \( \text{Al} \) → 1 molecule of \( \text{Al}_2\text{O}_3 \)

4.60 moles of \( \text{Al} \) → 2.30 mole of \( \text{Al}_2\text{O}_3 \)

\[
m_{\text{Al}_2\text{O}_3} = n_{\text{Al}_2\text{O}_3} \times MW_{\text{Al}_2\text{O}_3} = 2.30 \text{ mol} \times 101.96 \frac{g}{\text{mol}} = 234.5 \text{ g}
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\]

% Yield = \[
\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%
\]

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{maximum yield}} \times 100\% = \frac{200 \text{ g}}{234.5 \text{ g}} \times 100\% = 85.2\%
\]