Balancing Equations

\[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \]

reactants \hspace{2cm} \text{products}

+ "reacts with" \hspace{2cm} \rightarrow "to produce"

coefficients - indicate amount of substance

Equations must be balanced. Equal amounts of each element on each side of the equation.

\[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \]

\[ 4 \text{H} \hspace{2cm} 2\text{O} \hspace{2cm} 4\text{H}, 2\text{O} \]

When balancing an equation, subscripts should never change, as this changes the chemical identity.

Changing coefficients only changes amount of a substance, not the identity.

\[ \text{H}_2\text{O} \]

\[ 2 \text{H}_2\text{O} \]

\[ \text{H}_2\text{O}_2 \]

\[ \text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \ \text{ (unbalanced)} \]
CH\textsubscript{4} + O\textsubscript{2} → CO\textsubscript{2} + H\textsubscript{2}O

Guidelines:
1. write unbalanced equation
2. use coefficients to indicate how many formula units are required to balance equation
3. balance those species that occur in the fewest formulas on each side.
4. reduce coefficients to smallest whole number values
5. when balancing reactions involving organic compounds, balance in the order: C, H, O

Chemical Symbols on Different Levels

Chemical symbols represent both a microscopic and macroscopic level.

\[ 2 \text{H}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{H}_2\text{O}_\text{(g)} \]

2 molecules \hspace{1cm} 1 molecule \hspace{1cm} 2 molecules

2(6x10\textsuperscript{23} \text{molecules}) \hspace{1cm} 6x10\textsuperscript{23} \text{molecules} \hspace{1cm} 2(6x10\textsuperscript{23} \text{molecules})

2 moles \hspace{1cm} 1 mole \hspace{1cm} 2 moles

The coefficients in a balanced chemical equation can be interpreted both as the relative numbers of molecules (or formula units) involved in the reaction and as the relative numbers of moles.

Avogadro’s Number and the Mole

Atomic and Molecular Weights

The Atomic Mass Scale

1 amu = 1.66054x10\textsuperscript{-24} g and 1 g = 6.02214x10\textsuperscript{23} amu

\textsuperscript{1}H \hspace{1cm} 1.6735x10\textsuperscript{-24} g \hspace{1cm} 1.0078 \text{amu}
Average Atomic Masses

Determine average atomic mass by using masses of various isotopes and their relative abundances.

Carbon is 98.892% $^{12}\text{C}$ and 1.108% $^{13}\text{C}$

$^{12}\text{C}$ is 12 amu (exactly) and $^{13}\text{C}$ is 13.00335 amu

$\text{(0.98892)(12 amu) + (0.01108)(13.00335 amu)} = 12.011 \text{ amu}$

The average atomic mass of each element (expressed in amu) is also known as its atomic weight.

Formula and Molecular Weights (Masses)

The formula weight of a substance is merely the sum of the atomic weights of each atom in its chemical formula.

ex: $\text{H}_2\text{SO}_4$ has a formula weight of 98.1 amu

$\text{FW} = 2(\text{AW of H}) + (\text{AW of S}) + 4(\text{AW of O})$

If the chemical formula is the molecular formula, then the formula weight is also called the molecular weight.

ex: glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, has a molecular weight of 180.0 amu

$\text{MW} =$

With ionic substances such as NaCl, it is inappropriate to speak of molecules. We will use the formula weight

$\text{FW} =$
A mole is defined as the amount of matter that contains as many objects (atom, molecules, or whatever objects we are considering) as the number of atoms in exactly 12 g of $^{12}$C. That number is $6.0221421 \times 10^{23}$.

This number is given a special name: Avogadro's number, $N_A$.

We will use $6.022 \times 10^{23}$.

**Molar Mass**
The mass of single atom of an element (in amu) is numerically equal to the mass (in grams) of 1 mol of atoms of that element.

One $^{12}$C atom weights 12 amu $\rightarrow$ 1 mol of $^{12}$C weighs 12 g.

One $^{24}$Mg atom weights 24 amu $\rightarrow$ 1 mol of $^{24}$Mg weighs 24 g.

One $^{197}$Au atom weights 197 amu $\rightarrow$ 1 mol of $^{197}$Au weighs 197 g.

The mass in grams of 1 mol of a substance is called its molar mass.

The molar mass (in grams) of any substance is always numerically equal to its formula weight (in amu):

The molecular mass (also the molecular weight, MW) is the sum of all the atoms in a molecule.

The formula mass is the sum of all the atoms in the formula unit of any compound, molecular or ionic.

One H$_2$O molecule weighs 18.0 amu $\rightarrow$ 1 mol of H$_2$O weighs 18.0 g.

One NO$_3^-$ ion weighs 62.0 amu $\rightarrow$ 1 mol of NO$_3^-$ weighs 62.0 g.

One NaCl unit weighs 58.5 amu $\rightarrow$ 1 mol of NaCl weighs 58.5 g.
**Stoichiometry: Chemical Arithmetic**

Interconverting Masses, Moles, and Numbers of Particles

Conversion of mass to moles and moles to mass are made easy by use of dimensional analysis.

Example: How many moles of trinitrotoluene, C\(_7\)H\(_5\)N\(_3\)O\(_6\) (TNT) are there in 214.0 g of TNT?

Solution: The molar mass can be used as a conversion factor for converting grams to moles. The molar mass of TNT is 227.0 g.

Example: What is the mass of 3.125 mol of CO\(_2\)?

Solution: The molar mass of CO\(_2\) is 44.0 g. 1 mol of CO\(_2\)

Conversion factors can also be used to find the number of atoms, molecules, or ions.

Example: Find the number of copper atoms in a copper penny, which weighs 3 g.
for the balanced equation:
\[ a \, A + b \, B \rightarrow c \, C + d \, D \]

<table>
<thead>
<tr>
<th>Grams of A</th>
<th>Moles of A</th>
<th>Moles of B</th>
<th>Grams of B</th>
</tr>
</thead>
</table>

Use molar mass of A as a conversion factor
Use coefficients in the balanced equations to find mole ratios
Use molar mass of B as a conversion factor

\[ 2 \, H_2 + O_2 \rightarrow 2 \, H_2O \]

2 mol H\(_2\) = 1 mol O\(_2\) = 2 mol H\(_2\)O

They are stoichiometrically equivalent.

How much H\(_2\)O can be produced from 1.57 mol O\(_2\)?

Combustion of butane, C\(_4\)H\(_{10}\).

If we burn 1.00 g of butane, what mass of CO\(_2\) is produced?
Yields of Chemical Reactions

The quantity or product that is calculated to form when all of the limiting reactant reacts is called the **theoretical yield**.

The amount actually obtained in a reaction is called the **actual yield**.

The percent yield of a reaction relates the actual yield to the theoretical (calculated) yield:

\[
\text{Percent Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%
\]

**Example:** 4.92 g of Ba\(_3\)(PO\(_4\))\(_2\) should form when 3.50 g of Na\(_3\)PO\(_4\) is mixed with 6.40 g of Ba\(_3\)(NO\(_3\))\(_2\). This is the theoretical yield of the reaction. If the actual yield turned out to be 4.70 g, the percent yield would be?

Reactions with Limiting Amounts of Reactants

The reagent that is completely consumed in a reaction is called the **limiting reactant** or **limiting reagent**, because it determines, or limits, the amount of product formed.

The other reactants are sometimes called **excess reactants** or **excess reagents**.

Assume that 5.88 g of B\(_2\)S\(_3\) is mixed with 7.85 g of water. Which reactant is limiting and which reactant is in excess? How many grams of the excess reactant are consumed? How many grams of boric acid are produced?

\[\text{B}_2\text{S}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{BO}_3 + \text{H}_2\text{S}\] (Always make sure the chemical equation is balanced!)
Percent Composition and Empirical Formulas

Percentage Composition from Formulas

Percent composition- the percentage by mass contributed by each element in the substance.

\[ \text{C}_6\text{H}_{12}\text{O}_6 \]

Determining Empirical Formulas: Elemental analysis

Analysis gives the amount of each element as a percentage. If we assume the sample to be 100 g, we can divide these masses (the percentages in grams) by the appropriate atomic weight to obtain the number of moles of each element in 100 g.

We then divide the larger mole numbers by the smallest mole number to obtain the empirical formula. The ratios may not be exact due to experimental errors.
Combustion Analysis

When compounds that contain carbon and hydrogen are combusted, all of the carbon of the compound is converted to CO$_2$ and all the hydrogen to H$_2$O. The masses of CO$_2$ and H$_2$O can be used to determine the number of moles of C and H in the original compound and thereby the empirical formula.