Chemical Equations

Chemical Reaction: Interaction between substances that results in one or more new substances being produced

Example: hydrogen + oxygen → water

Reactants of a Reaction: Starting materials that undergo chemical change; written on the left side of the equation representing the reaction

Products of a Reaction: Substances formed as a result of the reaction; written on the right side of the equation representing the reaction

The arrow points towards the products formed by the reaction

Individual products and reactants are separated by a plus sign

Chemical Equation: A written statement using symbols and formulas to describe the changes that occur in a reaction

Example: 2H₂(g) + O₂(g) → 2H₂O(l)

Letter in parentheses indicates the state of the substance: gas (g), liquid (l), solid (s), dissolved in water (aq)

If heat is required for the reaction to take place, the symbol Δ is written over the reaction arrow

Balanced Equation: Equation in which the number of atoms of each element in the reactants is the same as the number of atoms of that element in the products

Law of Conservation of Mass: Atoms are neither created nor destroyed in chemical reactions

Example: CaS + H₂O → CaO + H₂S

Reactants Products

Is this equation balanced?
Example: \( \text{NO(g)} + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \)

Is this equation balanced?

Adjust the coefficient of the reactants and products to balance the equation

\[
2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2
\]

Reactants \hspace{1cm} \text{Product}

Practice: \( \text{SO}_2 + \text{O}_2 \rightarrow \text{SO}_3 \)

Is this equation balanced?

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

Practice: \( \text{H}_2 + \text{Cl}_2 \rightarrow \text{HCl} \)

Is this equation balanced?

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

Avogadro’s Number: The Mole

1 Dozen = 12
1 Gross = 144
1 Mole = \(6.02 \times 10^{23}\)

1 mole of atoms = \(6.022 \times 10^{23}\) atoms
1 mole of molecules = \(6.022 \times 10^{23}\) molecules
The Mole and Avogadro’s Number

**Example:** How many molecules are contained in 5.0 moles of carbon dioxide (CO₂)?

**Step [1]** Identify the original quantity and the desired quantity.

\[ 5.0 \text{ mol of CO}_2 \times \text{ conversion factor} = ? \text{ molecules of CO}_2 \]

**Step [2]** Write out the conversion factors.

\[ \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ molecules}} \text{ or } \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} \]

Choose this one to cancel mol.

**Step [3]** Set up and solve the problem.

\[ 5.0 \text{ mol} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.0 \times 10^{24} \text{ molecules CO}_2 \]

Unwanted unit cancels.

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**Molar Mass:** the mass (in grams) of one mole a particular substance

There is a unique relationship between molar mass and atomic weight:

Molar mass (in grams) **is always** equal to the atomic weight of the atom!

**Examples:**
- Atomic weight of carbon is 12.01 amu
  - Molar mass of carbon is 12.01 g/mol
- Atomic weight of helium is 4.00 amu
  - Molar mass of helium is 4.00 g/mol
Mole (mol): The number of particles (atoms or molecules) in a sample of element or compound with a mass in grams equal to the atomic (or molecular) weight

Example: Atomic weight of sodium = 22.99

22.99 g of sodium contains 1 mole
(6.02 x 10^23) atoms in 22.99 g of sodium

Mass to Mole Conversions: Relating Grams to Moles

Because molar mass relates the number of moles to the number of grams of a substance, molar mass can be used as a conversion factor.

Example: How many moles are present in 100. g of aspirin (C9H8O4, molar mass 180.2 g/mol)?

Step [1] Identify the original quantity and the desired quantity.

100. g of aspirin original quantity x conversion factor = ? mol of aspirin desired quantity


The conversion factor is the molar mass, and it can be written in two ways.

\[
\frac{180.2 \text{ g aspirin}}{1 \text{ mol}} \quad \text{or} \quad \frac{1 \text{ mol}}{180.2 \text{ g aspirin}}
\]

Choose this one to cancel g aspirin.
Step 3: Set up and solve the problem.

\[
\text{100 g aspirin} \times \frac{1 \text{ mol}}{180.2 \text{ g aspirin}} = 0.555 \text{ mol aspirin}
\]

Unwanted unit cancels.

Calculations Using the Mole

1 mole represents the mass of a sample that contains Avogadro’s number of particles

Example: Atomic wt. of Potassium (K) = 39 amu
1 mol K atoms = 6.02 x 10^{23} atoms = 39 g K

Example: Atomic wt. of Sulfur (S) = 32 u
1 mol S atoms = 6.02 x 10^{23} atoms = 32 g S
\[
\begin{align*}
1 \text{ mol S atoms} & = 6.02 \times 10^{23} \text{ atoms} \\
& = 32 \text{ g S} \\
& = 6.02 \times 10^{23} \text{ atoms} \\
& = 32 \text{ g S} \\
& = 1 \text{ mol S atoms} \\
& = 32 \text{ g S}
\end{align*}
\]

Practice:

1 mol S atoms = 6.02 x 10^{23} atoms
6.02 x 10^{23} atoms = 32 g S
1 mol S atoms = 32 g S

What is the mass in grams of 1 atom of Sulfur?

1 S atom = g
1 S atom x 32 g S / 6.02 x 10^{23} atoms = g
1 S atom = 5.32 x 10^{-23} g
Practice:
1 mol S atoms = 6.02 x 10²³ atoms
6.02 x 10²³ atoms = 32 g S
1 mol S atoms = 32 g S

How many moles of Sulfur in 98.6 grams?
98.6g = moles Sulfur
98.6g x 1 mole/32g = moles sulfur
= 3.07 moles sulfur

Practice:
1 mol S atoms = 6.02 x 10²³ atoms
6.02 x 10²³ atoms = 32 g S
1 mol S atoms = 32 g S

How many atoms in this sample of 98.6g of S?
98.6g = atoms of S
98.6g x 6.02 x 10²³ atoms/32g = atoms of S
= 1.85 x 10²⁴ atoms S

Practice:
1 mol S atoms = 6.02 x 10²³ atoms
6.02 x 10²³ atoms = 32 g S
1 mol S atoms = 32 g S

What is the mass of 1 atom of Sulfur?
1 atom = g
1 atom x 32g/6.02 x 10²³ atoms = g
= 5.33 x 10⁻²³ g S
The mole concept applies to molecules, as well as to atoms.

Chemical formulas indicate relative quantities of atoms within a compound.

Example: $\text{H}_2\text{O}$ has 2 H atoms for every 1 Oxygen atom.

$\text{C}_4\text{H}_12\text{O}_6$ has how many atoms of each element?

**Formula Weight (F.W.):** sum of all atomic weights of all atoms in a compound; expressed in amu.

A mole of a molecule will have a mass in grams equal to its formula weight.

Example: F.W. of $\text{CH}_4\text{N}_2\text{O}$ (urea) = ?

Atomic weights:
- $\text{C} = 12$ amu
- $\text{H} = 1$ amu
- $\text{N} = 14$ amu
- $\text{O} = 16$ amu

F.W. $\text{CH}_4\text{N}_2\text{O} = 12 + 4 (1) + 2 (14) + 16 = 60$ amu

Molar mass of $\text{CH}_4\text{N}_2\text{O}$ is 60 g/mol

**Practice:** Formula weight of $\text{H}_2\text{O}$ = ?

F.W. of $\text{H}_2\text{O} = \text{atomic weight of H atoms + atomic weight of O atom}$

Atomic weight of H =1 amu, O =16 amu

F.W. of $\text{H}_2\text{O} = 18$ amu

Molar mass of $\text{H}_2\text{O} = 18$ g/mol
More Practice:

Prozac, C₁₇H₁₈F₃NO, is a widely used antidepressant that inhibits the uptake of serotonin by the brain. What is the molar mass of Prozac?

17C (12.0) + 18H (1.0) + 3F (19.0) + 1N (14.0) + 1 O (16.0) =
204 + 18 + 57.0 + 14.0 + 16.0 = 309 g/mole

Chemical Equations and the Mole

Stoichiometry: Study of mass relationships in chemical reactions; ratios of different molecules

Example: CH₄ + 2O₂ → CO₂ + 2H₂O

10 CH₄ + 20 O₂ → 10 CO₂ + 20 H₂O
6.02 x 10²³ CH₄ + 12.0 x 10²³ O₂ → 6.02 x 10²³ CO₂ + 12.0 x 10²³ H₂O
1 mol CH₄ + 2 mol O₂ → 1 mol CO₂ + 2 mol H₂O
16.0g CH₄ + 64.0g O₂ → 44.0g CO₂ + 36.0g H₂O

How many moles of O₂ would be required to react with 1.72 mol CH₄?

1.72 mol CH₄ = mol O₂

1.72 mol CH₄ x 2 mol O₂/1 mol CH₄ = 3.44 mol O₂
How many grams of H₂O will be produced from 1.09 mol of CH₄?

2-part problem: first find moles of H₂O then convert moles to grams

Use this equation to obtain conversion factor for moles of CH₄ to moles of H₂O:

\[ \text{1 mol CH}_4 + 2 \text{ mol O}_2 \rightarrow \text{1 mol CO}_2 + 2 \text{ mol H}_2\text{O} \]

1.09 mol CH₄ = \text{mol H}_2\text{O}

1.09 mol CH₄ x 2 mol H₂O/1 mol CH₄ = 2.18 mol H₂O

2.18 mol H₂O = \text{grams H}_2\text{O}

2.18 mol H₂O x 18.0 g H₂O/1 mol H₂O = 39.2 grams H₂O

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Theoretical and Percent Yield

**Theoretical yield:** the maximum amount of product that would be formed from a particular reaction in an ideal world

**Actual yield:** the amount of product formed from a particular reaction in the real world (usually less than the theoretical yield)

**Percent yield:** ratio of actual yield to theoretical yield, times 100%

\% yield = actual yield/theoretical yield x 100%

**Example:**

theoretical yield = 39.2g water

actual yield = 35.5 g water

\% yield = 35.5g/39.2g x 100\% = 90.6 \% yield

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Oxidation and Reduction Reactions

- An oxidation-reduction reaction involves the transfer of electrons from one reactant to another.
- In oxidation, electrons are lost
  
  \[ \text{Zn} \rightarrow \text{Zn}^{2+} + 2e^- \text{ (loss of electrons--LEO)} \]
- In reduction, electrons are gained.
  
  \[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \text{ (gain of electrons--GER)} \]
Identify each of the following as an oxidation or a reduction reaction:

A. \( \text{Sn} \rightarrow \text{Sn}^{4+} + 4\text{e}^- \) Oxidation

B. \( \text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+} \) Reduction

C. \( \text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^- \) Reduction

**Practice:**

Identify each of the following as an oxidation or a reduction reaction:

A. \( \text{Sn} \rightarrow \text{Sn}^{4+} + 4\text{e}^- \) Oxidation

B. \( \text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+} \) Reduction

C. \( \text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^- \) Reduction

**Zn**<sup>2+</sup> + **Cu**<sup>2+</sup> → **Zn**<sup>2+</sup> + **Cu**

- **Zn** acts as a reducing agent because it causes **Cu**<sup>2+</sup> to gain electrons and become reduced.
- **Cu**<sup>2+</sup> acts as an oxidizing agent because it causes **Zn** to lose electrons and become oxidized.
Examples of Everyday Oxidation–Reduction Reactions:

Iron Rusting

\[ 4 \text{Fe}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Fe}_2\text{O}_3(s) \]

Fe loses e\(^{-}\) and is oxidized.

O gains e\(^{-}\) and is reduced.

Inside an Alkaline Battery

\[ \text{Zn} + 2 \text{MnO}_2 \rightarrow \text{ZnO} + \text{Mn}_2\text{O}_3 \]

Zn loses e\(^{-}\) and is oxidized.

Mn\(^{4+}\) gains e\(^{-}\) and is reduced.