Stoichiometric Relationships

Ms. Peace
Lesson 1

Introduction to Stoichiometry
We Are Here

Quantitative Chemistry

- Mole Ratios, Limiting Reactants and Theoretical Yields
- Particle Theory
- Chemical Equations
- Formulas
- Basic Concepts

The Mole
- Molar Mass

Gases
- Avogadro's Law
- Ideal Gases

Solutions
- Titrations
States of Matter

Matter

- Made up of particles, atoms, molecules, ions
- Occupies a volume in space
- Particles are in constant motion
- Has a mass
States of Matter

Solid → Liquid → Gas

ADD ENERGY

Solid

Liquid

Gas

REMOVE ENERGY
## Properties of States of Matter

<table>
<thead>
<tr>
<th></th>
<th>SOLID</th>
<th>LIQUID</th>
<th>GAS</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Distance</strong></td>
<td>Close together</td>
<td>Close but further apart than in solids</td>
<td>Particles far apart</td>
</tr>
<tr>
<td><strong>Arrangement</strong></td>
<td>Regular</td>
<td>Random</td>
<td>Random</td>
</tr>
<tr>
<td><strong>Shape</strong></td>
<td>Fixed shape</td>
<td>No fixed shape—take the shape of the container</td>
<td>No fixed shape—fill the container</td>
</tr>
<tr>
<td><strong>Volume</strong></td>
<td>Fixed</td>
<td>Fixed</td>
<td>Not Fixed</td>
</tr>
<tr>
<td><strong>Movement</strong></td>
<td>Vibrate</td>
<td>Move around each other</td>
<td>Move around in all directions</td>
</tr>
<tr>
<td><strong>Speed</strong></td>
<td>Slowest</td>
<td>Faster</td>
<td>Fastest</td>
</tr>
<tr>
<td><strong>Energy</strong></td>
<td>Lowest</td>
<td>Higher</td>
<td>Highest</td>
</tr>
<tr>
<td><strong>Forces of attraction</strong></td>
<td>Strongest</td>
<td>Weaker</td>
<td>Weakest</td>
</tr>
</tbody>
</table>
Temperature

- **Units:**
  - Fahrenheit
  - Celsius
  - Kelvin*

- SI units are a set of standard units that are used in science throughout the world

- Absolute zero: 0 K or -273°C
  - \( K = °C + 273 \)
Changes of States

- Endothermic
- Exothermic
- Sublimation
- Condensation
- Vaporization
- Melting (fusion)
- Freezing
- Deposition

Energy of System vs. Disorder of System

Solid ↔ Liquid ↔ Gas
Changes of States

- **Endothermic:**
  - Melting and boiling
  - Energy must be transferred from the surroundings to bring about these changes
  - The potential energy of molecules increases; they vibrate more and move faster

- **Exothermic:**
  - Condensation and freezing
  - Energy is transferred to the surroundings
  - The potential energy of molecules decreases; the vibrate less and move slower
Utilization

- Freeze-drying is a food preservation technique
- Freeze-drying uses the process of sublimation
- Foods that require dehydration are first frozen and then subjected to a reduced pressure
- The frozen water then sublimes directly to water vapor thus dehydrating the food

- How Refrigeration Works
Elements and Compounds

- Elements contain atoms of only one type
- Atoms of elements combine in a fixed ratio to form compounds composed of molecules or ions
- These are the fundamental basis in formulas and chemical equations
- Properties of compounds are very different from those of its constitutional elements
Mixtures

- A pure substance is matter that has a constant composition
  - $\text{N}_2$
  - $\text{H}_2\text{O}$
  - $\text{NaCl}$
  - $\text{C}_6\text{H}_{12}\text{O}_6$
- Pure substances combine physically to form a mixture
- Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.
Mixtures

Filtration

Fractional Distillation
Mixtures

Column Chromatography

Hydrophobic Interaction (hydrophobicity)

Normal-phase (polarity)

Ion-exchange (net charge)

Affinity (specific binding)

Size-exclusion (molecular size)
Types of Mixtures

- **Homogeneous**
  - One phase...you can’t see the separation between the parts of the mixture
  - For example: seawater is a mixture of salt and water, but you can’t see the salt.

- **Heterogeneous**
  - Multiple phases....you can see the different components of the mixture
  - For example: salad dressing is a mixture of oil and vinegar, and you can see the bits of oil and the bits of vinegar.
MYP Review

- State Symbols:
  - (s)
  - (l)
  - (g)
  - (aq)
Types of Chemical Reactions

Four Basic Types of Chemical Reaction

1. **Synthesis**
   
   ![Synthesis Diagram](image)
   
   **Description**: Elements are joined together.
   
   **Example**: \( 2H_2 + O_2 \rightarrow 2H_2O \)

2. **Decomposition**
   
   ![Decomposition Diagram](image)
   
   **Description**: A compound breaks into parts.
   
   **Example**: \( 2H_2O \rightarrow 2H_2 + O_2 \)

3. **Single Replacement**
   
   ![Single Replacement Diagram](image)
   
   **Description**: A single element replaces an element in a compound.
   
   **Example**: \( Zn + 2HCl \rightarrow H_2 + ZnCl_2 \)

4. **Double Replacement**
   
   ![Double Replacement Diagram](image)
   
   **Description**: An element from each of two compounds switch places.
   
   **Example**: \( H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O \)

<table>
<thead>
<tr>
<th>Types</th>
<th>Description</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Synthesis Reactions</td>
<td>Elements are joined together.</td>
<td>( 2H_2 + O_2 \rightarrow 2H_2O )</td>
</tr>
<tr>
<td>Decomposition Reactions</td>
<td>A compound breaks into parts.</td>
<td>( 2H_2O \rightarrow 2H_2 + O_2 )</td>
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Naming Compounds

Ionic or Covalent?

Yes

<table>
<thead>
<tr>
<th>IONIC</th>
<th>COVALENT</th>
</tr>
</thead>
<tbody>
<tr>
<td>Positive ion first</td>
<td>Is there a metal or NH₄⁺?</td>
</tr>
<tr>
<td>What is positive ion?</td>
<td>No</td>
</tr>
<tr>
<td>NH₄⁺ = Ammonium</td>
<td>COVALENT</td>
</tr>
<tr>
<td>Type I Type II</td>
<td>Is it an acid? Does it start with H?</td>
</tr>
<tr>
<td>Use roman numeral = ion charge</td>
<td>Yes</td>
</tr>
<tr>
<td>Use name of atom</td>
<td>ACID</td>
</tr>
<tr>
<td>What is negative ion?</td>
<td>No</td>
</tr>
<tr>
<td>MONATOMIC</td>
<td>NOT ACID</td>
</tr>
<tr>
<td>Add -ide to stem of atom name</td>
<td>USE PREFIXES for the number of each type of atom (mono, di, tri, etc.)</td>
</tr>
<tr>
<td>POLYATOMIC</td>
<td></td>
</tr>
<tr>
<td>Use ion name:</td>
<td>1. Use name of first atom with prefix (except mono)</td>
</tr>
<tr>
<td>_______ate</td>
<td>2. Add -ide to stem of name of second atom with prefix.</td>
</tr>
<tr>
<td>_______ite</td>
<td></td>
</tr>
<tr>
<td>One less O than -ate</td>
<td></td>
</tr>
<tr>
<td>per____ate</td>
<td></td>
</tr>
<tr>
<td>One more O than -ate</td>
<td></td>
</tr>
<tr>
<td>hypo____ite</td>
<td></td>
</tr>
<tr>
<td>One less O than -ite</td>
<td></td>
</tr>
</tbody>
</table>
Word Equations

- Hydrogen + Oxygen → Water

\[
\text{REACTANTS} \quad \text{PRODUCTS}
\]

- \( \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} \)

Is this balanced???
Symbol Equations

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

- The red numbers are called coefficients and tell you the number of each molecule involved in the reaction.
- Required to balance the equation.
- Without them the equation does not balance – each side of the reaction would have different numbers of each atom – which would break physics.
- You can’t change the subscript numbers in the formulas as this changes the chemical.
- If there is no coefficient, it is ‘1’.
Tips for Balancing Equations

1) Balance metals
2) Balance nonmetals
3) Balance oxygen
4) Balance hydrogen

- Try keeping a tally-chart of the numbers of atoms on each side of the equation

- BE PATIENT!!!!
Construct equations and then balance them for each of the following:

- Magnesium (Mg) reacts with hydrochloric acid (HCl) to make magnesium chloride (MgCl$_2$) and hydrogen gas

- Ethane (C$_2$H$_6$) reacts with oxygen gas to make carbon dioxide and water

- Lead nitrate (Pb(NO$_3$)$_2$) reacts with aluminium chloride (AlCl$_3$) to make aluminium nitrate (Al(NO$_3$)$_3$) and lead chloride (PbCl$_2$)

- Barium nitride (Ba$_3$N$_2$) reacts with water to make barium hydroxide (Ba(OH)$_2$) and ammonia (NH$_3$)

- Ammonium perchlorate (NH$_4$ClO$_4$) reacting with aluminium to make aluminium oxide (Al$_2$O$_3$), aluminium chloride (AlCl$_3$), nitric oxide (NO) and water

**Extension:** Write a flow chart that can be followed to let you to balance equations.
Green chemistry is the design of chemical products and processes that reduce or eliminate the generation of hazardous substances.
Atom Economy

- Utilization of synthetic reactions and industrial processes that must be increasingly efficient to preserve raw materials and produce fewer and less toxic emissions

- Developed by Professor Barry Trost of Stanford University

- Atom economy looks at the level of efficiency of chemical reactions by comparing the molecular mass of atoms in the reactants with the molecular mass of useful compounds
Atom Economy

- Atom economy is the percentage of reactants changed to useful products

\[
\text{atom economy} = \frac{\text{mass of atoms in desired product}}{\text{mass of atoms in reactants}} \times 100\%
\]

- In an ideal chemical process the amount of reactants = amounts of products produced
- What does it mean if there is an atom economy of 100%?
Key Points

- The properties of solids, liquids and gases are due to the arrangement and motion of their particles

- Mixtures can be homogeneous or heterogeneous

- Equations must be balanced to ensure that mass is conserved

- A high atom economy means fewer atoms are being wasted
Lesson 2

Formulas and Composition by Mass
We Are Here

Quantitative Chemistry

Mole Ratios, Limiting Reactants and Theoretical Yields

Particle Theory
Chemical Equations
Formulas

Basic Concepts

The Mole
Molar Mass

Gases
Avogadro's Law
Ideal Gases

Solutions
Titrations

Main
SI Units

- System of International Units was developed to transcend all languages and cultures

- The International Bureau of Weights and Measures (BIPM) monitors the correct use of SI units in all applications of science

The SI Units

<table>
<thead>
<tr>
<th>Base quantity</th>
<th>Name</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>meter</td>
<td>m</td>
</tr>
<tr>
<td>Mass</td>
<td>kilogram</td>
<td>kg</td>
</tr>
<tr>
<td>Time</td>
<td>second</td>
<td>s</td>
</tr>
<tr>
<td>Electrical Current</td>
<td>ampere</td>
<td>A</td>
</tr>
<tr>
<td>Thermodynamic temperature</td>
<td>kelvin</td>
<td>K</td>
</tr>
<tr>
<td>Amount of substance</td>
<td>mole</td>
<td>mol</td>
</tr>
<tr>
<td>Luminous intensity</td>
<td>candela</td>
<td>cd</td>
</tr>
</tbody>
</table>
Stoichiometry

- Stoichiometry uses the quantitative relationships between amounts of reactants and products in a chemical reaction.

- Avogadro’s constant, $6.02 \times 10^{23} \text{mol}^{-1}$, enables us to make comparisons between chemical species.
Relative Atomic Mass, $A_r$

- The periodic table tells you the relative atomic mass of each element
  - This is the mass of an element relative to a $\frac{1}{12}$ of the mass of $^{12}$C.
  - It is a relative value, which means it has no units.
- Relative atomic mass has the symbol ‘$A_r$’

- For example carbon: $A_r(C) = 12.01$
  - The reason it isn’t a whole number is due to isotopes
Relative Atomic Mass, $A_r$

- Relative abundance of each isotope is a measure of the percentage that occurs in a sample of the element.
Relative Molecular or Formula Mass, $M_r$

- **Relative Molecular Mass** is calculated by adding up the $A_r$ for each atom in a molecule.

- The related term **relative formula mass** refers to the relative mass of one unit of a formula and is used for empirical formulas.

- No units as this is a ratio.
Calculating $M_r$

- **HCl**
  - $A_r(H) = 1.01$
  - $A_r(Cl) = 35.45$
  - $M_r = 1.01 + 35.45 = 36.46$

- **C$_2$H$_4$**
  - $A_r(C) = 12.01$
  - $A_r(H) = 1.01$
  - $M_r = 2 \times 12.01 + 4 \times 1.01$
  - $= 16.06$

- **H$_2$SO$_4$**
  - $A_r(H) = 1.01$
  - $A_r(S) = 32.06$
  - $A_r(O) = 16.00$
  - $M_r = 2 \times 1.01 + 32.06 + 4 \times 16.00$
  - $= 98.08$

- **Mg(OH)$_2$**
  - $A_r(Mg) = 24.31$
  - $A_r(O) = 16.00$
  - $A_r(H) = 1.01$
  - $M_r = 24.31 + 2 \times 16.00 + 2 \times 1.01$
  - $= 58.33$
Calculate $M_r$ for:

- $\text{Br}_2$
- $(\text{NH}_4)_2\text{SO}_4$
- $\text{C}_3\text{H}_8$
- $\text{C}_6\text{H}_{12}\text{O}_6$
Molar Mass

- **Molar mass** is the mass of one mole of a substance. It has the units of grams per mole, g mol\(^{-1}\)

<table>
<thead>
<tr>
<th>Element</th>
<th>Molar Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\text{H}_2\text{O})</td>
<td>(2(1.0) + 16.0 = 18.0) g/mol</td>
</tr>
<tr>
<td>((\text{NH}_4)_2\text{CrO}_4)</td>
<td>(2(14.0)+8(1.0)+52.0+4(16.0) = 152) g/mol</td>
</tr>
<tr>
<td>(\text{Ba(NO}_3)_2)</td>
<td>(137.3 + 2(14.0) + 6 (16.0) = 261.3) g/mol</td>
</tr>
</tbody>
</table>
Mole Calculations

- Volume of Gas (STP)
- Molar Mass
- Mass
- Representative Particles

1.00 mol
22.4 L
1.00 mol
22.4 L
1.00 mol
6.02 x 10²³ particles
6.02 x 10²³ particles
1.00 mol
1.00 mol
Types of Formulas

- Qualitative analysis: focuses on determining which elements are present
- Quantitative analysis: focuses on determining the relative masses of elements allowing to determine the exact composition
- Empirical Formula: simplest whole-number ratio of atoms or amount of each element present in a compound
- Molecular Formula: the actual number of atoms or amount of elements in one structural unit or one mole of the compound
  - Empirical formulas and molecular formulas can be the same
  - Because of their structure, ionic (and giant covalent) compounds do not form molecules so empirical formula is the only one relevant
## Empirical formula: Table summary

<table>
<thead>
<tr>
<th>Name of compound</th>
<th>Empirical formula</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen peroxide</td>
<td>HO</td>
<td>H₂O₂</td>
</tr>
<tr>
<td>Water</td>
<td>H₂O</td>
<td>H₂O</td>
</tr>
<tr>
<td>Glucose</td>
<td>CH₂O</td>
<td>C₆H₁₂O₆</td>
</tr>
<tr>
<td>Oxalic acid</td>
<td>HCO₂</td>
<td>H₂C₂O₄</td>
</tr>
<tr>
<td>Ethanol</td>
<td>C₂H₆O</td>
<td>C₂H₆O</td>
</tr>
<tr>
<td>Ethane</td>
<td>CH₃</td>
<td>C₂H₆</td>
</tr>
<tr>
<td>Ethylene</td>
<td>CH₂</td>
<td>C₂H₄</td>
</tr>
<tr>
<td>Caffeine</td>
<td>C₄H₅N₂O</td>
<td>C₈H₁₀N₄O₂</td>
</tr>
</tbody>
</table>
Percentage Composition by Mass

- If we divide the total mass of each element in a compound by the number of atoms

- For example ethanol, $C_2H_5OH$, $M_r = 46.08$

<table>
<thead>
<tr>
<th></th>
<th>C</th>
<th>H</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number Present</td>
<td>2</td>
<td>6</td>
<td>1</td>
</tr>
<tr>
<td>Multiply by $A_r$</td>
<td>2 x 12.01 = 24.02</td>
<td>6 x 1.01 = 6.06</td>
<td>1 x 16.00 = 16.00</td>
</tr>
<tr>
<td>Divide by $M_r$, convert to %</td>
<td>24.02/46.08 x 100 = 52.1%</td>
<td>6.06/46.08 x 100 = 13.1%</td>
<td>16.00/46.08 x 100 = 34.7%</td>
</tr>
</tbody>
</table>
Calculate % composition by mass for:

1. $\text{H}_2\text{O}$
2. $\text{Mg(OH)}_2$
3. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
4. $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
Example: Calculate the percentage by mass of sulfur in $\text{H}_2\text{SO}_4$

$\text{H}_2\text{SO}_4$

$2(\text{H}) = 2(1.01) = 2.02$

$1(\text{S}) = 1(32.06) = 32.06$

$4(\text{O}) = 4(16.00) = 64.00$

$\frac{32.06 + 64.00}{98.08} \times 100\% = 32.69\%$
Example: A sample of a compound contains 20% hydrogen and 80% carbon by mass and $M_r = 30.08$

- C : H
- 80% : 20%
- $80/12.01 = 6.67$: $20/1.01 = 20$
- $6.67/6.67 = 1$: $20/6.67 = 3$

n/a since no awkward decimals above

Empirical formula = $\text{CH}_3$: $30.08/(12.01 + 3 \times 1.10) = 2$

Molecular formula = $\text{CH}_3 \times 2 = \text{C}_2\text{H}_6$
Example: a sample of a compound contains 8.4% hydrogen, 65.2% carbon and 29.1% nitrogen by mass, and $M_r = 288.5$

\[
\begin{align*}
C & : \quad H \quad : \quad N \\
62.5\% & : \quad 8.4\% : \quad 29.1\%
\end{align*}
\]

\[
\begin{align*}
62.5/12.01= 5.20 & \quad 8.4/1.01= 8.31 & \quad 29.1/14.01=2.08 \\
5.20/2.08 = 2.5 & \quad 8.31/2.08 = 4 & \quad 2.08/2.08 = 1
\end{align*}
\]

\[
\begin{align*}
2.5 \times 2 = 5 & \quad 4 \times 2 = 8 & \quad 1 \times 2 = 2
\end{align*}
\]

Empirical formula = $C_5H_8N_2$ so $288.5/96.2 = 3$

Molecular formula = $C_5H_8N_2 \times 3 = C_{15}H_{24}N_6$
Lesson 3

Mole Ratios and Theoretical Yields
We Are Here

Quantitative Chemistry

- Mole Ratios, Limiting Reactants and Theoretical Yields
- The Mole
  - Molar Mass
- Gases
  - Avogadro's Law
  - Ideal Gases
- Solutions
  - Titrations

Particle Theory
Chemical Equations
Formulas
Basic Concepts
Stoichiometry

- Stoichiometry is the quantitative method of examining relative amounts of reactants and products.
- Percentage yield is vital in monitoring the efficiency and profitability of industrial processes.
Mole Ratios

- This is the ratio of one compound to another in a balanced equation.

- For example, in the equation
  \[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \]

- Hydrogen, oxygen and water are present in 2:1:2 ratio.
  - 0.2 mol of \( \text{H}_2 \) reacts with 0.1 mol of \( \text{O}_2 \) to make 0.2 mol \( \text{H}_2\text{O} \)
  - 5 mol of \( \text{H}_2 \) reacts with 2.5 mol of \( \text{O}_2 \) to make 5 mol of \( \text{H}_2\text{O} \)
  - To make 4 mol of \( \text{H}_2\text{O} \) you need 4 mol of \( \text{H}_2 \) and 2 mol of \( \text{O}_2 \)
Mole Ratios in Calculations

\[ n(\text{wanted}) = n(\text{given}) \times \frac{\text{wanteds}}{\text{givens}} \]

- wanted = the substance you want to find out more about
- given = the substance you are given the full info for
- \( n(\text{wanted}) \) = the number of moles you are trying to find out
- \( n(\text{given}) \) = the number of moles of you are given in the question
- wanteds = the number of wants in the balanced equation
- givens = the number of givens in the balanced equation
Example 1

What quantity of Al(OH)$_3$ in moles is required to produce 5.00 mol of H$_2$O?

$$2 \text{Al(OH)}_3 + 3 \text{H}_2\text{SO}_4 \rightarrow \text{Al}_2(\text{SO}_4)_3 + 3 \text{H}_2\text{O}$$

$n(\text{wanted}) = n(\text{given}) \times \frac{\text{wanted}}{\text{given}}$

- H$_2$O is given, Al(OH)$_3$ is wanted.
- $n(\text{Al(OH)}_3) = 5.00 \times (2/3) = 3.33 \text{ mol}$
Example 2

- What quantity of \( O_2 \) in moles is required to fully react with 0.215 mol of butane \( (C_4H_{10}) \) to produce water and carbon dioxide?

\[
2 \ C_4H_{10} + 13 \ O_2 \rightarrow 8 \ CO_2 + 10 \ H_2O
\]

\[n(\text{wanted}) = n(\text{given}) \times \frac{\text{wanted}}{\text{givens}}\]

- \( C_4H_{10} \) is given, \( O_2 \) is wanted.
- \( n(O_2) = 0.215 \times (13/2) \)
- \( n(O_2) = 1.40 \text{ mol} \)
Limiting Reagent

- In a reaction, we can describe reactants as being ‘limiting’ or in ‘excess’
  - **Limiting** – this is the reactant that runs out
  - **Excess** – the reaction will not run out of this reactant

- The limiting reactant will be your ‘given’ in further calculations:
  - Determining amounts of products formed
  - Determining amounts of other reactants used
Limiting Reagent

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

- For example, if you have 2.0 mol \(\text{H}_2\) and 2.0 mol \(\text{O}_2\)
  - \(\text{H}_2\) is the limiting reactant – it will run out
  - \(\text{O}_2\) is present in excess – there is more than enough

- To determine this, divide the quantity of each reactant by its coefficient in the equation. The smallest number is the limiting reactant:
  - \(\text{H}_2\): \(2.0 / 2 = 1.0\) – smallest therefore limiting
  - \(\text{O}_2\): \(2.0 / 1 = 2.0\)
Example 1: What quantity, in moles, of MgCl₂ can be produced by reacting 10.5 g magnesium with 100 cm³ of 2.50 mol dm⁻³ hydrochloric acid solution?

- Mg + 2HCl → MgCl₂ + H₂
- Determine limiting reagent:
  - Mg: \( \frac{10.5 \text{ g}}{24.31 \text{ g}} = 0.432 \text{ mol} \)
  - HCl: \( 0.100 \text{ dm}^3 \times 2.50 \text{ mol dm}^{-3} = 0.250 \text{ mol} \)
  - \( \frac{0.250}{2} = 0.125 \text{ mol} \) (smallest therefore is L.R.)

\[
\begin{align*}
\text{Mg} & \quad + \quad 2\text{HCl} \quad \rightarrow \quad \text{MgCl}_2 \quad + \quad \text{H}_2 \\
\text{Initial} & \quad 0.432 \text{ mol} \quad 0.250 \text{ mol} \quad 0 \text{ mol} \quad 0 \text{ mol} \\
\text{Change} & \quad -0.125 \text{ mol} \quad -(2)0.125 \text{ mol} \quad +0.125 \text{ mol} \quad +0.125 \text{ mol} \\
\text{Final} & \quad 0.307 \text{ mol} \quad 0 \text{ mol} \quad 0.125 \text{ mol} \quad 0.125 \text{ mol}
\end{align*}
\]
Example 2: What quantity, in moles, of carbon dioxide would be formed from the reaction of 12.0 mol oxygen with 2.00 mol propane, and how much of which reactant would remain?

\[
C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O
\]

- Determine limiting reagent:
  - \(C_3H_8\): 2.00 mol
  - \(O_2\): \(\frac{12.0\text{ mol}}{5\text{ (coefficient)}} = 2.4\text{ mol}\)

<table>
<thead>
<tr>
<th></th>
<th>Initial</th>
<th>Change</th>
<th>Final</th>
</tr>
</thead>
<tbody>
<tr>
<td>(C_3H_8)</td>
<td>2.00 mol</td>
<td>-2.00 mol</td>
<td>0 mol</td>
</tr>
<tr>
<td>(5O_2)</td>
<td>12.0 mol</td>
<td>-5 (2.00 mol)</td>
<td>2.00 mol</td>
</tr>
<tr>
<td>(3CO_2)</td>
<td>0 mol</td>
<td>+3 (2.00 mol)</td>
<td>6.00 mol</td>
</tr>
<tr>
<td>(4H_2O)</td>
<td>0 mol</td>
<td>+4 (2.00 mol)</td>
<td>8.00 mol</td>
</tr>
</tbody>
</table>
Theoretical, actual and percentage yield

- **Theoretical yield** is the maximum amount of product you would make if the limiting reactant was fully converted to product.
  - Use the limiting reactants maths to work this out

- **Actual yield** is the actual amount of product collected in after a reaction
  - Can be different from the theoretical yield

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

\[
\text{Percent Yield} = \left( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) \times 100\%
\]

Theoretical yield is the mass of product that you calculate using stoichiometry; it’s what you are supposed to be able to get from the reaction (theoretically). Actual yield is the mass of product that you actually obtain from the reaction in the lab.
Lesson 9

Molar Volumes of Gases
We Are Here

Quantitative Chemistry

- Mole Ratios, Limiting Reactants and Theoretical Yields
- The Mole
  - Molar Mass
  - Gases
    - Avogadro's Law
    - Ideal Gases
  - Solutions
    - Titrations

Particle Theory
- Chemical Equations
- Formulas

Basic Concepts
Molar Volume of a Gas

- The kinetic theory of gases is a model used to explain and predict the behavior of gases at a microscopic level.
- Postulates of assumptions that must be true for this theory to hold:
  1. Gases are made up of very small particles, separated by large distances. Most of the volume occupied by gas is empty space.
  2. Gaseous particles are constantly moving in a straight lines, but random directions.
  3. Gaseous particles undergo elastic collisions with each other and the walls of the container. No loss of kinetic energy occurs.
  4. Gaseous particles exert no force of attraction on other gases.
Molar Volume of a Gas

- Under conditions of STP, an ideal gas obeys these postulates
- At high temperature and low pressure, gases respond in ways that depart from the ideal gas behavior and exhibit behaviors of real gases
The Molar Volume of an Ideal Gas

- At standard temperature and pressure:
  - (STP, \( T = 273K, P = 1.01 \times 10^5 \) Pa or 100kPa)
  - Molar Volume of Ideal Gas = 22.7 dm\(^3\) mol\(^{-1}\)
Avogadro’s Law

- Calculate the moles of oxygen in a 6.73dm$^3$ sample of oxygen gas at STP

\[
\frac{6.73\text{dm}^3}{1\text{mole}} = \frac{22.7\text{dm}^3}{0.296\text{mol}}
\]

\[
\frac{V_1}{n_1} = \frac{V_2}{n_2}
\]
Boyle's Law

\[ P_1 V_1 = P_2 V_2 \] At a constant temperature, the volume of a fixed mass of an ideal gas is inversely proportional to its pressure

\[ V \propto \frac{1}{P} \]
Charles’ Law

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]

The volume of a fixed mass of an ideal gas at a constant pressure is directly proportional to its kelvin temperature.

\[ V \propto T \]
Gay-Lussac’s Law

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

The pressure of a fixed mass of an ideal gas at a constant volume is directly proportional to its kelvin temperature

\[ P \propto T \]
Combined Gas Law Equation

\[ \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \]

All the Law’s Combined

- Boyle’s Law: \( PV = k \)
- Charles’ Law: \( \frac{V}{T} = k \)
- Ideal Gas Law: \( PV = nRT \)
- Combined Gas Law: \( \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} = k \)
Ideal Gas Law

- Relationship between pressure, volume, temperature, and the amount, in mol, of gas particles

\[ PV = nRT \]

\( R = 8.31 \text{ JK}^{-1}\text{mol}^{-1} \)
Conversions

1,000,000cm³ = 1m³
- To convert m³ to cm³ multiply by 1,000,000
- To convert cm³ to m³ divide by 1,000,000

1000dm³ = 1m³
- To convert m³ to dm³ multiply by 1000
- To convert dm³ to m³ divide by 1000
We Are Here

Quantitative Chemistry

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- The Mole
  - Molar Mass
- Gases
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  - Ideal Gases
- Solutions
  - Titrations

Particle Theory
- Chemical Equations
- Formulas
- Basic Concepts
A solution is a homogeneous mixture of a solute that has been dissolved in a solvent. When dissolved in water the solution is described as an aqueous solution.
Concentration

This is the strength of a solution.
Concentration/Molarity

- The molar concentration of a solution is defined as the amount (in mol) of a substance dissolved in dm$^3$ of a solvent.
Concentration/Molarity

- **Units:**
  - mass per unit volume, g dm\(^{-3}\)
  - mole per unit volume, mol dm\(^{-3}\)
  - Parts per million (ppm)
    - One part in \(1 \times 10^6\) parts
    - 1 ppm = 1 mg dm\(^{-3}\)

\[
\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6
\]

- Square brackets are used to denote molar concentration
Titrations

- Titration involves using a solution whose concentration is known, to find the concentration of another which isn’t known.

- An exact volume of one solution is in a conical flask, a second solution is added to it from a burette.

- When the reaction reaches its ‘endpoint’, we record how much was added
  - There is always some kind of indicator which changes colour to tell us when we have reached the end.

- Determine the concentration of acids/bases
- Determine concentrations of other reactants
- Following the rate of a reaction
- Determining equilibrium constants
Titration

- Titration involves using a solution whose concentration is known, to find the concentration of another which isn’t known.

- An exact volume of one solution is in a conical flask, a second solution is added to it from a burette.

- It doesn’t matter which solution is in the burette, and which is in the conical flask.

- When the reaction reaches its ‘endpoint’, we record how much was added
  - There is always some kind of indicator which changes colour to tell us when we have reached the end.

  - Determine the concentration of acids/bases
  - Determine concentrations of other reactants
  - Following the rate of a reaction
  - Determining equilibrium constants
The mathematics of titrations

\[
\frac{C_1 V_1}{n_1} = \frac{C_2 V_2}{n_2}
\]

Where:
- \( n \) = coefficient in balanced equation
- \( C \) = concentration
- \( V \) = volume
Water of Crystallisation

- Some substances crystallise with water, this is called water of crystallisation.
- These substances are described as hydrated.
Airbag Stoichiometry