



# Topic 1: Quantitative Chemistry

## ▼ 1.1 Introduction to the particulate nature of matter and chemical change



### Definitions

Matter: Anything that takes up space (can refer to particles or mixtures)

Pure Substance: Particles will all look the same with constant composition

Element: Atoms all having the same number of protons

Molecule: Two or more elements chemically joined together

Compounds: Two or more different elements chemically joined together in a fixed ratio

Mixture: A combination of pure substances (not chemically bonded)

### **Homogeneous vs Heterogeneous Mixtures**

- Both contain more than one element/compound that are not chemically bonded
- Homogeneous mixtures will be the same mixture throughout, having a uniform composition
- Heterogeneous mixtures will have a different mixture throughout, visually shown to be non-uniform and different in different parts of the mixture

### **Chemical Equation**

- A Chemical Equation describes what happens in a chemical reaction with reactants and products

- They are shown like this: Reactants → Products
- Chemical equations usually include state symbols

### State Symbols

- (s): solid
- (l): Liquid
- (g): gas
- (aq): aqueous solution

### Physical and Chemical Changes

- In a physical change, there is no altering of chemical composition in the substances (no products are formed)
- Examples of physical change include change in state
- In a chemical change, there is a change in chemical composition, in which the atoms in the reactants are rearranged to form new products.

## ▼ 1.2 The mole concept and Avogadro's constant

### The Mole

- A mole is a unit of the amount of substance that contains the same number of specified particles as there are atoms in 12g of Carbon-12
- The mole is used to quantify the amount of particles present in grams. These are really large numbers, and are therefore expressed in moles.
- A mole is given by the symbol n
- The number of particles in 1 mole is defined by the Avogadro's constant, which is  **$6.02 \times 10^{23}$  particles**
- The formula to find the number of particles is the following



Number of Particles = Number of Moles x Avogadro's constant



Example:

To calculate the number of O<sub>2</sub> molecules in 1.5 mole of oxygen will be calculated as following

$$N = 6.02 \times 10^{23} \text{ (Avogadro's constant)} \times 1.5(\text{mol})$$

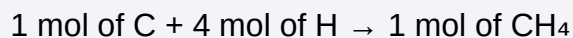
$$N = 9 \times 10^{23}$$

## Mole Relationships

- Moles of individual atoms that make up a moles of a molecule



Example:



- To find the number of mols in an element, you can use the following formula



Number of mols in an element = Number of mols of molecule x Number of elements within molecule

## Mole Concept

- One mole also weighs the element/compounds's mass number in grams (eg. 1 mol of oxygen is 15.999 grams, 1 atom of oxygen has a mass number of 15.999)
- In other words, the mass number of an element is equal to the relative molecular mass (molar mass) to an element as well
- The relative molecular mass (aka molar mass) has a unit of g mol<sup>-1</sup>

## Amount of moles

- To calculate the number of moles, this is the formula



$$\text{Moles} = \text{Mass} \div \text{Molar Mass}$$



Example:

To calculate the amount of mol in 32g of O<sub>2</sub>, we would do the following working

Since the molar mass of O<sub>2</sub> is 16, and total mass is 32g, the number of moles can be worked out by  $32 \div 16$

Therefore, there are 2 moles in 32g of O<sub>2</sub>

## Percentage Composition

- The values of molar masses of elements within compounds can be used to calculate percentage composition of a compound
- This can be worked out by the following equation



$$\text{Percentage Composition by Mass of Element} = \frac{\text{Molar Mass of the Element}}{\text{Molar Mass of the Compound}}$$

## Empirical Formula

- The empirical formula is the formula of a compound that shows the lowest whole number ratio of each type of atom
- This can be calculated through the following equation



Percentage Composition of Element/Mass : Percentage Composition of Element/Mass (to nearest whole number)

### **Molecular formula**

- The molecular formula is a formula of a compound that shows the number of each type of atom in a molecule
- The molecular formula is always a whole multiple of the empirical formula

### **Atom economy**

- The atom economy of a chemical reaction measures the amount of starting materials that become useful product, which acts as a measure of efficiency of a reaction
- Thus, a reaction with a higher atom economy has a better efficiency with less waste
- The formula for an atom economy is the following



Atom Economy =  $\frac{\text{Total Mass of Desired Products}}{\text{Total Mass of All Products or Reactants}}$

## **▼ 1.3 Reacting masses and volumes**

### **Limiting/Excess Reactants**

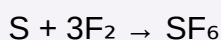
- The limiting reactant is the reactant that will be used up first in a chemical reaction
- The excess reactant is the reactant that will be left over after the reaction ends (when limiting reactant is used up)

- To find out which is the limiting reactant, divide the number of moles by the leading coefficient
- The reactant with the lower result is the limiting reactant



Example:

Find the limiting reactant when 5 moles of S react with 18 moles of F<sub>2</sub> in the following reaction



(For S):  $5 \div 1 = 5$

(For F<sub>2</sub>):  $18 \div 3 = 6$

As the answer for S was smaller, sulfur is the limiting reactant

### Percentage Yield

- The percentage yield of an experiment is the percent ratio between the actual and theoretical (predicted) yield in an experiment
- The actual experimental yield may differ from the theoretical yield due to a variety of factors such as
  - Measurement error
  - Some reactants may not react
  - Some of the reactants may form other products
- The formula to find percentage yield is the following



Percentage yield (%) = Actual Yield  $\div$  Theoretical Yield  $\times$  100

## Theory of an ideal gas

- The kinetic molecular theory of gases explains the behaviour of gases through the following properties
  - Particles are in continuous random motion in straight directions
  - Particles have perfect elastic collision (there is no net loss of kinetic energy from before and after collision)
  - The average kinetic energy of the particles are directly proportional to temperature
  - Volume of gas is negligible (so small they can be ignored)
  - Intermolecular forces are negligible (attraction is so small it can be ignored)
- Despite this theory, in reality, no gas is perfectly ideal and adheres to all behaviours above

## Ideal Gas Equation

- The ideal gas equation related different properties of gases as following
- With this equation, you can work out certain aspects (eg. pressure, temperature, etc.)



$$PV = nRT$$

P: Pressure (Pa)

V: Volume (m<sup>3</sup>)

n: Number of moles

T: Temperature (K)

R: Universal gas constant (8.31)

## Combined Gas Equation

- The combined gas law is an equation formed from the concepts that
  - Pressure is inversely proportional to volume at constant temperature
  - Volume is directly proportional to temperature (at constant pressure)
  - Pressure is directly proportional to temperature (at constant volume)
- The combined gas law is the following



$$(P_1V_1)/T_1 = (P_2V_2)/T_2$$

### Real vs Ideal Gases

- At high temperatures and low pressures, gases behave closest to an ideal gas
- An ideal gas has the following features
  - Particles are in continuous random motion in straight directions
  - Particles have perfect elastic collision (there is no net loss of kinetic energy from before and after collision)
  - The average kinetic energy of the particles are directly proportional to temperature
  - Volume of gas is negligible (so small they can be ignored)
  - Intermolecular forces are negligible (attraction is so small it can be ignored)

### Comparison Chart (Real Gases vs. Ideal Gases)

Real Gases	Ideal Gases
Gas particles have volume	Particle volume negligible
Intermolecular forces present	Intermolecular forces negligible

### Molar Volume



- Molar volume is the volume of one mole of gas at a certain temperature and pressure
- Avogadro's law states that 1 mol of any gas at standard temperature and pressure (273K and 100kPa), any gas will occupy 22.7dm<sup>3</sup>
- The formula to determine the volume of a gas is the following



Moles = Molar Volume ÷ Volume of Gas

## Molar Concentrations



### Definitions

Solute: The component in a solution that is dissolved

Solvent: The component in a solution that the solute dissolves in

Solution: The homogeneous mixture of the solute and solvent

Concentration: A measure of solute (mol) per solution (dm<sup>3</sup>)

- The following is the formula to calculate concentration



Concentration =  $n/v$

n: mole of solute

V: volume of solution

## Addition of solutions

- To calculate the new concentration in a mixture of solutions, add the number of moles from each individual solution and substitute in the formula above to find

the concentration of a solution

- Here are the steps to calculate the new concentration in added solutions



$$\text{Concentration of new solution} = ((c_1 \times v_1) + (c_2 \times v_2)) / (v_1 + v_2)$$

$c_1$ : concentration of first solution

$v_1$ : volume of first solution

$c_2$ : concentration of second solution

$v_2$ : volume of second solution

## Dilution

- Dilution is the process of adding more solvent to a solution
- By adding more solvent to a solution, solute particles will be more spaced out

### Dilution formula



$$C_1V_1 = C_2V_2$$