

Notes - HL/SL Chemistry - Exclusive for Chemyst Tuition Centre students only

# **Topic 1: Stoichiometric Relationships**

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







#### Explaining the heating curve :

Increasing temperature :

Melting & Boiling point linearity :

Particles are gaining energy and vibrating more violently. Heat energy supplied are used to overcome forces of attraction between

particles so that they can move around each other quickly.

#### Chemical Change:

Element :	A pure substance that contains only one type of atom.	НН
Atom :	An atom is the smallest part of an element that can exist.	Hydrogen (Molecule)
Compound :	Atoms of different elements chemically bond in fixed ratios to form compounds. E.g NaCl (ratio of Na: Cl is 1:1)	Au
Molecule :	A group of atoms held together by chemical bonding, representing the smallest fundamental unit of a chemical compound.	Gold (Atoms)
Homogenous mixture:	All components are in the <u>same</u> physical phase and have similar propert E.g. salt, water, air	ies.
Heterogeneous mixture:	Components are in <u>different</u> physical phases and have varying propertie E.g. Oil & Water – substances remain physically separate	s.
Physical Properties:	Melting & Boiling Point, Density, Electrical Conductivity, Hardness	
Chemical Properties :	Dictates how something reacts in a chemical reaction.	



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### **Chemical Equations**

A chemical reaction involves atoms joining together in different ways and electrons redistributing themselves between the atoms. However, atoms or electrons cannot be destroyed or created.

Balance with state symbols the following equations :

a) 
$$N_{2(\_)} + H_{2(\_)} \rightleftharpoons NH_{3(\_)}$$
  
b)  $C_4H_{10(\_)} + O_{2(\_)} \rightarrow CO_{2(\_)} + H_2O_{(\_)}$   
c)  $CaCO_{3(\_)} + HCl_{[\_]} \rightarrow LCO_{2(\_)} + H_2O_{(\_)}$ 



d) Zinc reacts with iron(III) chloride yielding zinc chloride plus iron.



f) Aluminum nitrate and sodium sulfide react to form aluminum sulfide and sodium nitrate.

#### Types of questions with balacing equation in IB HL/SL Chemistry Papers :

1) Sum of coefficients

2) Sum of lowest whole number coefficients

3) Ionic Equation balancing

### 1.2 The mole concept

Relative atomic mass  $(A_r)$  :Ratio of the average mass of the isotope of atoms to  $\frac{1}{12}$  mass of an atom of <sup>12</sup>C atom. No units.

Relative molecular mass  $(M_r)$  :Ratio of the average mass of the molecule to  $\frac{1}{12}$  mass of an atom of <sup>12</sup>C atom. No units.

Mole:

- A unit that describes the amount of substance. Substance could be atoms, molecules, ions or electrons.
- The term "mole" describes a number  $6.02 \times 10^{23}$
- One mole of substance =  $6.023 \times 10^{23}$  atoms, molecules, ions or electrons.
- 1 mole of  ${}^{12}C: 6.023 \times 10^{23} {}^{12}C$  atoms, 1 mole of NO<sub>3</sub>:  $6.023 \times 10^{23}$  of NO<sub>3</sub> ions



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Molar Mass (M)	:The mass of one mole of a substance. Units: g mol-1	
	E.g. Iron (Fe) has a Molar Mass of 55.85 g mol <sup>-1</sup> .	
	One mole of iron atoms (6.023 x 10 <sup>23</sup> iron atoms) has a mass of 55.85 g.	

Determining the no. of moles of substance (or Molar Mass, Sample mass):

no. of moles (mol) = 
$$\frac{\text{sample mass (g)}}{\text{molar mass (g mol^{-1})}}$$

a) Calculate the number of moles of magnesium atoms in 12.0 g of magnesium. [0.494 mol]

b) Calculate the mass of 0.50 mol of CH<sub>3</sub>COOH. [30.03 g]

Determining the mass of one molecule

mass of one molecule (g) = 
$$\frac{\text{molar mass } (\text{g mol}^{-1})}{\text{Avogadro's constant}}$$

Determining the percentage composition of a compound :

% by mass of an element = 
$$\frac{\text{no of atoms of the element x } A_r}{\text{relative molecular mass}}$$

a) Calculate the percentage by mass of each element present in  $C_6H_5NO_2$ . [C: 58.53%, H: 4.10%, N: 11.38%, O: 25.99%]

b) What mass of HNO3 contains 4.00 g of oxygen ? [5.62 g]

#### Determining the empirical formula, molecular formula :

Empirical formula: Simplest whole number ratio of the elements present in the compound (E.F of C2H4 is CH2)

Molecular formula: Actual number of atoms of the elements present in a molecule respectively.

Molecular formula: n x (Empirical Formula) & Molecular Mass: n x (Empirical Mass)

Structural formula: Shows the arrangement of atoms and bonds within a molecule. Used in organic chemistry.

You may be asked to find the empirical formula of a compound given its percentage composition by mass.

### Qn. A compound is found to be composed of 9% Hydrogen, 55% Carbon and 36% Oxygen by mass. Find

its empirical formula. Given that the molar mass of the compound is 132.18 g mol-1, determine its molecular

formula.



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	Carbon	Hydrogen	Oxygen
% in terms of grams	55	9	36
Number of moles of atoms (Divide by Ar)	$55 \div 12 = 4.58$	$9 \div 1 = 9$	$36 \div 16 = 2.25$
Divide by smallest	4.58 ÷ 2.25	9 ÷ 2.25	2.25 ÷ 2.25
Smallest ratio	2	4	1

So we have an empirical formula of  $C_2H_4O$ .

Empirical formula  $C_2H_4O$  has a molar mass of:  $(2 \times 12.01) + (4 \times 1.01) + (1 \times 16.00) = 44.06$  g mol<sup>-1</sup>

 $132.18 \div 44.06 = 3$ 

This tells us that the actual molar mass is 3 times larger than the empirical mass.

Therefore, the molecular formula is:C<sub>6</sub>H<sub>12</sub>O<sub>3</sub>

## Practice Questions from IBDP HL/SL Chemistry (May 2010-2011) for 1.1 & 1.2

- a) How many molecules are present in a drop of ethanol, C<sub>2</sub>H<sub>5</sub>OH, of mass  $2.3 \times 10^{-3}$  g? ( $L = 6.0 \times 10^{23}$  mol<sup>-1</sup>)
- b) The relative molecular mass of a gas is 56 and its empirical formula is CH<sub>2</sub>. What is the molecular formula of the gas?
- c) What is the total number of hydrogen atoms in 1.0 mol of benzamide,  $C_6H_5CONH_2$ ?
- d) Which is both an empirical and a molecular formula?

 $C_5H_{12} \qquad C_5H_{10} \qquad C_4H_8 \qquad C_4H_{10}$ 

e) What is the mass, in g, of one molecule of ethane, C<sub>2</sub>H<sub>6</sub>?

f) Which molecular formula is also an empirical formula?

 $PCl_3 \qquad C_2H_4 \qquad H_2O_2 \qquad C_6H_{12}O_6$ 

- g) What is the total number of nitrogen atoms in two mol of  $NH_4NO_3$ ?
- h) Which compound has the empirical formula with the largest mass?

 $C_2H_6 \qquad C_2H_4 \qquad C_2H_2 \qquad C_3H_6$ 



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i) An oxide of copper was reduced in a stream of hydrogen as shown below.



After heating, the stream of hydrogen gas was maintained until the apparatus had cooled. The following results were obtained:

	Mass of empty dish = 13.80 g Mass of dish and contents before heating = 21.75 g Mass of dish and contents after heating and leaving to cool = 20.15 g			
	(1)	Explain why the stream of hydrogen gas was maintained until the apparatus cooled.	(1)	
	(2)	Calculate the empirical formula of the oxide of copper using the data above, assuming complete reduction of the oxide.	(1)	
•••			(3)	
j)	Cro croc	cetin consists of the elements carbon, hydrogen and oxygen. Determine the empirical formula of cetin, if 1.00 g of crocetin forms 2.68 g of carbon dioxide and 0.657 g of water when it undergoes	(3)	
	com	iplete combustion.		
•••				
•••				
			(6)	



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# **Topic 1: Stoichiometric Relationships**

### 1.3 <u>Reacting Masses and Volume</u>

- Chemical Equations: Reactants react to produce substances which are known as products in a fixed relationship (i.e mole ratio).
- State symbols: (s) solid, (l) liquid (or molten), (g) gas and (aq) aqueous.
- Single arrow (→) : Reaction goes to completion. All the reactants have been consumed to form products completely.
- Reversible arrows (⇐): Reaction that has both reactant and product at the end (i.e equilibrium). The reaction
  proceeds in both directions, with reactants forming products and products reforming reactants
  simultaneously.
- Ionic equations: Used to describe the **dissociation of ionic compounds in water** or **precipitation** reactions.

For example:

NaCl (s) $\rightarrow$ Na <sup>+</sup> (aq) + Cl <sup>-</sup> (aq)	Dissociation of ionic compounds
$Ag^{+}(aq) + Cl^{-}(aq) \rightarrow AgCl(s)$	Precipitation of ionic compounds

### How to determine the limiting reagent / reactant ?

## $2 \text{ Ca}(s) + O_2(g) \rightarrow 2\text{CaO}(s)$

What does this chemical equation mean?

Moles: 2 moles of Ca reacts with 1 mole of  $O_2$  to produce 2 moles of CaO.

Mass: 80.2 g of Calcium reacts with 32.0 g of Oxygen to produce 112.0 g of Calcium Oxide.

What kind of questions will appear in exams?

Any random mass of the reactants will appear.

### Type 1: One of the reactant is in excess.

3.0 g of Calcium reacted in excess oxygen to produce Calcium Oxide. Find the mass of Calcium Oxide produced?

 $2 \text{ Ca}(s) + O_2(g) \rightarrow 2\text{CaO}(s)$ 

Answer:

No of moles or Amount of Calcium:  $\frac{\text{Sample Mass}}{\text{Molar Mass}} = \frac{3.0}{40.1} = 0.0748 \text{ mol}$ No of moles of CaO: 0.0748 mol (Mole ratio of Ca: CaO  $\rightarrow$  1:1)







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#### Type 2: Determining the limiting agent.

3.0 g of Calcium reacted with 4.0 g of oxygen to produce Calcium Oxide. Find the mass of Calcium Oxide produced.

Amount of Calcium: 0.0748 mol Amount of Oxygen : 4.0 / 32 .0 = 0.125 mol

Recall\* Moles: 2 mol of Ca reacts with 1 mol of O<sub>2</sub>

From Calcium's point of view: 0.0748 mol of Ca requires only 0.0374 mol of O<sub>2</sub>.

 $\rightarrow$  Oxygen is in excess as we have 0.125 mol.

From **Oxygen's point of view**: 0.125 mol of O<sub>2</sub> requires 0.250 mol of Ca. → Ca is limited as we only have 0.0748 mol.

Only Calcium will be completely used up to form Calcium Oxide. Mole ratio of Ca:CaO is 2:2 = 1:1Therefore, Amount of Ca = Amount of CaO = 0.0748

Mass of Calcium oxide: No of moles x Molar Mass =  $0.0748 \times 56.1 = 4.196 \text{ g}$ 

#### **Balancing Equations**

You will frequently be asked to balance equations in exams. Just ensure that each element has the same number of

moles present in the reactant and product side of the equation.

Example 1: Na (s) + H<sub>2</sub>O (l)  $\rightarrow$  NaOH (aq) + H<sub>2</sub> (g)

In this example we have 2 moles of hydrogen atoms on the reactant side, but 3 on the product side. Therefore, the

equation is not balanced. In order to balance it, we add numbers in front of each reactant/ product.

$$Na + 2H_2O \rightarrow 2NaOH + H_2$$

Now there are 4 moles hydrogen and 2 moles of Oxygen on both sides. However, the sodium is not balanced.

 $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ 

Practice Balancing Equations

- a)  $MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + CO_2 + H_2O$
- b)  $C_4H_8 + O_2 \rightarrow CO_2 + H_2O$
- c)  $Pb(NO_3)_2 + NaCl \rightarrow PbCl_2 + NaNO_3$



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#### Moles and Concentration

Concentration (mol dm<sup>-3</sup>) = 
$$\frac{\text{No of Mol (mol)}}{\text{Volume (dm3)}}$$

This equation is used to calculate the number of moles of a solute (dissolved chemical) in a particular sample of solution.

Example: Find the number of moles of NaCl present in 0.300 dm<sup>-3</sup> of 2.0 moldm<sup>-3</sup> aqueous NaCl solution.

Moles = Conc.  $\times$  Volume =  $2.0 \times 0.300 = 0.60$  mol

Percentage Yield

Percentage Yield (%) = 
$$\frac{\text{Actual Yield (g)}}{\text{Theoretical Yield (g)}} \times 100$$

Qn. 10.0g of calcium carbonate is decomposed and 3.0g of calcium oxide is obtained. Calculate the percentage yield.

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

No of moles of  $CaCO_3 = Mass / Molar Mass = 10.0g / 100.09 g mol<sup>-1</sup> = 0.0999 mol$ 

No of moles of CaO (if all of the CaCO<sub>3</sub> is decomposed) = 0.0999 mol (since mole ratio of CaCO<sub>3</sub> to CaO is 1:1)

Theoretical Yield of CaO = No of moles x Molar Mass =  $0.0999 \times (40.08 + 16.00) = 5.603 \text{ g}$ 

Percentage Yield =  $(3.0 / 5.60) \ge 100 = 53.5 \%$ 

#### Atom Economy

Atom Economy (%) = 
$$\frac{\text{Molar Mass of desired product (g)}}{\text{Molar mass of all reactants (g)}} \times 100$$

Qn. Calculate the atom economy of the decomposition reaction of calcium carbonate.

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

Molar Mass of  $CaCO_3 = 100.09 \text{ g mol}^{-1}$ 

Molar Mass of CaO =  $56.03 \text{ g mol}^{-1}$ 

Atom Economy =  $(56.03 / 100.09) \times 100 = 56.0\%$ 



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# **Topic 1: Stoichiometric Relationships**

#### Calculating Moles of Gas

At room temperature (25°C / 298K) and pressure, 1 mole of gas has a volume of 24dm3.

Moles of Gas= 
$$\frac{\text{Volume (dm^3)}}{24}$$

 $1 \text{ dm}^3 = 1000 \text{ cm}^3$ 

However, at standard temperature (0°C / 273K) and pressure, 1 mole of gas has a volume of 22.7dm<sup>3</sup>.

Moles of Gas= 
$$\frac{\text{Volume (dm^3)}}{22.7}$$

These two equations are only to be used to calculate the moles of GAS molecules. Example: Find the number of moles of hydrogen molecules in a container of volume 300 cm<sup>3</sup> at:

(a) Room temperature

(b) Standard temperature.

Recall:  $300 \text{ cm}^3 = 0.300 \text{ dm}^3$ (a)  $0.300 \div 24 = 0.0125 \text{ mol}$ (b)  $0.300 \div 22.7 \approx 0.0132 \text{ mol}$ 

#### Ideal Gases

When working with gases, we often approximate their behaviour to an "ideal gas" to simplify any necessary

calculations. There are 2 important assumptions to note about an ideal gas:

#### 1. The gas molecules do not interact with one another.

#### 2. The volume of each gas particle is zero.

Real gases **deviate** from this ideal behaviour at:

#### 1. High pressure -

The gas particles are pushed so closely together that they occupy a larger volume collectively, to the point where the volume of each particle is not negligible.

#### 2. Low temperature -

Each particle has very little kinetic energy, thus when they move close to other particles, they are more likely to experience some interaction due to intermolecular forces.



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The ideal gas equation:

PV = nRT

Symbol	Represents	Unit
Р	Pressure	Pascals (Pa)
V	Volume	$m^3$ (= 1000dm <sup>3</sup> )
n	no. of moles	mol
R	Gas constant = 8.31	J K <sup>-1</sup> mol <sup>-1</sup>
Т	Temperature	К
	-	

This equation allows us to calculate the molar mass of a gas, provided its mass is given.

**Example**: A sample of a gas is kept in a cylinder of volume 1.00 dm<sup>3</sup>, at 100 kPa and a temperature of 18.0°C. It is found to have a mass of 1.65g. Calculate the molar mass of the gas.

Convert all given data into appropriate units:

- $1 \text{ dm}^3 = 0.00100 \text{ m}^3$
- 100 kPa = 100 000 Pa
- 18.0°C = 291 K

PV = nRT

Rearrange to make *n* the subject:  $n = \frac{PV}{RT}$ 100000 × 0.001

$$n = \frac{100000 \times 0.001}{8.31 \times 291} \approx 0.0414$$

Recall: *Molar Mass* = 
$$\frac{Mass}{Moles}$$

$$Molar \ Mass = \frac{1.65}{0.414} \approx 39.9 gmol^{-1}$$

Alternatively, we can rearrange this equation to allow us to predict the response of a set amount of gas to a change in conditions.

$$\frac{PV}{T} = nR$$

For a fixed sample of gas (no. of moles does not change), the variables on the right side are both constants. This tells us that the final value of  $\frac{PV}{T}$  will not change even if one of the variables involved does.



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Example1 : A fixed sample of gas has the initial conditions of 80 kPa and a temperature of 278 K. The pressure is then increased to 110 kPa, with the volume remaining constant. Calculate the final temperature of the gas.

$$\frac{\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}}{T_2}$$

$$T_2 = \frac{T_1P_2V_2}{P_1V_1} = \frac{278 \times 110000}{80000} \approx 382K$$
**1 = Initial 2 = Final**

Example 2: A fixed mass of gas has a certain volume at a temperature of 50 °C. What temperature is required (in °C) to double its volume while keeping the pressure constant?

$$\frac{V_1}{V_2} = \frac{T_1}{T_2} = \frac{V_1}{2V_1} = \frac{50 + 273}{T_2}$$

 $T_2 = 2 (323) = 646 \text{ K} = 373 \text{ }^{\circ}\text{C}$ 





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# **Topic 1: Stoichiometric Relationships**

### Acid-Base Back Titration

A 0.8000 g of impure sample of K<sub>2</sub>CO<sub>3</sub> (138.2055 g/mol) is dissolved in enough water to make

 $200.0 \text{ cm}^3$  of solution A.

A 20.00 cm<sup>3</sup> aliquot of solution A is taken and put into an Erlenmeyer (conical) flask. To the flask is

added 20.00 mL of 0.1700 M HCl:

 $K_2CO_3(aq) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2O(l) + CO_2(g).$ 

The resulting solution is then titrated with 0.1048 M NaOH required 24.16 cm<sup>3</sup>.

 $NaOH(aq) + HCl(aq) \rightarrow H_2O(l) + NaCl(aq)$ 

Calculate the percentage purity of the impure sample. (Ans : 75.0%)





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# **Topic 1: Stoichiometric Relationships**

## Redox Back Titration

A 25.00 cm<sup>3</sup> of bleach (consisting of hypochlorite ions ClO<sup>-</sup>) was diluted in enough deionised water

to make 250.0 cm<sup>3</sup> of bleach solution inside a volumetric flask.

A 25.00 cm<sup>3</sup> aliquot of this solution is taken and put into an Erlenmeyer (conical) flask. Excess

iodide in acid was added to the flask:

 $ClO^{-}(aq) + 2I^{-}(aq) + 2H^{+}(aq) \rightarrow Cl^{-}(aq) + I_{2}(aq) + H_{2}O(l).$ 

The resulting solution is then titrated with 0.10 mol dm<sup>-3</sup> of sodium thiosulphate using the starch

indicator.

 $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^{-}(aq) + S_4O_6^{2-}(aq)$ 

Calculate the concentration of the bleach used originally. [0.442 mol dm<sup>-3</sup>]



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