O Level Pure Chemistry (G3) Topical Notes
(Latest 6092 Syllabus)

The O Level Chemistry Specialist

Topic 6: Chemical Calculations

CHEMICAL REACTIONS

Topic 6: Chemical Calculations

Learning Outcomes:

Formulae and Equation Writing

(a) state the symbols of the elements and formulae of the compounds mentioned in the syllabus.

(b) deduce the formulae of simple compounds from the relative numbers of atoms present and vice versa.

(c) deduce the formulae of ionic compounds from the charges on the ions present and vice versa.

(d) interpret chemical equations with state symbols.

(e) construct chemical equations, with state symbols, including ionic equations.

The Mole Concept and Stoichiometry

(a) define relative atomic mass, A_r.

(b) define relative molecular mass, M_r, and calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses.

(c) define the term mole in terms of the Avogadro constant.

(d) calculate the percentage mass of an element in a compound when given appropriate information.

(e) calculate empirical and molecular formulae from relevant data.

(f) calculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies $24dm^3$ at room temperature and pressure); calculations involving the idea of limiting reactants may be set

(knowledge of the gas laws and the calculations of gaseous volumes at different temperatures and pressures are not required).

(g) apply the concept of solution concentration (in mol/dm^3 or g/dm^3) to process the results of volumetric experiments (e.g. titration) and to solve simple problems

(appropriate guidance will be provided where unfamiliar reactions are involved).

(h) calculate % yield and % purity.

6.1 Chemical Formulae of Common Elements

Type of substance	Chemical Name	Chemical Formula
Metals	Sodium	Na
	Magnesium	Mg
Noble Gases	Argon	Ar
	Helium	He
Diatomic Molecules	Hydrogen	H_2
	Fluorine	F_2
	Chlorine	Cl_2
	Bromine	Br_2
	Iodine	I_2
	Nitrogen	N_2
	Oxygen	O_2

6.2 Polyatomic Ions

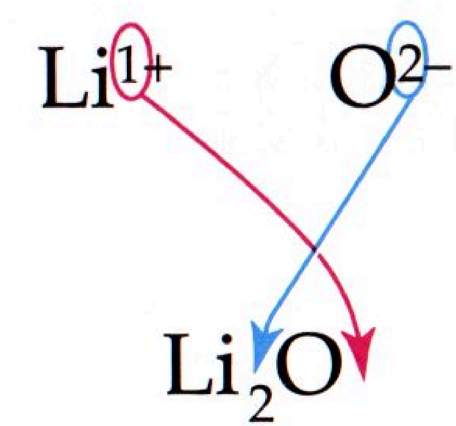
Polyatomic ions are groups of atoms covalently bonded together, but the group of atoms as a whole has a charge.

Polyatomic Ions
Carbonate (CO_3^{2-})
Nitrate (NO_3^-)
Sulfate (SO_4^{2-})
Hydroxide (OH^-)
Ammonium (NH_4^+)

6.3 Ionic Compounds

They are electrically neutral as charges of the cation and anion must balance each other.

We can deduce the chemical formula of the ionic compound from a cation and an anion using the 'criss-cross' method. Each O^{2-} gives a -2 charge. To balance this out to 0, we will need 2 Li^{+} atoms to produce a +2 charge.



6.4 Naming Ionic Compounds

The name of an ionic compound consists of the cation and the anion present. The cation is written before the anion.

The name of the cation doesn't change. However, modifications are made to the name of the anion.

Atom	Atom Name	Ion	Ion Name
Na	Sodium	Na^{+}	Sodium
Mg	Magnesium	Mg^{2+}	Magnesium
Cl	Chlorine	Cl^{-}	Chloride
F	Fluorine	F^{-}	Fluoride
Br	Bromine	Br^{-}	Bromide
I	Iodine	I^{-}	Iodide
O	Oxygen	O^{2-}	Oxide

6.5 Covalent Compounds

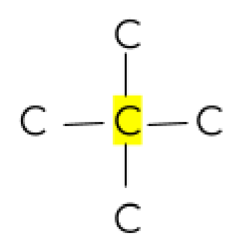
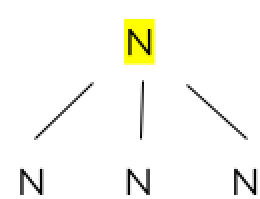
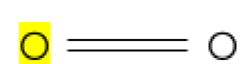
They are stable compounds and exist as simple covalent structures. Diatomic molecules (H_2, O_2, F_2) have covalent bonds, but are considered elements.

For other covalent compounds, the naming depends on the composition of the number of atoms in the molecule.

Prefix	Number of atoms
Mono	1
Di	2
Tri	3
Tetra	4

Eg. CO is Carbon Monoxide, CO_2 is Carbon Dioxide, CCl_4 is Tetrachloromethane.

The number of bonds a non-metal forms depends on how many more electrons they need to obtain a stable duplet or octet electron configuration.

Element	Max number of bonds	Illustration
C (Group 4)	4	
N (Group 5)	3	
O (Group 6)	2	
Cl (Group 7)	1	

6.6 Chemical Formulae of Common Compounds

Chemical Name	Chemical Formula
Carbon Dioxide	CO_2
Carbon Monoxide	CO
Ammonia	NH_3
Water	H_2O
Hydrogen Chloride	HCl
Sulfur Dioxide	SO_2
Nitrogen Dioxide	NO_2
Methane	CH_4

6.7 The State Symbols

State	Keywords
Solid (s)	Solid, Chips, Powdered, Precipitate
Liquid (l)	Molten, Water
Gas (g)	Effervescence
Aqueous (aq)	Solution, Dissolved in Water

Note:

Aqueous and Liquid are different! Aqueous is used for compounds that are dissolved in water; we call this a solution. Liquid is for melted substances. H_2O *is a liquid.*

State symbols must be written on the same line as the chemical formula.

NaCl (s) ✅

NaCl ❌
(aq)

6.8 Understanding Chemical Equations

$$2HCl\ (aq) + 2Na\ (s) \rightarrow 2NaCl\ (aq) + H_2\ (g)$$

Substances on the left-hand side of the arrow (HCl and Na) are known as the reactants; they represent the start of the reaction.	Substances on the right-hand side of the arrow (NaCl and H_2) are known as the products; they represent the outcome of the reaction.
Chemical equations must be balanced. State Symbols are optional unless required by the question.	

6.9 Balancing Chemical Equations

An equation is balanced when the number of atoms and/or the charges of the reactants and products are the same.

$$H_2 + O_2 \rightarrow H_2O$$

When you balance a chemical equation, imagine that there is a bubble surrounding the substances. This means that you cannot edit the numbers inside the bubble. In order to balance, we should only add numbers outside the bubbles, known as coefficients!

Steps	Actions
1	Write the equation: Start with the chemical equation given.
2	Count atoms: List the number of each type of atom on both sides.
3	Adjust coefficients: Change the coefficients in front of compounds to balance the number of atoms of each element.
4	Verify: Double-check that all atoms are balanced (same number on both sides).
5	Reduce coefficients (optional): If possible, simplify coefficients by finding the greatest common divisor.

Worked Example 1

$H_2 + O_2 \rightarrow H_2O$

Step 1: Write the equation: $H_2 + O_2 \rightarrow H_2O$

Step 2: Count atoms on the left and right sides of the equation:

	Left	Right
Hydrogen	2	2
Oxygen	2	1

Step 3: Adjust coefficients:

Balance oxygen first by adding a 2 in front of H_2O: $H_2 + O_2 \rightarrow 2H_2O$

However, now our H is unbalanced

	Left	Right
Hydrogen	2	4
Oxygen	2	2

Step 4: Continue balancing until all atoms are balanced (left and right numbers are the same).

Balance hydrogen by adding a 2 in front of H_2: $2H_2 + O_2 \rightarrow 2H_2O$

	Left	Right
Hydrogen	2	4
Oxygen	2	2

Step 5: Reduce coefficients (if needed): Coefficients are already in their simplest form.

Suppose we have: $4H_2 + 2O_2 \rightarrow 4H_2O$

Can be simplified by the smallest factor of 2

Final Chemical Equation: $2H_2 + O_2 \rightarrow 2H_2O$

Tricks to balance tough equations:

Suppose we have the equation: $C_3H_7 + O_2 \rightarrow CO_2 + H_2O$

Look for the least occurring element first. In this case, it is Carbon.

$C_3H_6 + O_2 \rightarrow 3CO_2 + H_2O$ (We put a coefficient of 3 in front of CO_2 to balance the number of C)

Next, we balance the next least common element. In this case, it is Hydrogen.

$C_3H_6 + O_2 \rightarrow 3CO_2 + 3H_2O$ (We put a coefficient of 3 in front of H_2O to balance the number of H)

However, there is a problem. The number of O atoms on the left is 2, while on the right is 9. They do not have a common factor. In such cases, use the 'temporary fractions' technique.
$C_3H_6 + \frac{9}{2}O_2 \rightarrow 3CO_2 + 3H_2O$ (We put a coefficient of $\frac{9}{2}$ in front of O_2 to balance the number of O, since $\frac{9}{2}$ x 2 = 9)

However, do not leave your answer in fractions, unless the question specifically wants 1 mole of C_3H_6. Thus, we will multiply the denominator (2) throughout.

Final Chemical Equation: $2C_3H_6 + 9O_2 \rightarrow 6CO_2 + 6H_2O$

6.10 What are Ionic Equations?

Ionic equations are chemical equations in which the dissolved ions are written as separated ions.

Only substances in the aqueous state have to be separated into their ions. This is not needed for solid, liquid and gaseous states.

Ionic equations must be balanced, and state symbols are not optional.

Steps	Actions
1	Write the balanced chemical equation: Start with the chemical equation given.
2	Separate aqueous compounds: Change aqueous compounds into their constituent ions, taking into account their composition. Eg. H_2SO_4 (aq) → $2H^+$ (aq) + SO_4^{2-} (aq)
3	Cancel like terms: Cancel any like terms you see on the left and right-hand side of the equation
4	Verify: Double-check that all elements are balanced (same number on both sides, charges are the same).
Worked Example 2 HCl (aq) + NaOH (aq) → NaCl (aq) + H_2O (l) Step 1: Write a balanced chemical equation Step 2: Separate aqueous compounds H^+ (aq) + Cl^- (aq) + Na^+ (aq) + OH^- (aq) → Na^+ (aq) + Cl^- (aq) + H_2O (l)	

Step 3: Cancel like terms

H^+ (aq) + ~~Cl^- (aq)~~ + ~~Na^+ (aq)~~ + OH^- (aq) → ~~Na^+ (aq)~~ + ~~Cl^- (aq)~~ + H_2O (l)

Step 4: Verify (balanced in terms of BOTH atoms and charges)

H^+ (aq) + OH^- (aq) → H_2O (l)

Worked Example 3

NaCl (aq) + $Ag(NO_3)$ (aq) → AgCl (s) + $NaNO_3$ (aq)

Step 1: Write a balanced chemical equation

Step 2: Separate aqueous compounds

Ag^+ (aq) + NO_3^- (aq) + Na^+ (aq) + Cl^- (aq) → AgCl (s) + Na^+ (aq) + NO_3^- (aq)

Step 3: Cancel like terms

Ag^+ (aq) + ~~NO_3^- (aq)~~ + ~~Na^+ (aq)~~ + Cl^- (aq) → AgCl (s) + ~~Na^+ (aq)~~ + ~~NO_3^- (aq)~~

Step 4: Verify (balanced in terms of BOTH atoms and charges)

Ag^+ (aq) + Cl^- (aq) → AgCl (s)

Pro Tips:

1. Neutralisation reactions (acid + alkali) will always lead to the same ionic equation of H^+ *(aq)* + OH^- *(aq)* → H_2O *(l)!*

2. Precipitation reactions involve only the ions of the precipitate
- Ag^+ *(aq)* + Cl^- *(aq)* → *AgCl (s)*
- Cu^{2+} *(aq)* + OH^- *(aq)* → $Cu(OH)_2$ *(s)*

3. Displacement reactions involve only the displaced ions
- *Zn (s)* + Cu^{2+} *(aq)* → Zn^{2+} *(aq)* + *Cu (s)*

Note: *Ions that were cancelled out are known as the 'spectator ions'! They do not take part in the reaction to form meaningful products.*

Try It Out!

N2017/P1/Q5

2. Iron(III) sulfate is composed of Fe^{3+} and SO_4^{2-} ions.
Which values of x and y in $Fe_x(SO_4)_y$ give the correct formula of iron(III) sulfate?
(N2017/P1/Q5)

	x	y
A	2	2
B	2	3
C	3	2
D	3	3

( )

N2019/P1/Q6

6. Which ions are present in solid chromium sulfate, $Cr_2(SO_4)_3$? (N2019/P1/Q6)

A Cr^{2+} and SO_4^{2-} **B** Cr^{3+} and SO_4^{2-}
C Cr^{2+} and SO_4^{3-} **D** Cr^{3+} and SO_4^{3-} ( )

N2020/P1/Q7

8. Calcium phosphate has the formula, $Ca_3(PO_4)_2$.
What is the charge on the phosphate ion? (N2020/P1/Q7)

A 2– **B** 3–
C 2+ **D** 3+ ( )

6.11 What is the Mole Concept?

A mole is a unit of measurement in Chemistry, just like grams or seconds.

As atoms and molecules are so small, the mole concept allows us to count atoms and molecules by weighing macroscopically small amounts of matter.

We use it to calculate the amount of reagents needed, or predict the amount of products we get.

Mole

One mole of any substance contains 6.02×10^{23} particles. This value is known as Avogadro's Constant.

Note:

Particles is a general word. It can refer to atoms, molecules, ions or even electrons.

It can be represented by n.
Eg. n_{Ca} has the same meaning as no. of moles of Ca.

$$n = \frac{no.\,of\ particles}{6.02 \times 10^{23}}$$

6.12 Mole Ratio

It can be derived from the coefficient of a balanced chemical equation.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

From the balanced chemical equation above, the mole ratio of Zn to HCl is 1:2. The mole ratio of Zn to $ZnCl_2$ is 1:1, etc.

6.13 Relative Atomic Mass (A_r)

A_r is defined as the average mass of an atom of an element compared to 1/12 the mass of a carbon-12 atom.

The A_r of an element can be determined by looking at its nucleon (mass) number in the periodic table. Eg. A_r of F = 19

6.14 Relative Molecular Mass (M_r)

M_r is defined as the average mass of a molecule of a substance compared to 1/12 the mass of a carbon-12 atom.

Eg. M_r of $CaCO_3$ = A_r of Ca + A_r of C + 3(A_r of O)
= 40 + 12 + 3(16)
= 100

Note:

Carbon-12 is used as the benchmark since it is the most commonly found element on Earth. A_r and M_r do not have units as they are relative to the mass of a Carbon-12 atom, effectively making them ratios rather than absolute masses. A Carbon-12 atom is considered to have 12 units of relative mass!

Common Mistake:

Thinking that A_r and M_r have units! A_r and M_r are ratios and do not have units!

6.15 Calculating how much of an Element is in a Compound

Percentage by Mass
$\frac{No.\ of\ atoms\ of\ the\ element\ in\ the\ compound \times Ar}{Mr\ of\ the\ compound}$ x 100%
% of O in $CaCO_3$ = $\frac{3 \times Ar\ of\ O}{Mr\ \ of\ CaCO3}$ x 100%   = $\frac{3 \times 16}{100}$ x 100%   = <u>48.0%</u>

Absolute Mass
$\frac{No.\ of\ atoms\ of\ the\ element\ in\ the\ compound \times Ar}{Mr\ of\ the\ compound}$ x mass of compound
Mass of O in 5g of $CaCO_3$ = $\frac{3 \times Ar\ of\ O}{Mr\ of\ CaCO3}$ = $\frac{3 \times 16}{100}$ x 5g   = <u>2.40g</u>

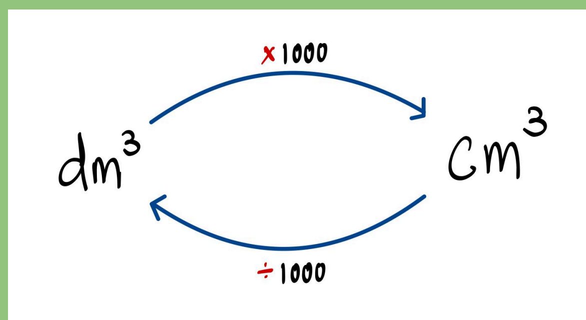
6.16 Molar Volume for Gases

At room temperature and pressure (25 °C and 1 atm), one mole of any gas occupies 24 dm^3 of space. This is regardless of their chemical formula and M_r.

Room Temperature and Pressure are usually abbreviated as r.t.p.

Pro Tip:

All calculations of volume MUST be done in dm^3. To quickly convert cm^3 to dm^3, simply divide cm^3 by 1000!



6.17 Molar Mass

Molar mass refers to the mass of 1 mol (6.02×10^{23} particles) of substance. The magnitude of the molar mass is equivalent to the relative atomic mass or relative molecular mass.

However, molar mass has units! (g/mol)

Using O_2 as an example

Molar Mass	Relative Atomic Mass (A_r)
32 g/mol	32

$$n = \frac{mass}{molar\ mass}$$

6.18 Stoichiometry for Gases

Regardless of their M_r, all gases occupy the same volume of $24dm^3$ under r.t.p. Thus, the volume of gas is directly proportional to the number of moles.

Therefore, the mole ratio of gases in a chemical reaction will also tell us about the volume ratio of gases.

$$3H_2\ (g) + N_2\ (g) \rightleftharpoons 2NH_3\ (g)$$

Suppose we have $10cm^3$ of N_2 gas, $30cm^3$ of H_2 gas is needed to fully react with N_2 since the mole ratio of N_2 to H_2 is 1:3.

$$n = \frac{volume}{24}$$

Common Mistake:

Using the above formula for solutions. The volume in this formula is used to represent the volume of gases, not solutions!

6.19 Concentration

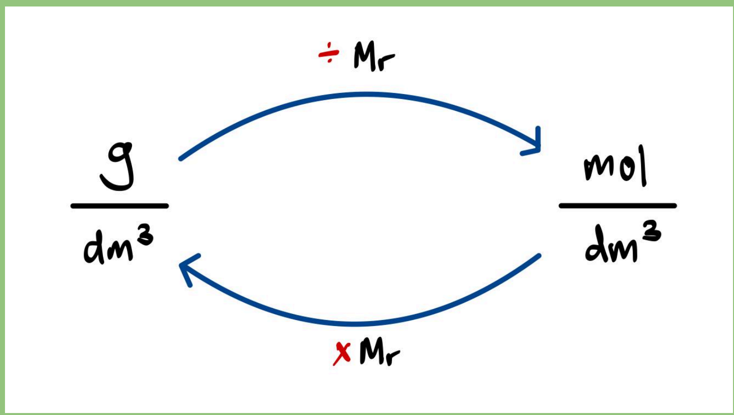
Concentration of a solution refers to the amount of solute in a solvent. We can measure concentration in g/dm^3 or mol/dm^3.

$$n = cV$$

g/dm^3	mol/dm^3
$\frac{mass}{volume}$	$\frac{mole}{volume}$

Pro Tip:

To convert between g/dm^3 and mol/dm^3, remember that Mass = Mole x M_r.



$$\frac{n1}{n2} = \frac{C1 \times V1}{C2 \times V2}$$

This is a shortcut method when calculating reactions that involve mole (n), concentration (c) and volume (V), especially useful in MCQs.

Worked Example 1

$10.0cm^3$ of 1 mol/dm^3 of NaOH is reacted fully with 2 mol/dm^3 of HCl. What was the volume of HCl needed to completely neutralise the NaOH?

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

$$\frac{n1}{n2} = \frac{C1 \times V1}{C2 \times V2}$$

Let the unknown be the top

n_1 = mole of HCl
c_1 = conc. of HCl
V_1 = volume of HCl

n_2 = mole of NaOH
c_2 = conc. of NaOH
V_2 = volume of NaOH

$$\frac{1}{1} = \frac{2 \times V1}{1 \times \frac{10}{1000}}$$

$V_1 = 0.00500dm^3 = 5cm^3$

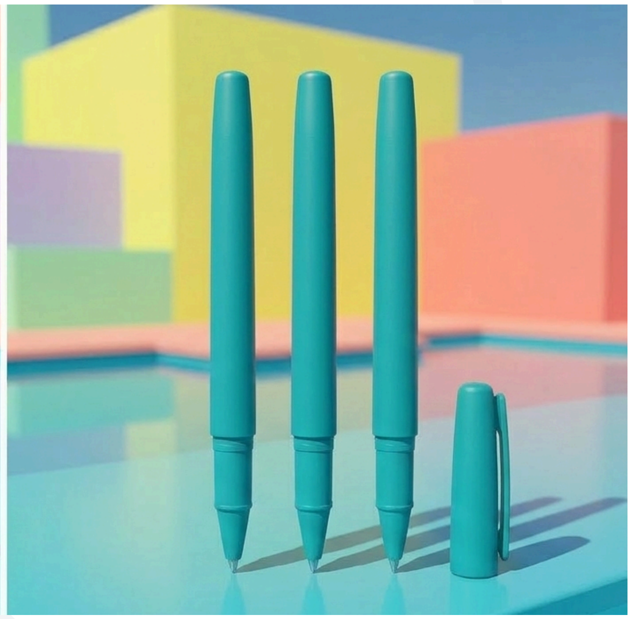
6.20 Limiting and Excess Reagents

In chemical reactions, reagents are not always used up fully.

The reaction will stop if one reactant is used up, even if other reactants are still available.

Analogy for Limiting and Excess Reagents

Suppose we need 1 pen and 1 cap to form a complete pen that is ready to be sold.



However, we have 3 pens and 1 cap. Regardless, we can only form 1 complete pen, despite having 3 pens.

How did we know that only 1 complete pen can be formed?
Internally, we deduced that the caps are limiting (not enough), and the pens were in excess. Therefore, we knew that we should use the number of caps to determine how many complete pens we can form. This is the idea of limiting and excess reagent.

Limiting Reagent	Excess Reagent
The reactant that is completely used up. It limits how much product is produced.	The other reactants that are not completely used up.

Note:

When calculating the moles of products produced, we will use the moles of the limiting reagent, not the excess!

Worked Example

$$2SO_2 + O_2 \rightarrow 2SO_3$$

Suppose we have 20cm^3 of SO_2 reacting with 5cm^3 of O_2.

Based on the mole ratio, 2 mol of SO_2 needs to react with 1 mol of O_2. This means that for 20cm^3 of SO_2, we will need 10cm^3 of O_2 for a complete reaction.

However, we only have 5cm^3 of O_2, making O_2 the limiting reagent, and SO_2 the excess!

***Pro Tip*:**

To identify whether a question requires us to determine the limiting/excess reagent first, we have to see if all the data are given to us to determine the moles of all reactants. If so, the question is a limiting/excess type of question!

6.21 Percentage Yield

Experimental (Actual) Yield	Theoretical (Expected) Yield
The actual amount of product obtained from the experiment	The maximum amount of products that could have been obtained in an ideal scenario based on chemical calculations.
Obtained from doing the experiment itself	Obtained by doing chemical calculations

Note:

The actual yield tends to be lower due to the incomplete reactions, impure reactants, or even side reactions.

$$\text{Percentage Yield} = \frac{Actual\ Yield}{Expected\ Yield} \times 100\%$$

Common Mistake:

Swapping the expected yield and actual yield, leading to a % yield > 100%, which is incorrect.

Worked Example 3

When 3.7g of ethanol was reacted with 6.0g of ethanoic acid, the mass of ethyl ethanoate obtained was 5.8g. What is the percentage yield?

$C_2H_5OH + CH_3COOH \rightarrow CH_3COOC_2H_5 + H_2O$

Firstly, this is a limiting/excess type of question, as the moles of both reactants can be found based on the data given

$n_{ethanol}$ = 3.7 / (2x12 + 6 + 16) = 0.08043 mol
$n_{ethanoic\ acid}$ = 6.0 / (2x12 + 4 + 2x16) = 0.1 mol

Since 1 mol of ethanol reacts with 1 mol of ethanoic acid, based on the mole ratio of 1:1, Ethanol is the limiting reagent!

Theoretical Yield of Ethyl Ethanoate = 0.008043 x (4x12 + 2x16 + 8) = 7.0778g

Percentage Yield = (5.8 / 7.0778) x 100% = <u>81.9%</u> (to 3 s.f)

6.22 Percentage Purity

Percentage purity measures how pure a substance is. The higher, the better.

$$\text{Percentage Purity} = \frac{\textit{Mass of Pure Substance}}{\textit{Mass of Impure Substance}} \times 100\%$$

Note:

The mass of impure reactant and pure product are usually given by the question.

Common Mistake:

Swapping the mass of pure substance and mass of impure substance, leading to a % purity > 100%, which is incorrect.

Worked Example 4

An impure sample of NaCl of mass 0.50g, given treatment with excess of $AgNO_3$ solution, produced 0.90g of AgCl as precipitate. Calculate the percentage purity of the sample.

$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

Firstly, this is not a limiting/excess type of question, as the moles of both reactants cannot be found based on the data given

$$n_{AgCl} = 0.90 \ / \ (108 + 35.5) = 0.006272 \text{ mol}$$

Since 1 mol of NaCl reacts with 1 mol of $AgNO_3$ based on the mole ratio of 1:1,

$$n_{NaCl} = n_{AgCl} = 0.006272 \text{ mol}$$

$$m_{pure\ NaCl} = 0.006272 \times (23 + 35.5) = 0.3669\text{g}$$

Percentage Purity of NaCl = (0.3669 / 0.5) x 100% = 73.4% (to 3 s.f)

6.23 Empirical and Molecular Formula

Empirical Formula	Molecular Formula
The simplest ratio of the constituent elements of a compound	Contains the total number of atoms of each element in a substance
Eg. CH_2O	Eg. $C_6H_{12}O_6$

Worked Example 5

Vitamin C (a carboxylic acid) contains 40.92% of C, 4.58% of H by mass. The experimentally determined relative molecular mass is 176. What is the empirical and molecular formula for Vitamin C?

First, ensure that all percentages add up to 100%
In this case, it is not. Since Vitamin C is a carboxylic acid, it is missing 54.50% of O.

Element	C	H	O
Mass in 100g (%)	40.92	4.58	54.50
A_r	12	1	16
No. of moles (n)	40.92 / 12 = 3.41	4.58 / 1 = 4.58	54.50 / 16 = 3.406
Simplest Ratio (Multiply by the smallest number of moles calculated above)	3.41 / 3.406 = 1	4.58 / 3.406 = 1.344	3.406 / 3.406 = 1

In this step, multiply the lowest possible number throughout to convert the simplest mole ratio to a whole number! **Optional if the simplest ratio is already a whole number!**	1 x 3 = 3	1.344 x 3 = 4.02 ≈ 4	1 x 3 = 3

Therefore, the empirical formula is $C_3H_4O_3$

Note:

Do not round off the simplest ratio unnecessarily. Eg. 4.12 cannot be rounded off to 4, but 4.02 can be rounded off to 4!

Let the molecular formula be $(C_3H_4O_3)n$

$(3 \times 12)n + (4 \times 1)n + (3 \times 16)n = 176$

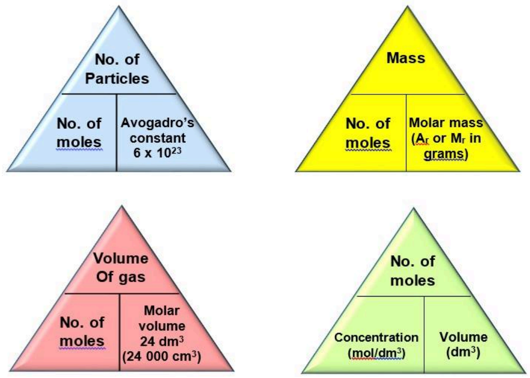
$88n = 176$

$n = 2$

Therefore, the molecular formula is $C_6H_8O_6$

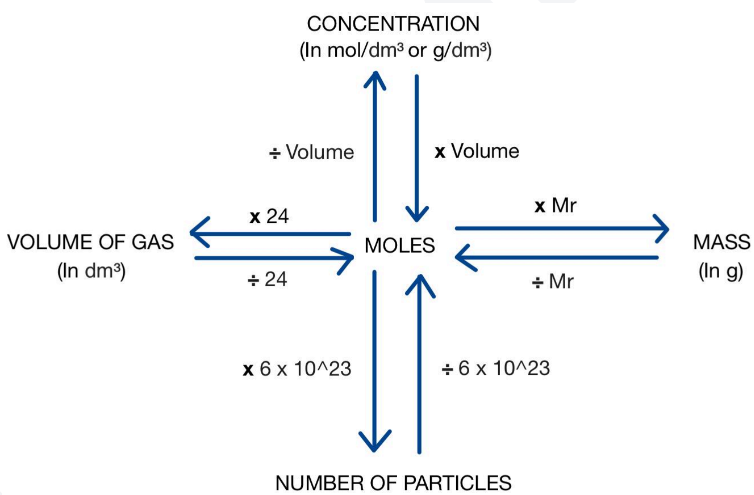
Common Mistake: *Not checking if the percentages of elements add up to 100%.*

6.24 Formula Triangles



To find one variable, cover up the variable you're solving for. The remaining visible variables will show you the relationship you need to use.

If the remaining variables are in a top-down relationship, divide the top by the bottom.	If the remaining variables are in a left-right relationship, multiply them together.



Pro Tips:

1. Always write out the balanced chemical equation to get the mole ratio of all compounds first! If data to find the mole of all reactants are given, find the limiting reactant!

2. Any intermediate working (working used for further calculations) should be expressed as 4 s.f on paper, but the full decimal number on the calculator should be used during calculations. Only the final answer should be expressed in 3 s.f!

Try It Out!

N2023/P1/Q26

13. What is the relative molecular mass, M_r, of $(NH_4)_2CO_3$? (N2023/P1/Q26)

A 77 **B** 80
C 94 **D** 96 ( )

N2016/P3/A5

1. Acid **J** has a relative molecular mass of 98. A 200 cm³ aqueous sample contains 196 g of **J**.
   (a) Calculate the concentration of **J** in g/dm³. [1]
   (b) Calculate the concentration of **J** in mol/dm³. [1]
   (N2016/P3/A5a)

N2019/P1/Q7

7. 0.1 mol/dm³ hydrochloric acid reacts with 25 cm³ of 0.2 mol/dm³ aqueous sodium carbonate.
   The equation for this reaction is shown.

$$2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$$

   What is the volume of acid required to completely react with this volume of sodium carbonate? (N2019/P1/Q7)

A 6.25 cm³ **B** 25 cm³
C 50 cm³ **D** 100 cm³ ( )

N2016/P1/Q5

1. 2.0 g of calcium are completely burnt in pure oxygen.

$$2Ca + O_2 \rightarrow 2CaO$$

   Which volume of oxygen is used in this reaction at room temperature and pressure? (N2016/P1/Q5)

A 0.025 dm³ **B** 0.050 dm³
C 0.60 dm³ **D** 1.2 dm³ ( )

N2024/P1/Q7

15. 25.6 g of metal M reacts with sulfur to form 32.0 g of metal sulfide, M_2S.
   What is the relative atomic mass of metal M? (N2024/P1/Q7)

A 32 **B** 64
C 128 **D** 256 ( )

N2017/P1/Q6 / N2020/P1/Q8 / N2024/P1/Q8

3. Glucose has a relative molecular mass, M_r, of 180.
How many grams of glucose are added to 50 cm^3 of water to make a solution of concentration 0.4 mol/dm^3? (N2017/P1/Q6 / N2020/P1/Q8 / N2024/P1/Q8)

A 3.6 **B** 9

C 36 **D** 72 ( )

N2018/P1/Q5

4. 20 cm^3 of ethene are reacted with 70 cm^3 of oxygen.
The equation for the reaction is shown.

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

[All volumes are measured at r.t.p.]
What is the **total** volume of gas remaining at the end of the reaction? (N2018/P1/Q5)

A 40 cm^3 **B** 50 cm^3

C 80 cm^3 **D** 90 cm^3 ( )

N2018/P1/Q12

7. Bromobutane, C_4H_9Br, can be made from butanol using the reaction shown.

$$C_4H_9OH + HBr \rightarrow C_4H_9Br + H_2O$$

In an experiment, 10 g of butanol produced 12 g of bromobutane.
What is the percentage yield of bromobutane?
[M_r: C_4H_9OH, 74; C_4H_9Br, 137] (N2018/P1/Q12)

A 45% **B** 54%

C 65% **D** 83% ( )

N2024/P1/Q29

31. Sample X contains 0.050 moles of chlorine gas.
Sample Y contains 0.10 moles of carbon tetrachloride, CCl_4.
How many chlorine atoms does each sample contain? (N2024/P1/Q29)

	sample X	sample Y
A	3.01×10^{22}	6.02×10^{22}
B	3.01×10^{22}	2.41×10^{23}
C	6.02×10^{22}	6.02×10^{22}
D	6.02×10^{22}	2.41×10^{23}

( )

N2023/P1/Q14

23. What contains the smallest number of particles? (N2023/P1/Q14)

A 2 g of hydrogen molecules
B 4 g of helium atoms
C 8 g of oxygen molecules
D 22 g of carbon dioxide molecules ( )

N2023/P1/Q15

24. A chemical in a paint has the following composition.

cobalt	phosphorus	oxygen
48.2%	16.9%	34.9%

What is its empirical formula? (N2023/P1/Q15)

A Co_2P_4O **B** Co_3PO_2
C $Co_3P_2O_8$ **D** $Co_5P_7O_2$ ( )