Chapter 6
The Mole Concept & Stoichiometry

CHAPTER ANALYSIS

FOCUS

- Relatively straight forward chapter
- 7 **key** concepts with Practical Tips

EXAM

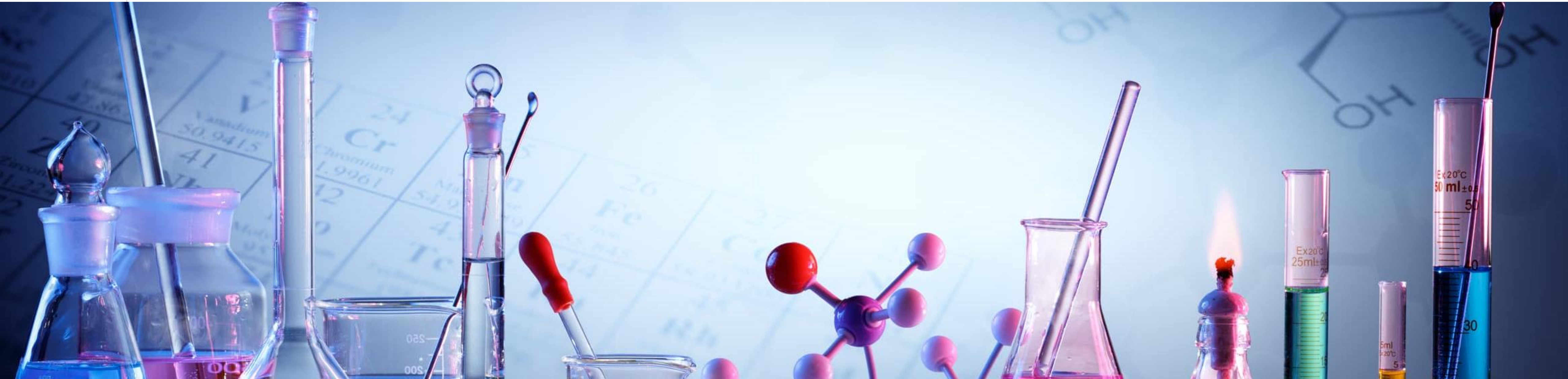
- Always tested
- The foundations for all calculations

WEIGHTAGE

- Heavy overall weightage as it is always incorporated into other chapters

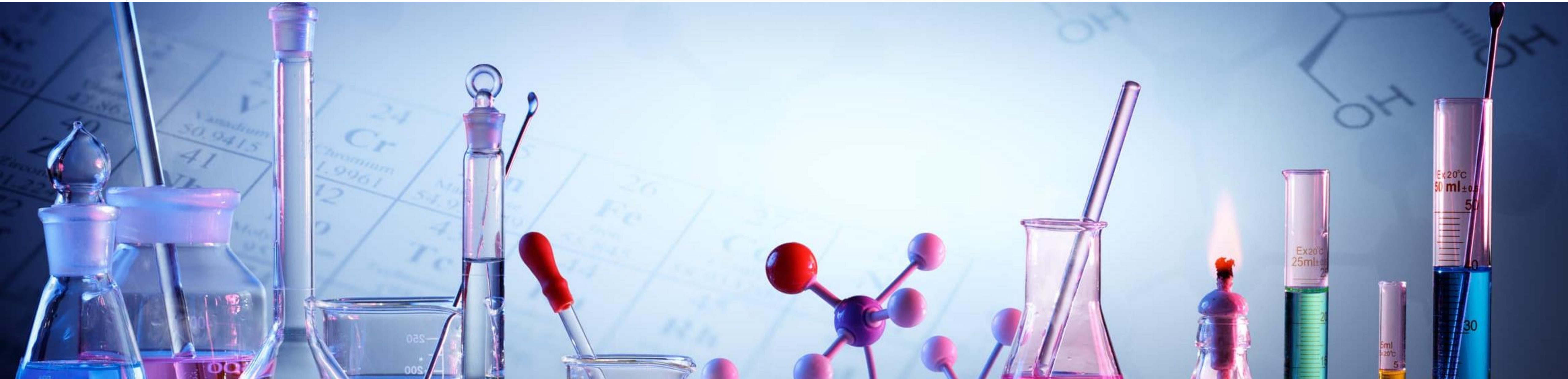
KEY CONCEPT

Atoms & Other Particles
Atomic Mass & Molecular Mass
Mole & Related Formula
Empirical Formula & Molecular Formula



KEY CONCEPT

Balancing Chemical Equation & Stoichiometric Mole Ratio
Concentration of Solution
Titrations (Volumetric Analysis)
Practical



Atoms & Other Particles

ATOMS

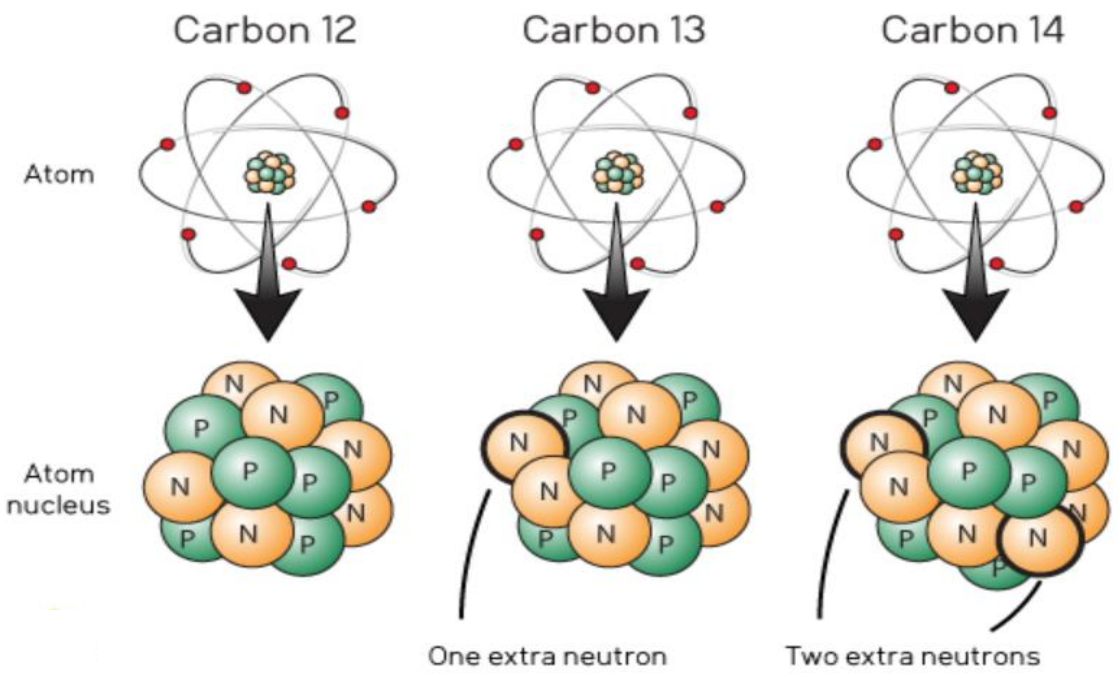
The **smallest particle** of an element.

MOLECULES

A **group of atoms** covalently bonded together in an element or a compound.

ISOTOPES

Atoms of the same element with the **same number of protons/electrons** but **different number of neutrons**



Atomic Mass & Molecular Mass

<u>**PROTON NUMBER**</u>	**VS**	<u>**NUCLEON NUMBER**</u>
Refer to the number of **protons/electrons** in an atom.		Refer to the **total number of protons and neutrons** in an atom.

<u>RELATIVE ATOMIC MASS,</u> A_r

Average mass of **one atom** of the element as compared to $\frac{1}{12}$ of the mass of one atom of carbon-12.

<u>RELATIVE ISOTOPIC MASS,</u> A_r

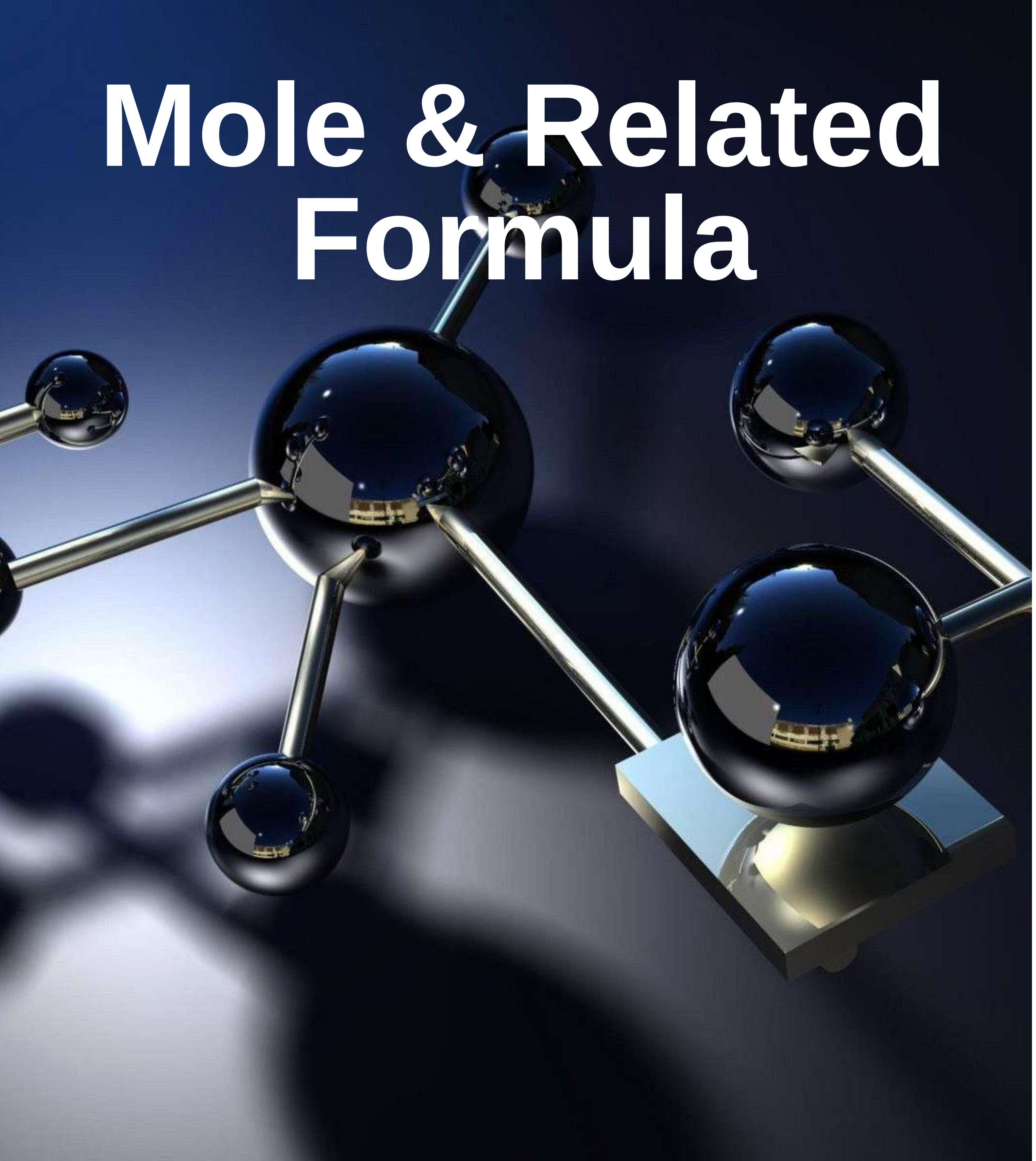
Mass of **one atom** of an isotope of an element as compared to $\frac{1}{12}$ of the mass of one atom of carbon-12.

Average Isotopic Mass = Σ [(relative isotopic mass)(% abundance)]

<u>RELATIVE MOLECULAR MASS</u>, M_r

Average mass of **one molecule** of the substances as compared to $\frac{1}{12}$ of the mass of one atom of carbon-12.

Mole & Related Formula



MOLES

One mole contains exactly 6.02×10^{23} (Avogadro constant) elementary particles.

$$\text{Mol} = \frac{\text{Number of molecules}}{\text{Avogadro Constant}}$$

MOLAR MASS

It is numerically equal to A_r or M_r with units of g mol^{-1}.

$$\text{Mol} = \frac{\text{Mass of substance}}{\text{Molar mass of substance}}$$

MOLAR VOLUME

The volume occupied by one mole of gas with units of $\text{dm}^3\ \text{mol}^{-1}$.
(Commonly used values = 24.0 $\text{dm}^3\text{mol}^{-1}$ & 22.7 $\text{dm}^3\text{mol}^{-1}$)

$$\text{Mol of gas} = \frac{\text{Volume of gas}}{\text{Molar volume}}$$

EMPIRICAL FORMULA

Simplest formula which the **ratio of atoms of the different elements** in the compound.

Determination of Empirical Formula

Step 1: Write down the mass ratio of the elements involved

Step 2: Convert the mass ratio into mole ratio

Step 3: Determine the smallest whole number values for the mole ratio obtained.

MOLECULAR FORMULA

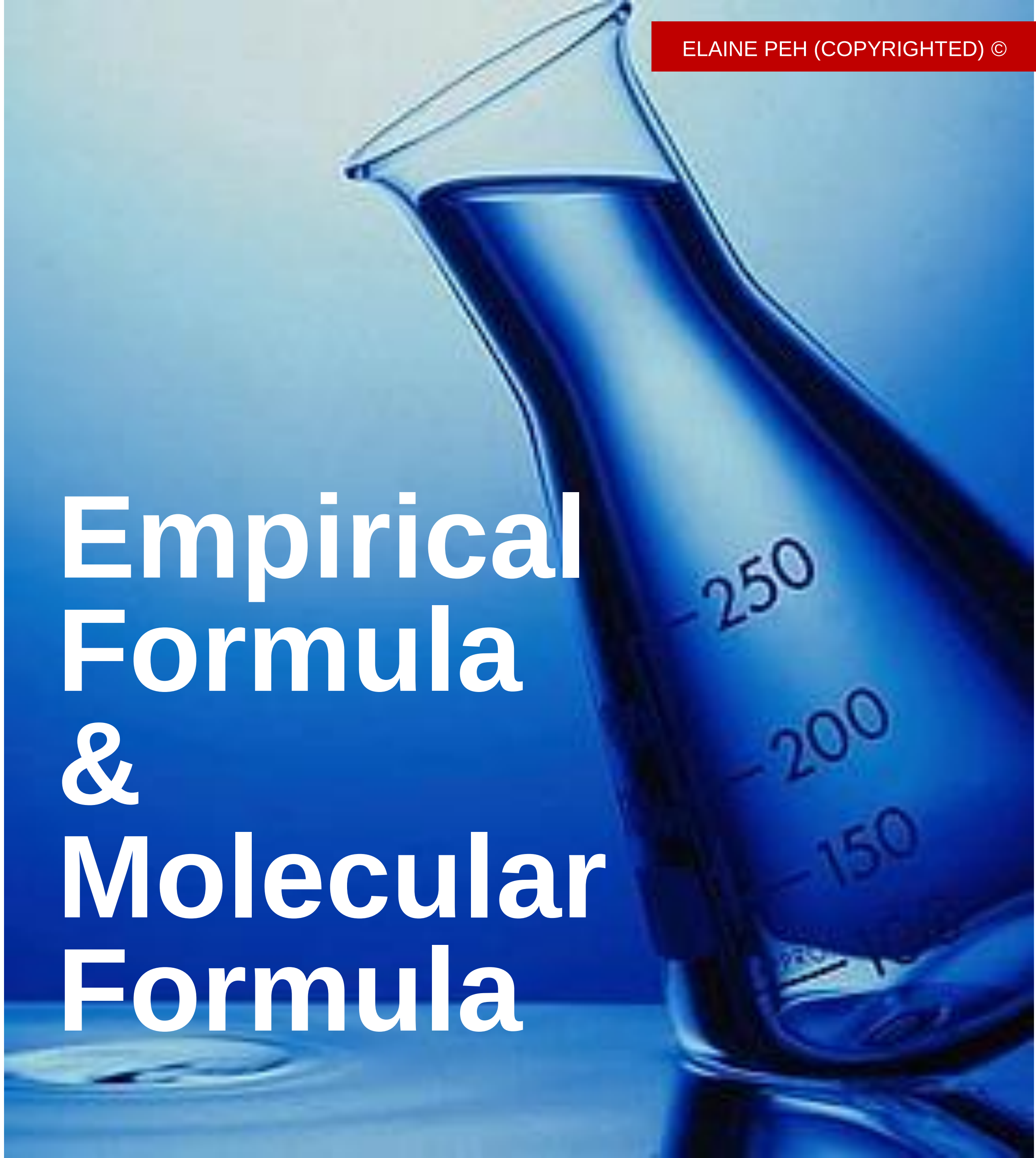
Actual formula which shows the **actual** number of atoms of each element in **one molecule** of the compound.

Determination of Molecular Formula from Empirical Formula

Step 1: Determine the empirical formula first

Step 2: Determine $_n$ using the given M_r*

*M_r of molecular formula = (M_r of empirical formula)$_n$

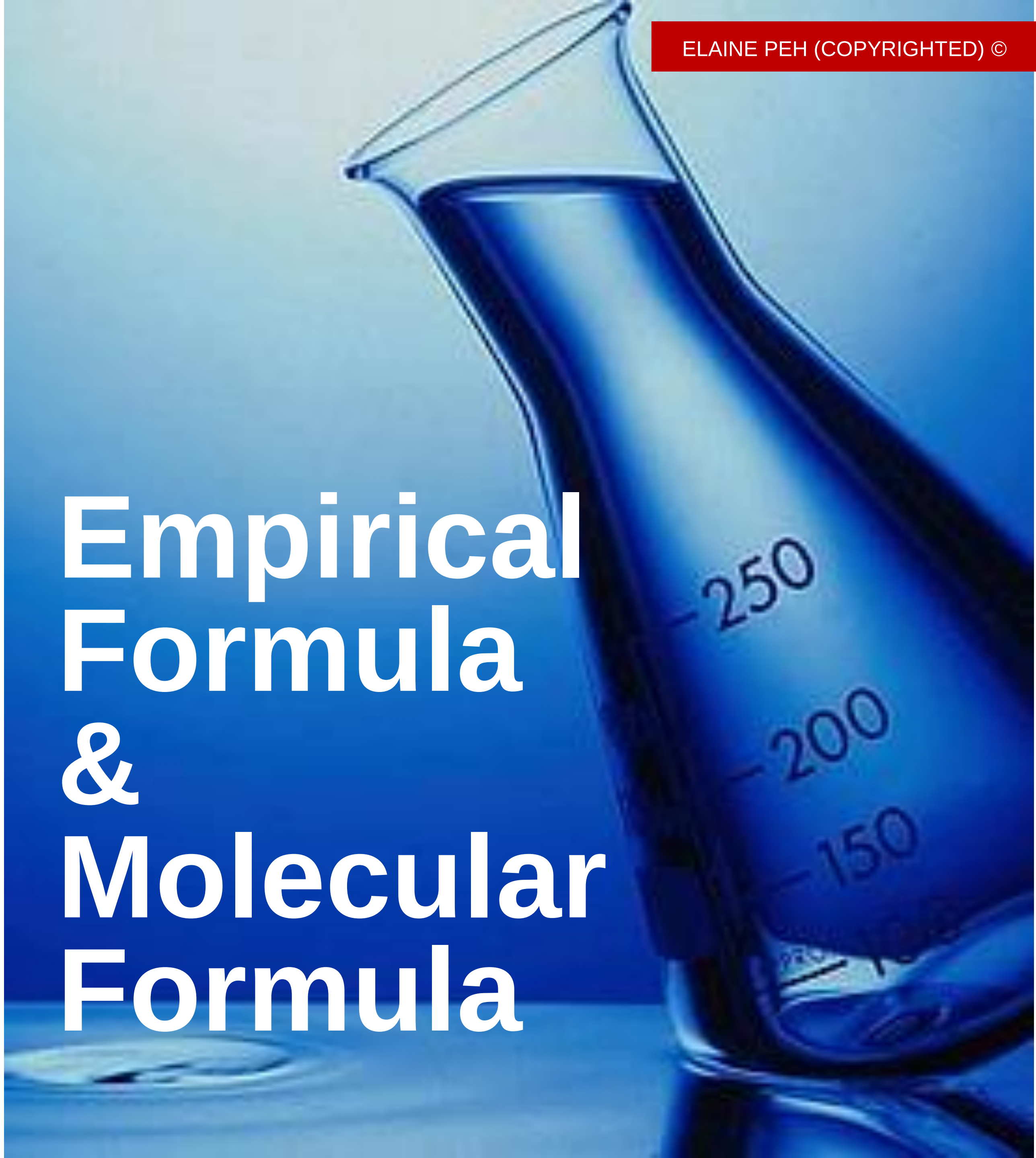


MOLECUALR FORMULA COMBUSTION DATA OF HYDROCARBON

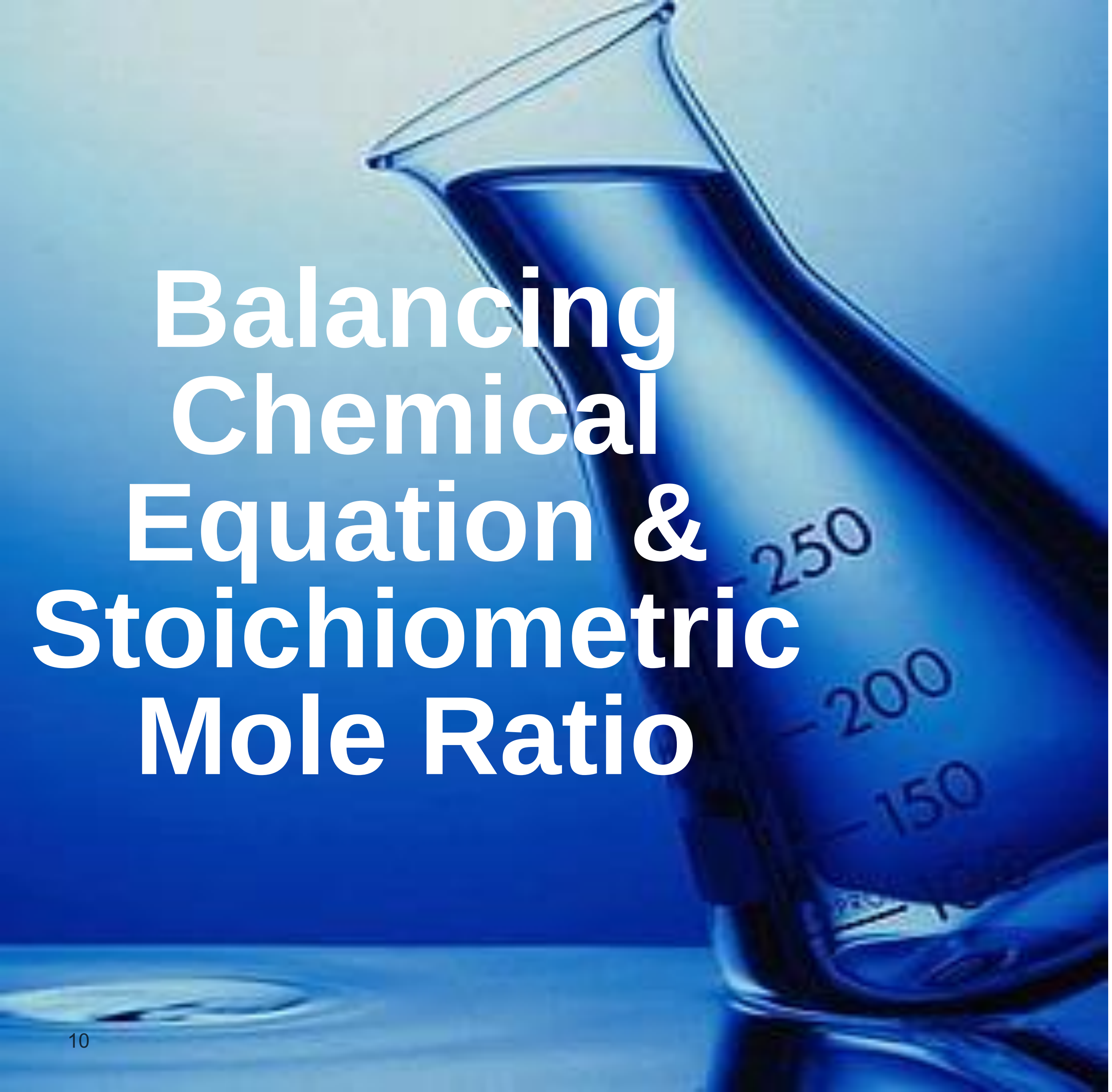
1) By the use of Mass Ratio
- All the weight from a pure element originated from the compound

2) By the use of Volume Ratio
- Only for complete combustion

$$C_xH_y + (x + \frac{y}{4})O_2\ (g) \rightarrow xCO_2\ (g) + \frac{y}{2}H_2O\ (l)$$



Empirical Formula & Molecular Formula



Balancing Chemical Equation & Stoichiometric Mole Ratio

CHEMICAL EQUATION

A **balanced chemical equation** shows the **mole ratio (stoichiometric ratio)** of reactants and products.

LIMITING REAGENT

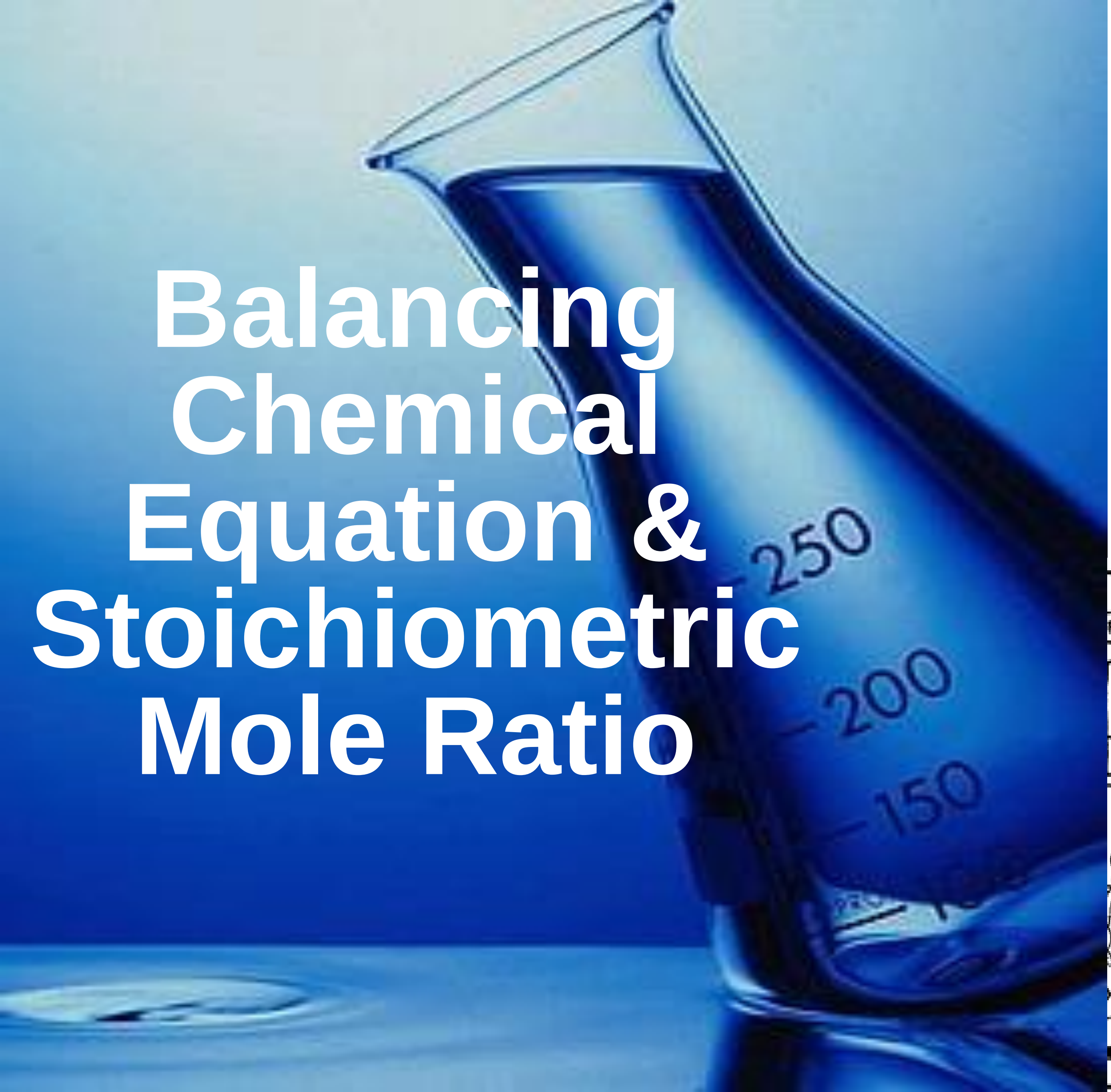
Sometimes **reactants are not always added in stoichiometric amounts**. When **only one reactant is completely consumed** at **the end** of the reaction, this **reactant is the limiting reagent**.

THEORETICAL YIELD

It is the yield calculated **based on the chemical equation** and the **amount of reactants present.**

EXPERIMENTAL (ACTUAL) YIELD

It is the actual yield calculated **based on the amount of products actually obtained** in a reaction.
Note: The difference between theoretical yield and experimental yield is due to experimental losses

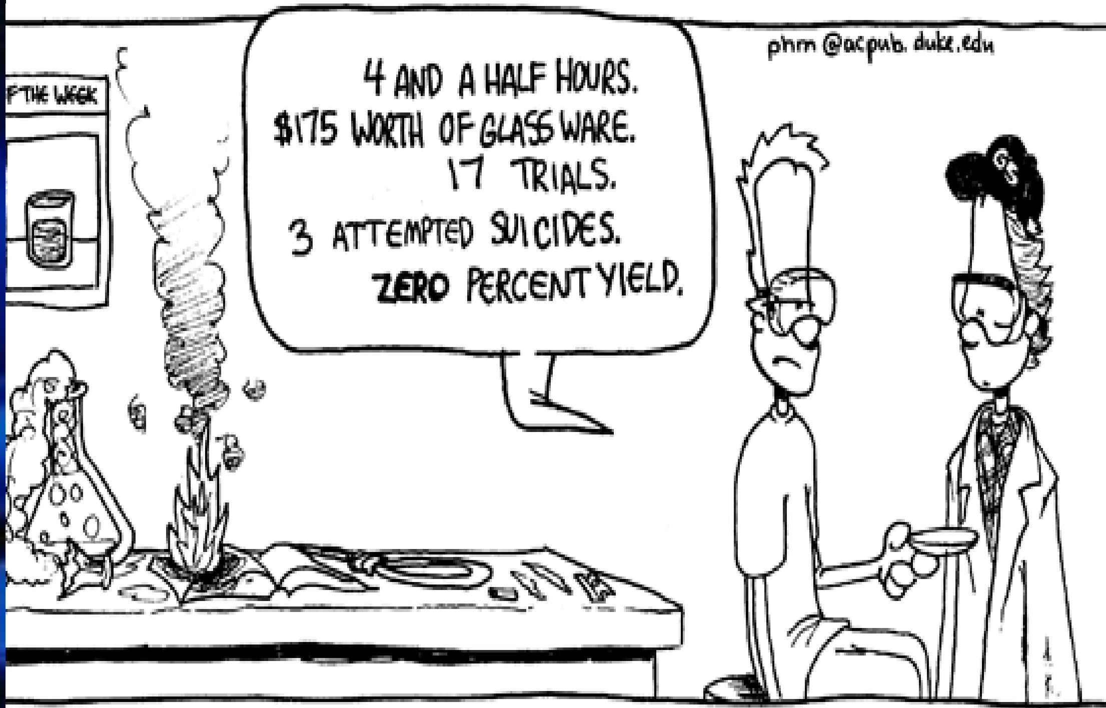


Balancing Chemical Equation & Stoichiometric Mole Ratio

PERCENTAGE YIELD

Step 1: Write out the balanced chemical equation
Step 2: Write out the mol ratio
Step 3: Identify any limiting reagent (if have)
Step 4: Calculate the theoretical and experimental yield
Step 5: Calculate the percentage yield using the formula

$$\text{Percentage Yield} = \frac{\text{Experimental Yield}}{\text{Theoretical Yield}} \times 100\%$$



Concentration of Solution

SOLUTION
A **homogeneous mixture** that consists of a substance dissolved in another.

SOLVENT
The substance in **greater quantity**, usually a liquid.

SOLUTE
The substance in **smaller quantity**.

MOLAR CONCENTRATION
The amount of solute (in **mol**) per unit volume of solution, with units of mol dm^{-3} .

$$\text{Molar concentration} = \frac{\text{Mol of substance}}{\text{Volume}}$$

MASS PER UNIT VOLUME
The amount of solute (in **mass**) per unit volume of solution, with units of g dm^{-3}.

$$\text{Mass Concentration} = \frac{\text{Mass of substance}}{\text{Volume}}$$

Concentration of Solution

DILUTION

Dilution is the process of **adding solvent** to a **known volume of the solution** to **lower its concentration**. This **increases the volume of the solution** but as **no additional solute is added**, the **number of moles of solute remains unchanged (*n*)**.

$$n_{original} = n_{diluted}$$

$$c_{original} \times V_{original} = c_{diluted} \times V_{diluted}$$

DILUTION FACTOR

It is a measure of how much the solution is diluted, and is given by the ratio of the volume after dilution to the original volume.

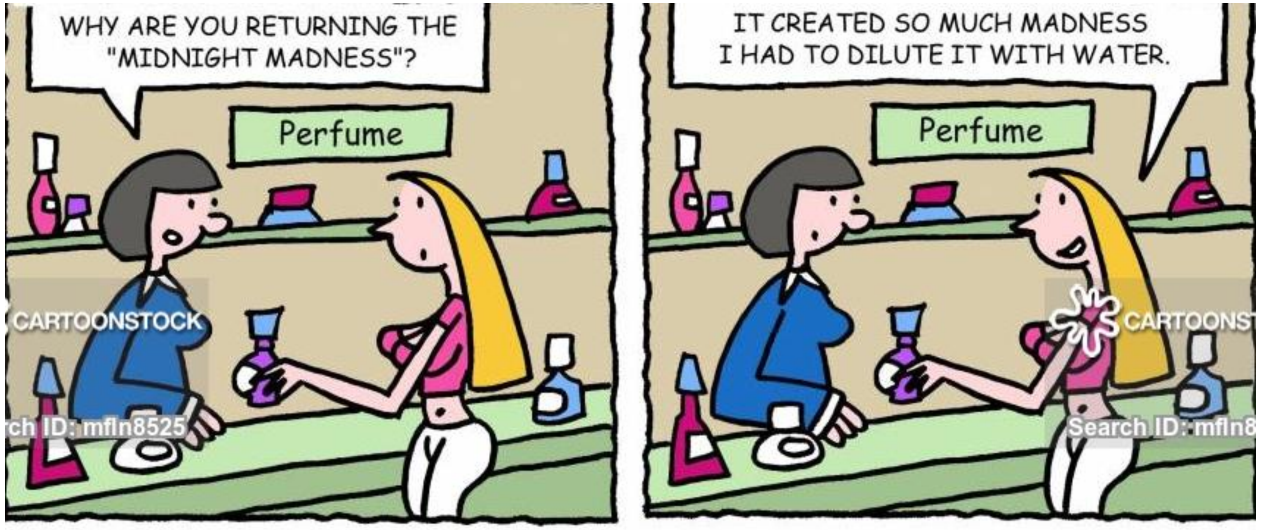
$$c_{original} = c_{diluted} \times \frac{V_{diluted}}{V_{original}} \text{ (dilution factor)}$$

MIXING

Mixing of two separate solutions can also result in dilution of the chemical species present.

PORTIONING

Portioning occurs when a **small portion of solution** is taken from a **stock solution**. The **concentration of the solute remains unchanged**.



Titrations (Volumetric Analysis)

STANDARD SOLUTION

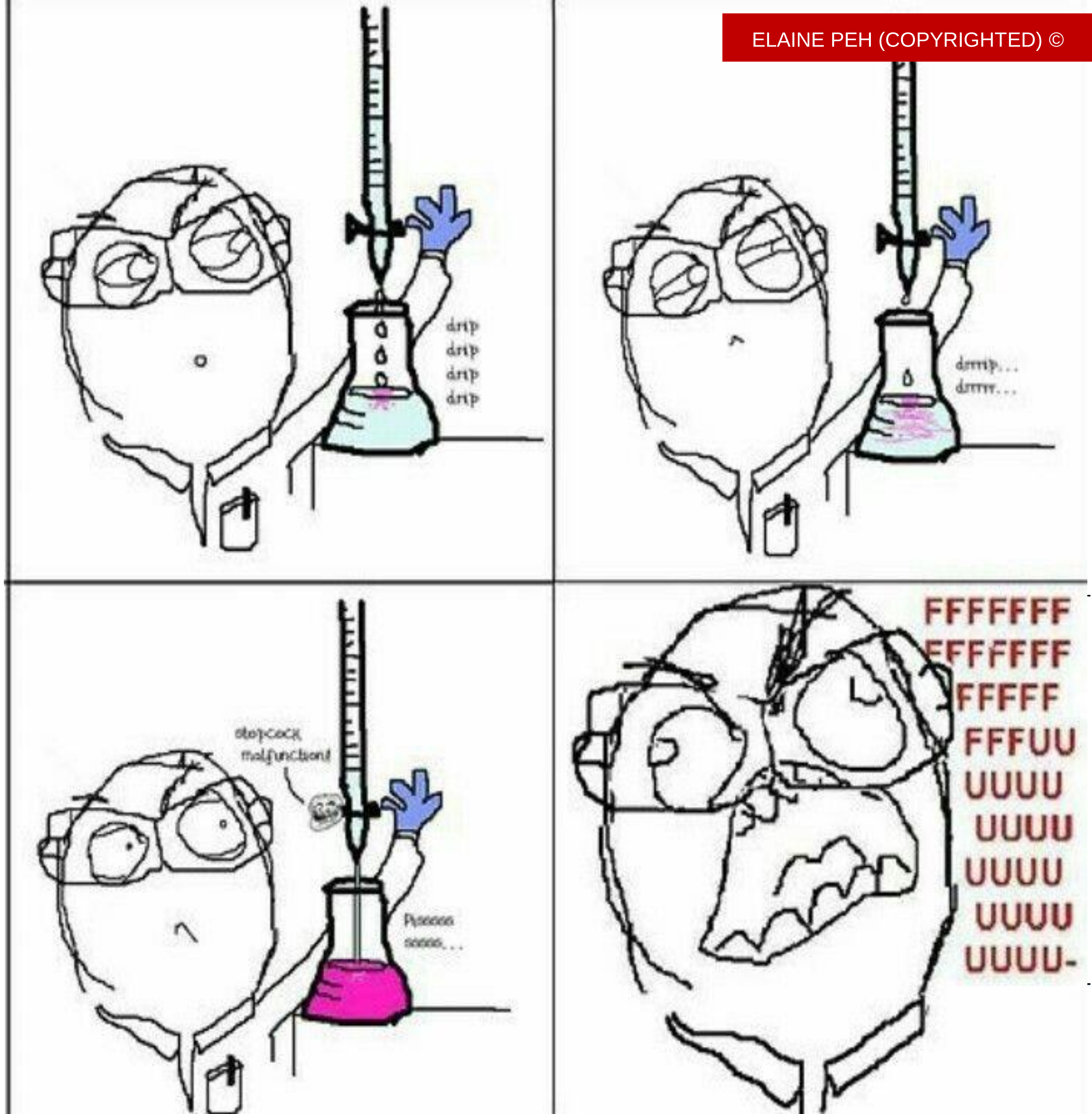
A solution of **known concentration**.

EQUIVALENCE POINT

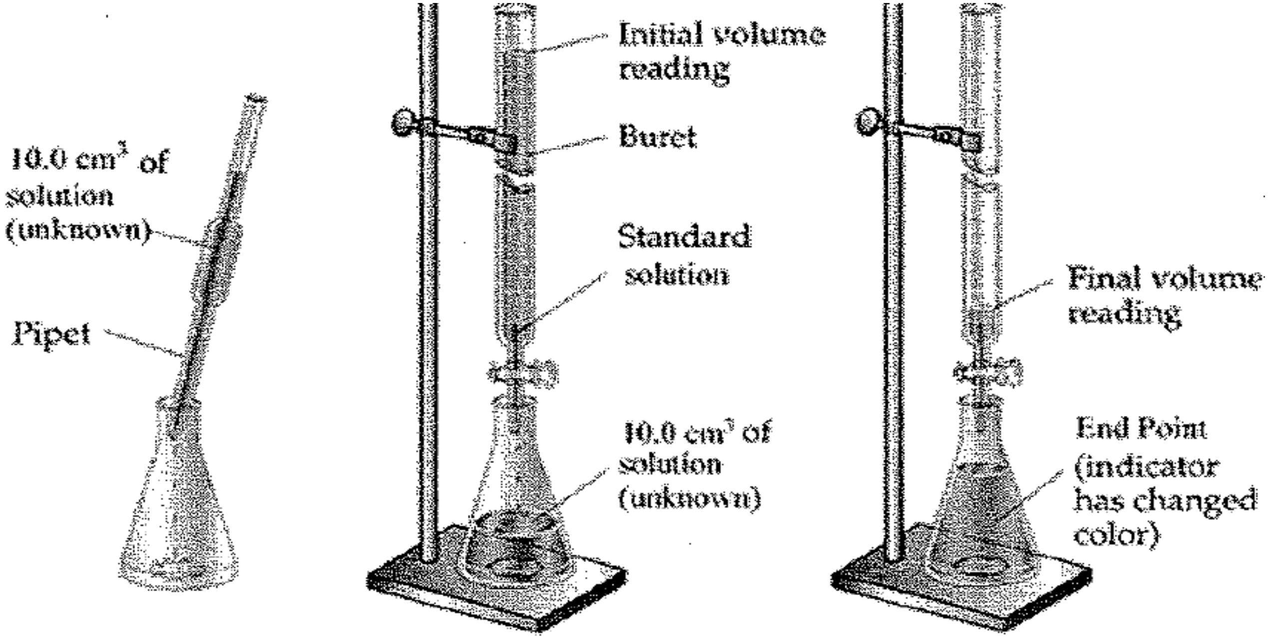
It **marks the end of the titration** when **reactants have just reacted with each other** according to the stoichiometric ratio. *It is the **theoretical value**.*

END POINT

It **marks the end of the titration** when the indicator in the titration has just changed its colour. *It is the **experimental value**.*



Titrations (Volumetric Analysis)

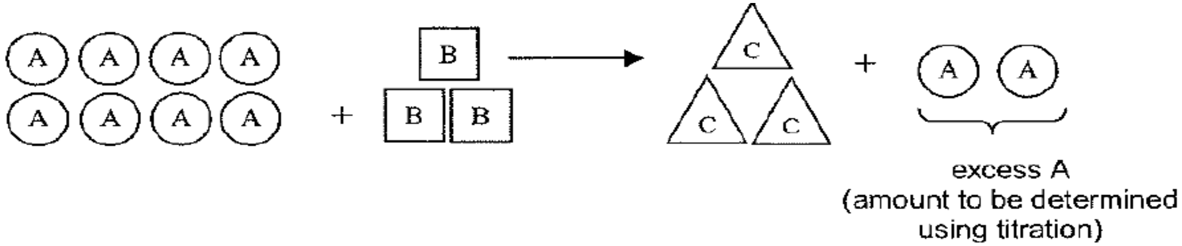


TITRATION

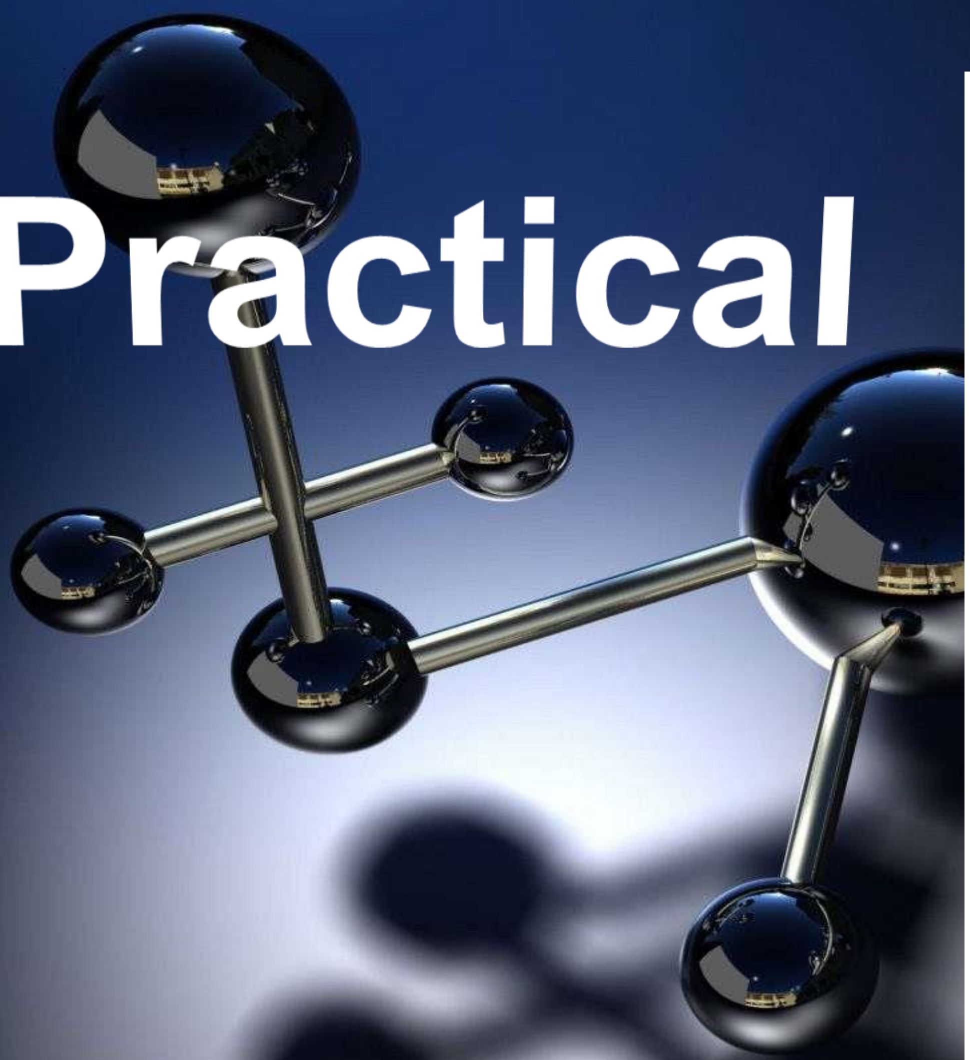
Titration is a type of **volumetric analysis** where a **solution of known concentration** (commonly refer to as standard solution) is **added to another** to **determine the unknown concentration** of that solution.

BACK TITRATION

It is an indirect method that first calculate the amount of excess (unreacted) reagent **(A)** left after the original reaction by titration with another standard solution.



Practical

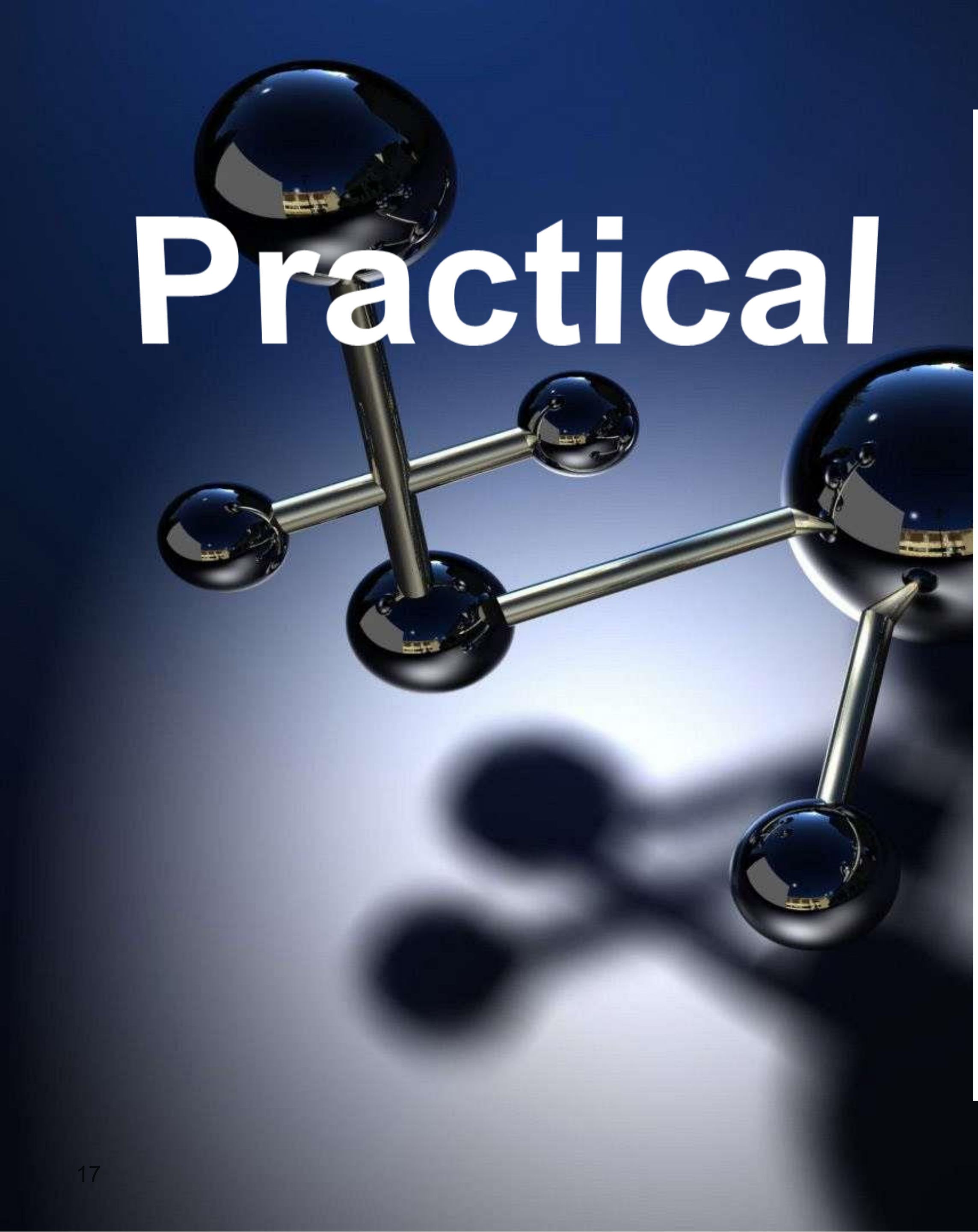


GENERAL LEARNING POINTS FOR TITRATION

1. Washing - with distilled water for conical flask, and ONLY solution of interest for burette and pipette and FUNNEL!
2. Perform the titration on a white tile for better colour change
3. Tabulation table of results (write in volumes in PENCIL separately first - do 3 titrations for approximate result):

Final burette reading / cm^3	25.20	35.20
Initial burette reading / cm^3	0.00	10.00
Volume of FA1 (SPECIFY) used / cm^3	25.20	25.20
Selected readings (✓)	✓	✓

Selecting readings: as long as readings are within $0.10cm^3$ of each other, select TWO readings and calculate average reading using $\frac{25.20 + 25.20}{2} = 25.20\text{cm}^3$



REDOX TITRATION

Using known concentration of thiosulfate (titre) to determine no. of moles of I_2 liberated from oxidising agent mixed with I^-

Common questions:

- *Explain the colour changes in the titration of thiosulfate with I_2 produced: Solution first turns from deep purple to reddish brown as the oxidising agent is reduced with I_2 is produced; as more thiosulfate is added, the brown I_2 solution turns yellow as thiosulfate reduces I_2 to I^- (colourless) and there is less I_2 in the mixture. When starch is added, it forms a blue-black complex with I_2; eventually turns colourless as all the I_2 is reduced by thiosulfate*
- *Suggest why the I_2 produced must be immediately titrated with thiosulfate: I_2 is a volatile compound and easily vaporizer out of the conical flask, resulting in decreased no. of moles of I_2 present*
- *Explain why it is not necessary to measure the volume of sulfuric acid accurately using a pipette: Sulfuric acid in the experiment is used to provide an acidic medium and does not undergo a REDOX reaction / sulfuric acid in excess*
- *Reasons why excess acid in solution must be neutralise: Sodium thiosulfate reacts with acid to form sulfur and other products hence titre volume of thiosulfate required would increase to more than expected*

Practice Questions

Question: Lead occurs as isotope in nature, determine the relative atomic mass of lead given the following data:

Relative isotopic mass	% abundance
204	1.41
206	24.97
207	21.59
208	52.03

Answer: 207.2

Question: Determine the empirical formula of a sugar that contain 40% by mass of carbon, 67% by mass of hydrogen, and 53.5% by mass of oxygen.

Answer: CH_2O

Question: Determine the empirical formula of an oxide of iron that contains 70.0% Fe by mass.

Answer: Fe_2O_3

Practice Questions

Question: Determine the molecular formula for ascorbic acid that contains 40.92% carbon, 4.58% hydrogen and 54.5% oxygen by mass with a determined molar mass of 176.0 g mol^{-1}.

Answer: *$C_6H_8O_6$*

Question: Complete combustion of an unknown organic compound of mass 1 g containing only carbon, hydrogen and oxygen gave 1.5 g of carbon dioxide and 0.405 g of water. Determine the empirical formula of the organic compound.

Answer: $C_3H_4O_3$

Question: When 10 cm^3 of a gaseous hydrocarbon was completely burned in excess oxygen, 20 cm^3 of carbon dioxide and 30 cm^3 of steam was product. Determine the molecular formula of the hydrocarbon.

Answer: C_2H_6

Practice Questions

Question: Determine the limiting reagent when 3 mol of Mg are added to 4 mol of HCl

Answer: HCl

Question: Determine the percentage yield of barium sulfate when 15 g of barium chloride was added to 10 g of iron (III) sulfate with a precipitation of 15.6 g of barium sulfate obtained.

Answer: 92.8%

Question: Calculate the concentration in both g dm^{-3} and mol dm^{-3} when 6.5 g of $CaCl_2$ is dissolved in water and the solution is made up to 500 cm^3.

Answer: 13.0 g dm^{-3}; 0.1117 mol dm^{-3}

Question: 10 cm^3 of a NaOH solution of concentration 1.5 mol dm^{-3} was diluted to 250.0 cm^3 using distilled water. Calculate the concentration of the NaOH solution after dilution.

Answer: 0.0600 mol dm^{-3}

Question: A concentrated solution of H_2SO_4 has a concentration of 3.35 mol dm^{-3}. What volume of this acid is required to prepare 250.0 cm^3 of 0.130 mol dm^{-3} H_2SO_4 solution.

Answer: 9.70 cm^3

Practice Questions

Question: A student dissolved 0.500 mol NaOH in 250 cm^3 of water and pipetted 20.0 cm^3 of the solution into a conical flask. What is the amount of NaOH in the flask?

Answer: 0.0400 mol

Question: 20.0 cm^3 of a solution of barium hydroxide of unknown concentration was placed in a conical flask and titrated with a standard solution of hydrochloric acid which had a concentration of 0.0600 mol dm^{-3}. The volume of acd required was 25.0 cm^3. Calculate the molar concentration of barium hydroxide.

Answer: 0.0375 mol dm^{-3}

Question: A sample of calcium carbonate was added to 50.0 cm^3 of 1.00 mol dm^{-3} of hydrochloric acid. The resulting solution was made up to 250.0 cm^3 with deionized water. When 25.0 cm^3 of this solution was titrated with 0.100 mol dm^{-3} of sodium hydroxide solution, 25.0 cm^3 of sodium hydroxide solution was needed for complete reaction. Calculate the mass of calcium carbonate.

Answer: 1.25 g

Question: A 4.30 g of an impure sample of NH_4Cl was warmed with 100.0 cm^3 of 1.00 mol dm^{-3} of NaOH. The mixture was boiled until all the ammonia was driven off. The excess NaOH in the mixture required 25.0 cm^3 of 0.500 mol dm^{-3} of sulfuric acid for neutralization. Calculate the percentage purity of the NH_4Cl.
Answer: 93.3%

Test yourself!

(a) define the terms *relative atomic*, *isotopic*, *molecular* and *formula mass*

(b) define the term *mole* in terms of the Avogadro constant

(c) calculate the relative atomic mass of an element given the relative abundances of its isotopes

(d) define the terms *empirical* and *molecular formula*

(e) calculate empirical and molecular formulae using combustion data or composition by mass

Test yourself!

(f) write and/or construct balanced equations

(g) perform calculations, including use of the mole concept, involving:

i. reacting masses (from formulae and equations)
ii. volumes of gases (e.g. in the burning of hydrocarbons)
iii. volumes and concentrations of solutions

[when performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question]

(h) deduce stoichiometric relationships from calculations such as those in (g)