Moles

A mole (abbreviated **mol**) **is a unit of measure in chemistry**.

One (1) mole of any element = it's Relative Atomic Mass (R.A.M. or A_r)

e.g. 1 mol Cl = 35.5 g
1 mol N = 14 g

Avogadro's number (L) is the <u>number of particles (atoms, molecules or ions)</u> in one (1) mole of an element or compound.

1 mole = L = 6.023×10^{23} particles (atoms, molecules or ions)

e.g. **(1)** How many atoms are there in 0.5 moles of sodium?
1 mol = L
1 mol Na = 6.023×10^{23} atoms

0.5 mols Na atoms = $0.5 \times (6.023 \times 10^{23})$ atoms = 3.01×10^{23} atoms

To calculate **number of moles of an element**:

$$\text{No. of mols} = \frac{\text{mass}}{A_r}$$

e.g. **(2)** How many moles are there in 57.5 g of sodium?

$$\text{No. of mols} = \frac{57.5}{23} = 2.5 \text{ mols Na}$$

To calculate **mass of an element**:

Mass = no. of mols x A_r

e.g. **(3)** What is the mass of 5 moles of sodium?
mass = 5 x 23 = 115 g Na

To calculate the **Relative Molecular Mass** (R.M.M. or M_r) of a compound:

e.g. **(4)** M_r of H_2SO_4 = (2 x 1) + (1 x 32) + (4 x 16)

= 98 g

To calculate **number of moles of a compound**:

$$\text{No. of mols} = \frac{\text{mass}}{M_r}$$

e.g. **(5)** How many moles are there in 1.17 g of sodium chloride?

M_r of NaCl = 23 + 35.5 = 58.5 g

$$\text{No. of mols} = \frac{1.17}{58.5} = 0.02 \text{ mols NaCl}$$

To calculate **mass of a compound**:

Mass = no. of mols x M_r

e.g. **(6)** What is the mass of 0.25 moles of sulfuric acid (H_2SO_4)?

Mass = 0.25 x 98 = 24.5 g H_2SO_4 [M_r of H_2SO_4 = 98 g]

*Note that if the **volume** of a compound is given, then using:

Density = mass/ volume

The mass of the compound can be calculated first, and then the no. of mols.

Moles applied to gases

A gas takes up a certain **volume** in a container.

One (1) mole of gas occupies approximately 22.4 dm^3 (or 22400 cm^3) at standard temperature and pressure (s.t.p.). S.t.p. is 273K (0^{o}C) and 1 atmosphere (atm.)

One (1) mole of gas occupies approximately 24 dm^3 (or 24000 cm^3) at room temperature and pressure (r.t.p.). R.t.p. is 298K (25^{o}C) and 1 atm.

[1 dm^3 = 1000 cm^3]

e.g. **(7)** What volume is occupied by 8 g of hydrogen (H_2) at s.t.p?

First, find no. of mols:

M_r of H_2 = 2 g

Mols. of H_2 = $\frac{\text{mass}}{M_r} = \frac{8}{2}$ = 4 mols H_2

Then, find volume:

1 mol of H_2 occupies 22.4 dm^3 at s.t.p.

Therefore 4 mols of H_2 occupies (4 x 22.4) dm^3 = 89.6 dm^3

e.g. **(8)** How many moles of CO_2 are present in 250 cm^3 of gas at r.t.p.?

1 mol CO_2 occupies 24 dm^3

1 mol CO_2 occupies 24000 cm^3

If 24000 cm^3 contains 1 mol CO_2

1 cm^3 contains $\frac{1}{24000}$ mol CO_2

250 cm^3 contains $\left(\frac{1}{24000} \times 250\right)$ mols CO_2 = 0.01 mols CO_2

Moles and Concentration

<u>The amount of solute dissolved in a given quantity of solution is known as concentration.</u>

Concentration is expressed in terms of:

The <u>number of moles</u> of solute in 1000 cm^3 (or 1 dm^3) of solution. This is called **molar concentration**. The **units** of molar concentration are: **mol dm^{-3} or mol L^{-1}**, abbreviated **M**.

E.g. If the concentration of a solution of NaCl is 1.0 mol dm^{-3}, then this means that there is 1 mole of NaCl dissolved in 1 dm^3 or 1000 cm^3 of <u>solution</u>.

Concentration is also expressed in terms of:

The <u>mass</u> of solute in 1000 cm^3 (or 1 dm^3) of solution. This is called **mass concentration**. The **units** of mass concentration are: **gdm^{-3} or gL^{-1}**.

e.g. If the concentration of a solution of NaCl is 1.0 gdm^{-3}, then this means that there is 1 gram of NaCl dissolved in 1 dm^3 or 1000 cm^3 of <u>solution</u>.

To calculate **molar concentration**:

Find the number of moles of solute present in 1 dm^3 of solution.

e.g. **(9)** Find the molar concentration of a solution of HCl if 25 cm^3 of solution contains 0.015 moles of HCl.

25 cm^3 contain 0.015 mol HCl

1 cm^3 contains $\frac{0.015}{25}$ mol HCl

1000 cm^3 contain $\frac{0.015 \times 1000}{25}$ mol = 0.6 mol HCl

Therefore, the molar conc. of the HCl solution is 0.6 mol dm^{-3}.

To calculate **number of moles of solute** in a volume of solution using molar concentration:

e.g. **(10)** Calculate the number of moles in 25 cm^3 of a 0.05 mol dm^{-3} H_2SO_4 solution.

1000 cm^3 contain 0.05 mol H_2SO_4

1 cm^3 contains $\frac{0.05}{1000}$ mol H_2SO_4

25 cm^3 contain $\frac{0.05 \times 25}{1000}$ mol = 0.00125 mol H_2SO_4

To calculate **mass concentration**:

Find the mass of solute present in 1 dm^3 of solution.

e.g. **(11)** Find the mass concentration of a solution of NaCl if 25 cm^3 of the solution contains 0.5 g of NaCl.

25 cm^3 contain 0.5 g NaCl

1 cm^3 contains $\frac{0.5}{25}$ g NaCl

1000 cm^3 contain $\frac{0.5 \times 1000}{25}$ g = 20 g NaCl

Therefore, the mass conc. of the NaCl solution is 20 gdm^{-3}.

To calculate **mass of solute** in a volume of solution using concentration:

e.g. **(12)** Calculate the mass of NaCl present in 45 cm^3 of a 5.0 g dm^{-3} NaCl solution.

1000 cm^3 contain 5.0 g NaCl

1 cm^3 contains $\frac{5.0}{1000}$ g NaCl

45 cm^3 contain $\frac{5.0 \times 45}{1000}$ mol = 0.225 g NaCl

To **convert molar concentration to mass concentration**:
Multiply molar concentration by the $\mathbf{M_r}$ of the solute.
e.g. **(13)** What is the mass concentration of a 2 mol dm^{-3} NaCl solution?

M_r of NaCl = 58.5 g
Mass of 2 moles of NaCl = 2 x 58.5 = 117 g

Therefore, the mass concentration of NaCl is 117 gdm^{-3}.

To **convert mass concentration to molar concentration**:
Divide the mass concentration by the $\mathbf{M_r}$ of the solute.
e.g. **(14)** What is the molar concentration of a 2.52 gdm^{-3} solution of HNO_3?

M_r of HNO_3 = 63 g

No. of moles of HNO_3 in 2.52 g = $\frac{\text{mass}}{M_r} = \frac{2.52}{63}$ = 0.04 mol HNO_3

Therefore, the molar concentration of HNO_3 is 0.04 mol dm^{-3}.

N.B. Concentration is expressed as the amount of solute in 1 dm^3 of solution NOT solvent.

When asked to find concentration it is usually molar concentration unless otherwise indicated.

Concentration can also be expressed in terms of:

- Percentage by mass or percentage mass:

  % mass = mass of solute/ mass of solution x 100

- **% w/w** or **% m/m** – mass of solute in 100 g of solution
- **% w/v** or **% m/v** – mass of solute in 100 mL of solution
- **% v/v** – volume of solute in 100 mL of solution

e.g. A 10% w/v KI solution means that there are 10 g of KI in 100 mL of KI solution.

Preparation of Standard Solutions.

A solution of a particular/ known concentration is called a __standard solution__.

Standard solutions are made up in **two ways**:

1) By **calculating the amount of solute required** to make up the solution and dissolving this amount in the solvent or
2) By **using a solution of a different (higher) concentration and diluting it to get the standard solution**. The more concentrated solution is called the **stock solution**.

However, only a portion of the stock is used therefore, calculations need to be done in order to find out exactly <u>how much of the stock solution is required to make up the standard solution</u>.

1) Calculating the **amount of solute** required:

e.g. **(15)** What mass of Na_2CO_3 is needed to make 500 cm^3 of a 0.5 mol dm^{-3} solution?

First **calculate the number of moles** in the standard solution:

$$\text{No. of moles of } Na_2CO_3 = \frac{\textbf{vol. x conc.}}{\textbf{1000}} = \frac{500 \times 0.5}{1000} = 0.25 \text{ mol } Na_2CO_3$$

Next, **convert this number of moles to mass**:

Mass of Na_2CO_3 = mol x M_r = 0.25 x 106 = 26.5 g

Therefore, in order to prepare the 0.5 mol dm^{-3} solution of Na_2CO_3, 26.5 g of solid Na_2CO_3 must be weighed out and dissolved in enough solvent to make up 500 cm^3 of solution.

* Note that standard solutions are always prepared in a **<u>volumetric flask</u>** because it must be done <u>accurately</u>.

The solid solute is always weighed out accurately as well (using an analytical balance) and the mass is usually recorded to 4 decimal places (d.p.).

2) Calculating the **volume of stock solution required for dilution**:

e.g. **(16)** How can 200 cm^3 of a 0.1 mol dm^{-3} solution of NaOH be prepared using an NaOH solution of concentration 5.0 mol dm^{-3}?

First calculate the **number of moles of solute** in the **diluted (standard) solution**:

No. of moles of NaOH in diluted solution = $\dfrac{\textbf{vol. x conc.}}{\textbf{1000}} = \dfrac{200 \times 0.1}{1000}$ = 0.02 mol NaOH

Then calculate the **volume of the concentrated (stock) solution** that contains this number of moles:

1000 cm^3 of stock contains 5.0 mol NaOH

$\dfrac{1000}{5.0}$ cm^3 of stock NaOH contains 1 mol NaOH

$\dfrac{1000 \times 0.02}{5.0}$ cm^3 of stock contains 0.02 mol NaOH = 4 cm^3 of stock solution

Alternatively:

Use $\mathbf{C_1V_1 = C_2V_2}$:

C_1 = 0.1 mol dm^{-3} C_2 = 5.0 mol dm^{-3}
V_1 = 200 cm^3 V_2 = ?
(0.1 mol dm^{-3})(200 cm^3) = (5.0 mol dm^{-3})V_2
V_2 = 4 cm^3

Therefore, in order to make up the standard solution of 0.1 mol dm^{-3} NaOH, use 4 cm^3 of the 5.0 mol dm^{-3} solution and add enough solvent to make up 200 cm^3 of solution.

* Note that a **pipette** is used to transfer the required portion of stock solution to the volumetric flask because the volume of stock solution must be transferred accurately.

Empirical and molecular formulae

Empirical formula – shows the simplest whole no. ratio of atoms in the compound

Molecular formula – shows the actual no. of atoms in the compound

e.g. **(17)** The hydrocarbon benzene contains 92.3% by mass of carbon. The mass of 1 mole of benzene is 78 g. Find the empirical and molecular formulae of benzene.

	C	H
Convert % to mass	92.3 g	7.7 g
Convert mass to moles	$\frac{92.3}{12}$ = 7.69 mols.	$\frac{7.7}{1}$ = 7.7 mols.
Divide through by the smallest no.	$\frac{7.69}{7.69}$ = 1	$\frac{7.7}{7.69}$ = 1
Mole ratio	1 : 1	

Empirical formula = CH

Mass of 1 formula unit = 1(12) + 1(1) = 13 g

No. of formula units present in 78 g = 78/ 13 = 6 units
Molecular formula of benzene = 6(CH) = C_6H_6

N.B.: In preparation for the following sections, please review:

- Writing and balancing equations, including state symbols
- Solubility of compounds: nitrates, sulfates, chlorides, carbonates, hydroxides (to be able to write correct state symbols)
- Writing different types of reactions

E.g. Neutralization reactions – acid + base $\longrightarrow$ salt + water

– acid + carbonate $\longrightarrow$ salt + water + carbon dioxide

- Limiting and excess reagents; percentage yield

e.g. **(18)** Calcium carbide, CaC_2 reacts with water to form calcium hydroxide and the flammable gas ethyne (acetylene), according to the equation:

$$CaC_2\,(s) + 2H_2O\,(l) \longrightarrow Ca(OH)_2\,(aq) + C_2H_2\,(g)$$

(i) Which is the limiting reagent when 50 g of water reacts with 50 g of calcium carbide?

(ii) What mass of excess reagent remains after the reaction is complete?

(iii) Given that 20 g of ethyne were produced, calculate the percentage yield of ethyne.

(i) Mols. of CaC_2 = 50/ 64 = 0.781 mols. [M_r of CaC_2 = 64 g]

Mols. of H_2O = 50/ 18 = 2.78 mols. [M_r of H_2O = 18 g]

$CaC_2 : H_2O$

1 : 2

0.781 mols.: 1.562 mols. [Using the mols. of CaC_2]

1.562 mols. of H_2O would be required to react with all of the CaC_2

Since 2.78 mols. of H_2O were present in the rxn., H_2O is in excess.

Therefore, CaC_2 is the limiting reagent.

(ii) Mols. of H_2O remaining = 2.78 – 1.562 = 1.218 mols.

Mass of H_2O remaining = 1.218 x 18 = 21.92 g

(iii) $CaC_2 : C_2H_2$

1 : 1

0.781 mols.: 0.781 mols.

Mols. of C_2H_2 produced = 0.781 mols.

M_r of C_2H_2 = 26 g

Mass of C_2H_2 produced = 0.781 x 26 = 20.31 g

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{20}{20.31} \times 100 = 98\%$$

e.g. **(19)** Suppose that 50.0 mL each of 0.150 M $AgNO_3$ and 0.200 M Na_2CrO_4 solutions are mixed. The chemical equation for the reaction is:

$$2AgNO_3\ (aq) + Na_2CrO_4\ (aq) \longrightarrow 2NaNO_3\ (aq) + Ag_2CrO_4\ (s)$$

(i) Which of the two reactants determines the mass of Ag_2CrO_4 which precipitates out of solution?

(ii) What mass of Ag_2CrO_4 is formed in this reaction?

(iii) What mass of excess reactant remains in solution?

(i) Mols. of $AgNO_3$ = 50.0 x 0.150/ 1000 = 0.0075 mols.

Mols. of Na_2CrO_4 = 50.0 x 0.200/ 1000 = 0.01 mols.

$AgNO_3$: Na_2CrO_4

2 : 1

0.0075 mols.: 0.00375 mols. [Using the mols. of $AgNO_3$]

0.00375 mols. of Na_2CrO_4 would be required to react with all of the $AgNO_3$

Since 0.01 mols. of Na_2CrO_4 were present in the rxn., Na_2CrO_4 is in excess.

Therefore, $AgNO_3$ is the limiting reagent and determines the mass of Ag_2CrO_4 that precipitates out of solution.

(ii) $AgNO_3$: Ag_2CrO_4

2 : 1

0.0075 mols.: 0.00375 mols.

Mols. of Ag_2CrO_4 produced = 0.00375 mols.

M_r of Ag_2CrO_4 = 332 g

Mass of Ag_2CrO_4 produced = 0.00375 x 332 = 1.245 g

(iii) Mols. of Na_2CrO_4 remaining = 0.01 – 0.00375 = 0.00625 mols.

M_r of Na_2CrO_4 = 162 g

Mass of Na_2CrO_4 remaining = 0.00625 x 162 = 1.01 g

Titrations

A titration is a technique usually used to find the concentration of a particular solution.

For example, an **acid-base titration** can be used to find the concentration of either an acid or base. It is a neutralization reaction that generally produces a salt and water.

e.g. **(20)** 25 mL of a 0.1 M NaOH solution required 23.5 mL of HCl for complete reaction. Calculate the concentration of the HCl solution.

Mols. of NaOH reacting in the titration = 25 x 0.1/ 1000 = 0.0025 mols.

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

NaOH : HCl

1 : 1

0.0025 mols. : 0.0025 mols.

Therefore, the mols. of HCl that reacted in the titration = 0.0025 mols.

23.5 mL of solution contains 0.0025 mols. HCl

1000 mL of solution contains 0.0025/ 23.5 x 1000 mols. HCl = 0.106 mols. Conc. of the HCl solution = 0.106 M

Back-titrations

When a reaction during a titration is incomplete, direct titration becomes impossible. In such cases, a back titration is used. Instead of titrating the original sample, a known excess of a standard reagent is added to the solution, and the excess reagent titrated. A back titration is useful if the endpoint of the reverse titration is easier to identify than the endpoint of the normal titration.

e.g. **(21)** A scientist wishes to determine the percentage composition of sodium ethanoate (CH_3COONa) in a powder. Since sodium ethanoate is the salt of NaOH and CH_3COOH, she performs the following steps:

(1) 10.0 g of powder are dissolved and made up to the mark with distilled water in a 250 mL volumetric flask, (2) 25 mL of this solution is reacted completely with 50 mL of 0.1 M H_2SO_4, (3) this reaction mixture is then titrated against 0.2 M NaOH which yields an average titre volume of 13.35 mL.

(i) How many moles of excess H_2SO_4 remain after Step (2)?

(ii) How many moles of H_2SO_4 were consumed in Step (2)?

(iii) What mass of CH_3COONa is present in 10.0 g of powder?

(i) Mols. of NaOH reacting in Step (3) = 13.35 x 0.2/ 1000 = 0.00267 mols.

$$2NaOH\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2H_2O\ (l)$$

$NaOH : H_2SO_4$

2 : 1

0.00267 mols. : 0.001335 mols.

Therefore, the mols. of H_2SO_4 that reacted in Step (3) = 0.001335 mols.

So the mols. of H_2SO_4 that remain after Step (2) (i.e. the excess) = 0.001335 mols.

(ii) Mols. of H_2SO_4 added initially in Step (2) = 50 x 0.1/ 1000 = 0.005 mols.

Mols. of H_2SO_4 consumed in Step (2) = 0.005 – 0.001335 = 0.003665 mols.

(iii) $$2CH_3COONa\ (aq) + H_2SO_4\ (aq) \longrightarrow 2CH_3COOH\ (aq) + Na_2SO_4\ (aq)$$

$CH_3COONa : H_2SO_4$

2 : 1

0.00733 mols. : 0.003665 mols.

Therefore, the mols. of CH_3COONa that reacted in Step (2) = 0.00733 mols.

25 mL of solution contains 0.00733 mols. CH_3COONa

250 mL of solution contains 0.00733/ 25 x 250 mols. CH_3COONa = 0.0733 mols.
M_r of CH_3COONa = 82 g
Mass of CH_3COONa in 10 g of powder = 0.0733 x 82 = 6 g