
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 number of particles (atoms, molecules or ions) 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:**

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

A_r

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

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

To calculate **mass of an element:**

Mass = no. of mols $\times A_r$

e.g. (3) What is the mass of 5 moles of sodium?

mass = $5 \times 23 = 115$ g Na

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

e.g. (4) M_r of $\text{H}_2\text{SO}_4 = (2 \times 1) + (1 \times 32) + (4 \times 16)$
 $= 98 \text{ g}$

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?

$$M_r \text{ of NaCl} = 23 + 35.5 = 58.5 \text{ g}$$

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

To calculate **mass of a compound**:

$$\text{Mass} = \text{no. of mols} \times M_r$$

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

$$\text{Mass} = 0.25 \times 98 = 24.5 \text{ g H}_2\text{SO}_4$$

$$[M_r \text{ of H}_2\text{SO}_4 = 98 \text{ g}]$$

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

$$\text{Density} = \text{mass} / \text{volume}$$

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 dm³ (or 22400 cm³) at standard temperature and pressure (s.t.p.). S.t.p. is 273K (0°C) and 1 atmosphere (atm.)

One (1) mole of gas occupies approximately 24 dm³ (or 24000 cm³) at room temperature and pressure (r.t.p.). R.t.p. is 298K (25°C) and 1 atm.

[1 dm³ = 1000 cm³]

e.g. (7) What volume is occupied by 8 g of hydrogen (H₂) at s.t.p?

First, find no. of mols:

M_r of H₂ = 2 g

$$\text{Mols. of H}_2 = \frac{\text{mass}}{M_r} = \frac{8}{2} = 4 \text{ mols H}_2$$

Then, find volume:

1 mol of H₂ occupies 22.4 dm³ at s.t.p.

Therefore 4 mols of H₂ occupies (4 x 22.4) dm³ = 89.6 dm³

e.g. (8) How many moles of CO₂ are present in 250 cm³ of gas at r.t.p.?

1 mol CO₂ occupies 24 dm³

1 mol CO₂ occupies 24000 cm³

If 24000 cm³ contains 1 mol CO₂

1 cm³ contains $\frac{1}{24000}$ mol CO₂

$$250 \text{ cm}^3 \text{ contains } \left[\frac{1}{24000} \times 250 \right] \text{ mols CO}_2 = 0.01 \text{ mols CO}_2$$

Moles and Concentration

The amount of solute dissolved in a given quantity of solution is known as concentration.

Concentration is expressed in terms of:

The **number of moles of solute in 1000 cm³ (or 1 dm³) 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 dm³ or 1000 cm³ of solution.

Concentration is also expressed in terms of:

The **mass of solute in 1000 cm³ (or 1 dm³) 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 dm³ or 1000 cm³ of solution.

To calculate **molar concentration**:

Find the number of moles of solute present in 1 dm³ of solution.

e.g. (9) Find the molar concentration of a solution of HCl if 25 cm³ of solution contains 0.015 moles of HCl.

25 cm³ contain 0.015 mol HCl

1 cm³ contains $\frac{0.015}{25}$ mol HCl

1000 cm³ 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 cm³ of a 0.05 mol dm⁻³ H₂SO₄ solution.

$$1000 \text{ cm}^3 \text{ contain } 0.05 \text{ mol H}_2\text{SO}_4$$

$$1 \text{ cm}^3 \text{ contains } \frac{0.05}{1000} \text{ mol H}_2\text{SO}_4$$

$$25 \text{ cm}^3 \text{ contain } \frac{0.05 \times 25}{1000} \text{ mol} = 0.00125 \text{ mol H}_2\text{SO}_4$$

To calculate **mass concentration**:

Find the mass of solute present in 1 dm³ of solution.

e.g. (11) Find the mass concentration of a solution of NaCl if 25 cm³ of the solution contains 0.5 g of NaCl.

$$25 \text{ cm}^3 \text{ contain } 0.5 \text{ g NaCl}$$

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

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

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 cm³ of a 5.0 g dm⁻³ NaCl solution.

$$1000 \text{ cm}^3 \text{ contain } 5.0 \text{ g NaCl}$$

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

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

To **convert molar concentration to mass concentration**:

Multiply molar concentration by the M_r of the solute.

e.g. (13) What is the mass concentration of a 2 mol dm^{-3} NaCl solution?

$$M_r \text{ of NaCl} = 58.5 \text{ g}$$

$$\text{Mass of 2 moles of NaCl} = 2 \times 58.5 = 117 \text{ g}$$

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

To **convert mass concentration to molar concentration**:

Divide the mass concentration by the M_r of the solute.

e.g. (14) What is the molar concentration of a 2.52 gdm^{-3} solution of HNO_3 ?

$$M_r \text{ of HNO}_3 = 63 \text{ g}$$

$$\text{No. of moles of HNO}_3 \text{ in } 2.52 \text{ g} = \frac{\text{mass}}{M_r} = \frac{2.52}{63} = 0.04 \text{ mol HNO}_3$$

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:

$$\% \text{ mass} = \text{mass of solute} / \text{mass of solution} \times 100$$

- **% 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 how much of the stock solution is required to make up the standard solution.

- 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:

$$\text{No. of moles of Na}_2\text{CO}_3 = \frac{\text{vol.} \times \text{conc.}}{1000} = \frac{500 \times 0.5}{1000} = 0.25 \text{ mol Na}_2\text{CO}_3$$

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

$$\text{Mass of Na}_2\text{CO}_3 = \text{mol} \times M_r = 0.25 \times 106 = 26.5 \text{ g}$$

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 **volumetric flask** because it must be done accurately.

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 cm³ 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:**

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

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

1000 cm³ of stock contains 5.0 mol NaOH

$\frac{1000}{5.0}$ cm³ of stock NaOH contains 1 mol NaOH

$\frac{1000 \times 0.02}{5.0}$ cm³ of stock contains 0.02 mol NaOH = 4 cm³ of stock solution

Alternatively:

Use $C_1V_1 = C_2V_2$:

$$C_1 = 0.1 \text{ mol dm}^{-3} \quad C_2 = 5.0 \text{ mol dm}^{-3}$$

$$V_1 = 200 \text{ cm}^3 \quad V_2 = ?$$

$$(0.1 \text{ mol dm}^{-3})(200 \text{ cm}^3) = (5.0 \text{ mol dm}^{-3})V_2$$

$$V_2 = 4 \text{ cm}^3$$

Therefore, in order to make up the standard solution of 0.1 mol dm⁻³ NaOH, use 4 cm³ of the 5.0 mol dm⁻³ solution and add enough solvent to make up 200 cm³ 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 = $6(\text{CH}) = \text{C}_6\text{H}_6$

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

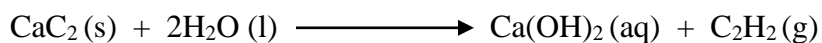
- Writing and balancing equations, including state symbols
- Solubility of compounds: nitrates, sulfates, chlorides, carbonates, hydroxides (to be able to write correct state symbols)
- Writing different types of reactions

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

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

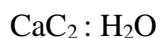
➤ Limiting and excess reagents; percentage yield

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



- (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 $\text{CaC}_2 = 50 / 64 = 0.781$ mols. [M_r of $\text{CaC}_2 = 64$ g]
Mols. of $\text{H}_2\text{O} = 50 / 18 = 2.78$ mols. [M_r of $\text{H}_2\text{O} = 18$ g]



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 = $1.218 \times 18 = 21.92$ g

- (iii) $\text{CaC}_2 : \text{C}_2\text{H}_2$

1 : 1

0.781 mols. : 0.781 mols.

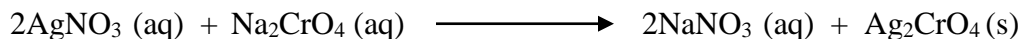
Mols. of C_2H_2 produced = 0.781 mols.

M_r of $\text{C}_2\text{H}_2 = 26$ g

Mass of C_2H_2 produced = $0.781 \times 26 = 20.31$ g

% 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 AgNO₃ and 0.200 M Na₂CrO₄ solutions are mixed. The chemical equation for the reaction is:



- (i) Which of the two reactants determines the mass of Ag₂CrO₄ which precipitates out of solution?
- (ii) What mass of Ag₂CrO₄ is formed in this reaction?
- (iii) What mass of excess reactant remains in solution?

(i) Mols. of AgNO₃ = 50.0 x 0.150 / 1000 = 0.0075 mols.

Mols. of Na₂CrO₄ = 50.0 x 0.200 / 1000 = 0.01 mols.

AgNO₃ : Na₂CrO₄

2 : 1

0.0075 mols. : 0.00375 mols.

[Using the mols. of AgNO₃]

0.00375 mols. of Na₂CrO₄ would be required to react with all of the AgNO₃

Since 0.01 mols. of Na₂CrO₄ were present in the rxn., Na₂CrO₄ is in excess.

Therefore, AgNO₃ is the limiting reagent and determines the mass of Ag₂CrO₄ that precipitates out of solution.

(ii) AgNO₃ : Ag₂CrO₄

2 : 1

0.0075 mols. : 0.00375 mols.

Mols. of Ag₂CrO₄ produced = 0.00375 mols.

M_r of Ag₂CrO₄ = 332 g

Mass of Ag₂CrO₄ produced = 0.00375 x 332 = 1.245 g

(iii) Mols. of Na₂CrO₄ remaining = 0.01 – 0.00375 = 0.00625 mols.

M_r of Na₂CrO₄ = 162 g

Mass of Na₂CrO₄ remaining = 0.00625 x 162 = 1.01 g

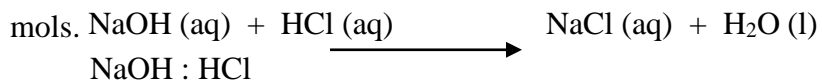
Titration

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 = $25 \times 0.1 / 1000 = 0.0025$



1

:

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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 $0.0025 / 23.5 \times 1000$ 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:

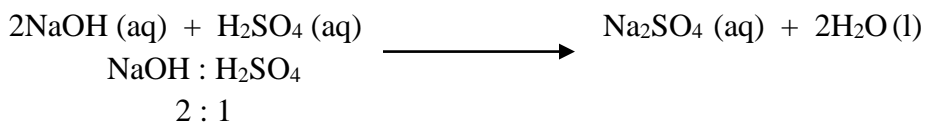
(1) 10.0 g of powder are dissolved and made up to the mark with distilled water in a 250 mL volumetric flask, (2) 25 mL of this solution is reacted completely with 50 mL of 0.1 M H_2SO_4 , (3) this reaction mixture is then titrated against 0.2 M NaOH which yields an average titre volume of 13.35 mL.

(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) = $13.35 \times 0.2 / 1000 = 0.00267$ mols.



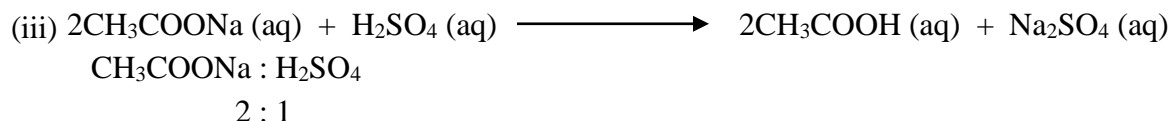
0.00267 mols. : 0.001335 mols.

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) = $50 \times 0.1 / 1000 = 0.005$ mols.

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



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. $\text{CH}_3\text{COONa} = 0.0733$ mols.

M_r of $\text{CH}_3\text{COONa} = 82$ g

Mass of CH_3COONa in 10 g of powder = $0.0733 \times 82 = 6$ g