HL IB Chemistry

Your notes

Counting Particles by Mass: The Mole

Contents

The Mole Unit

Your notes

The Mole

- The **Avogadro constant** (N_A or L) is the number of particles equivalent to the relative **atomic mass** or **molecular mass** of a substance in grams
  - The Avogadro constant applies to atoms, molecules and ions
  - The value of the Avogadro constant is **6.02×10^{23} mol^{-1}**
- The mass of a substance with this number of particles is called the **molar mass**
  - **One mole** of a substance contains the same number of fundamental units as there are atoms in exactly 12.00 g of ^{12}C
  - If you had 6.02×10^{23} atoms of carbon-12 in your hand, you would have a mass of exactly 12.00 g
  - One mole of water would have a mass of $(2 \times 1.01 + 16.00) = 18.02$ g

Your notes

Worked example

Determine the number of atoms, molecules and the relative mass of 1 mole of:

1. Na
2. H_2
3. NaCl

Answer 1:

- The relative atomic mass of Na is 22.99
- Therefore, 1 mol of Na has a mass of 22.99 g mol^{-1}
- 1 mol of Na will contain **6.02 x 10^{23} atoms of Na** (Avogadro's constant)

Answer 2:

- The relative atomic mass of H is 1.01
- Since there are 2 H atoms in H_2, the mass of 1 mol of H_2 is (2 x 1.01) 2.02 g mol^{-1}
- 1 mol of H_2 will contain **6.02 x 10^{23} molecules of H_2**
- However, since there are 2 H atoms in each molecule of H_2, 1 mol of H_2 molecules will contain **1.204 x 10^{24} H atoms**

Answer 3:

- The relative atomic masses of Na and Cl are 22.99 and 35.45 respectively
- Therefore, 1 mol of NaCl has a mass of (22.99 + 35.45) 58.44 g mol^{-1}
- 1 mol of NaCl will contain **6.02 x 10^{23}** formula units of NaCl
- Since there is both an Na and a Cl atom in NaCl, 1 mol of NaCl will contain **1.204 x 10^{24} atoms** in total

Summary:

1 mole of	Number of atoms	Number of molecules/ formula units	Relative mass
Na	6.02 x 10^{23}	-	23.99
H_2	1.204 x 10^{24}	6.02 x 10^{23}	2.02
NaCl	1.204 x 10^{24}	6.02 x 10^{23}	58.44

Your notes

Relative Atomic Mass

Relative atomic mass, A_r

- The **relative atomic mass** (A_r) of an element is the weighted average mass of one atom compared to one twelfth the mass of a carbon-12 atom
- The relative atomic mass is determined by using the weighted average mass of the **isotopes** of a particular element
- The A_r has **no units** as it is a ratio and the units cancel each other out

$$A_r = \frac{weighted\ average\ mass\ of\ one\ atom\ of\ an\ element}{\frac{1}{12} mass\ of\ one\ atom\ of\ carbon-12}$$

Table of Relative Molecular Mass Calculations

Substance	Atoms present	M_r
Hydrogen (H_2)	2 x H	(2 x 1.01) = 2.02
Water (H_2O)	(2 x H) + (1 x O)	(2 x 1.01) + 16.00 = 18.02
Potassium Carbonate (K_2CO_3)	(2 x K) + (1 x C) +(3 x O)	(2 x 39.10) + 12.01+ (3 x 16.00) = 138.21
Calcium hydroxide ($Ca(OH)_2$)	(1 x Ca) + (2 x O) +(2 x H)	40.08 x (2 x 16.00) + (2 x 1.01) = 74.10
Ammonium Sulfate ($(NH_4)_2SO_4$)	(2 x N) + (8 x H) +(1 x S) + (4 x O)	(2 x 14.01) + (8 x 1.01) + 32.07 + (4 x 16.00) = 132.17

Relative formula mass, M_r

- The **relative formula mass** (M_r) is used for compounds containing **ions**
- It is calculated in the same way as **relative molecular mass**
- In the table above, the M_r for potassium carbonate, calcium hydroxide and ammonium sulfate are relative formula masses

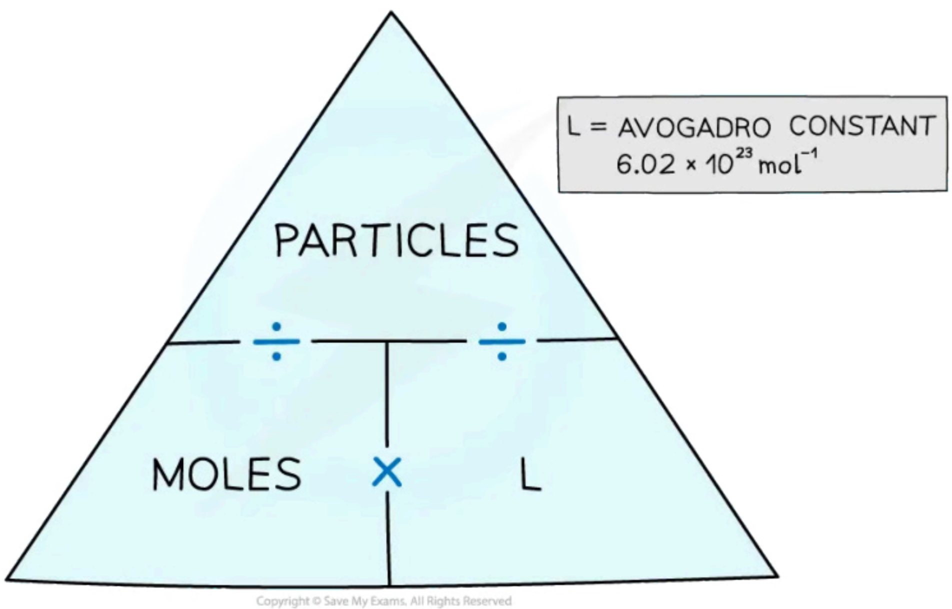
Your notes

Molar Mass

Molar Mass

- Since atoms are so small, any sensible laboratory quantity of substance must contain a huge number of atoms
- Such numbers are not convenient to work with, so using **moles** is a better unit to deal with the sort of quantities of substance normally being measured
- When we need to know the number of particles of a substance, we usually count the number of **moles**
- The number of **moles** or particles can be calculated easily using a formula triangle

Formula triangle diagram linking moles, particles and Avogadro's constant



The moles and particles formula triangle – cover with your finger the one you want to find out and follow the directions in the triangle

Worked example

How many hydrogen atoms are in 0.010 moles of CH_3CHO?

Answer:

- There are 4 H atoms in 1 molecule of CH_3CHO
- So, there are 0.040 moles of H atoms in 0.010 moles of CH_3CHO
- The number of H atoms is the **amount in moles x L**
- This comes to 0.040 x (6.02 x 10^{23}) = **2.4 x 10^{22} atoms**

Worked example

How many moles of hydrogen atoms are in 3.612 x 10^{23} molecules of H_2O_2?

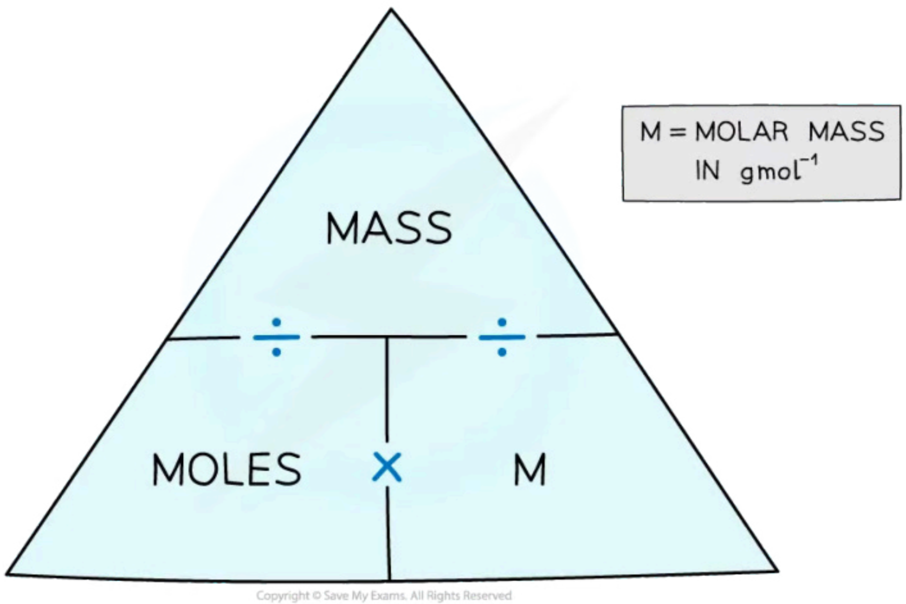
Answer:

- In 3.612 x 10^{23} molecules of H_2O_2 there are 2 x (3.612 x 10^{23}) atoms of H
- So, there are 7.224 x 10^{23} atoms of H
- The number of moles of H atoms is the **number of particles ÷ *L***
- This comes to 7.224 x 10^{23} ÷ (6.02 x 10^{23}) = **1.20 moles of H atoms**

Moles and Mass

- We count in **moles** by weighing the mass of substances
- The number of **moles** can be calculated by using a formula triangle
- The **molar mass** of a substance is its relative atomic mass, Ar, or its relative formula mass, M_r expressed in grams
- Molar mass has the units **g mol^{-1}**

Formula triangle diagram linking moles, mass and molar mass



The moles and mass formula triangle – cover with your finger the one you want to find out and follow the directions in the triangle

Worked example

What is the mass of 0.250 moles of zinc?

Answer:

- From the periodic table the relative atomic mass of Zn is 65.38
- So, the molar mass is 65.38 g mol^{-1}
- The mass is calculated by **moles x molar mass**
- This comes to 0.250 mol x 65.38 g mol^{-1} = **16.3 g**

Worked example

How many moles are in 2.64 g of sucrose, $C_{12}H_{11}O_{22}$ (M_r = 342.3)?

Answer:

- The molar mass of sucrose is 342.3 g mol^{-1}
- The number of moles is found by **mass ÷ molar mass**
- This comes to 2.64 g ÷ 342.3 g mol^{-1} = **7.71 x 10^{-3} mol**

Examiner Tip

Always show your workings in calculations as its easier to check for errors and you may pick up credit if you get the final answer wrong.

Your notes

Empirical Formula

Your notes

Empirical Formula

- The **molecular formula** is the formula that shows the **number** and **type** of each atom in a molecule
  - E.g. the molecular formula of ethanoic acid is $C_2H_4O_2$
- The **empirical formula** is the simplest whole number ratio of the atoms of each element present in one molecule or formula unit of the compound
  - E.g. the empirical formula of ethanoic acid is CH_2O
  - It can be deduced from data that give the **percentage composition by mass** of the elements in a compound
- **Organic molecules** often have **different** empirical and molecular formulae
- The formula of an **ionic compound** is always an **empirical formula**

Worked example

Determine the empirical formula of a compound that contains 10 g of hydrogen and 80 g of oxygen.

Answer:

	Hydrogen	Oxygen
Note the mass of each element	10 g	80 g
Divide the masses by atomic masses	$\frac{10}{1.01}$ = 10 mol	$\frac{80}{16.00}$ = 5 mol
Divide by the lowest figure to obtain nearest whole number ratio	$\frac{10}{5.0}$ =2	$\frac{5.0}{5.0}$ =1
Empirical formula	H_2O	

Your notes

Worked example

Determine the empirical formula of a compound that contains 85.7% carbon and 14.3% hydrogen.

Answer:

	Carbon	Hydrogen
Note the X by mass of each element	85.6	14.3
Divide the X by atomic masses	$\frac{85.7}{12.01}$ =7.14 mol	$\frac{14.3}{1.01}$ = 14.2 mol
Divide by the lowest figure to obtain nearest whole number ratio	$\frac{7.14}{7.14}$ = 1	$\frac{14.2}{7.14}$ = 2
Empirical formula	CH_2	

Molecular formula

- The **molecular formula** gives the actual numbers of each element present in the formula of the compound
- The molecular formula can be found by dividing the **relative molecular mass** by the **relative mass** of the **empirical formula** and finding the multiple that links the empirical formula to the molecular formula
- **Multiply** the empirical formula by this number to find the molecular formula

Worked example

The empirical formula of X is $C_4H_{10}S$ and the relative molecular mass of **X** is 180.42.

What is the molecular formula of **X**?

Relative Atomic Mass Carbon: 12.01 Hydrogen: 1.01 Sulfur: 32.07

Answer:

- **Step 1:** Calculate the relative mass of empirical formula
  Relative empirical mass = (C x 4) + (H x 10) + (S x 1)

  Relative empirical mass = (12.01 x 4) + (1.01 x 10) + (32.07 x 1)

  Relative formula mass = 90.21
- **Step 2:** Divide relative molecular mass of **X** by relative mass of empirical formula
  The multiple between X and the empirical formula = 180.42/90.21 = 2
- **Step 3:** Multiply the empirical formula by 2
  2 x $C_4H_{10}S$ = $C_8H_{20}S_2$

The molecular formula of **X** is $\mathbf{C_8H_{20}S_2}$

Your notes

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Molar Concentration

Molar Concentration

Volumes & concentrations of solutions

- The **concentration** of a solution is the amount of **solute** dissolved in a **solvent** to make 1 dm^3 of **solution**
  - The solute is the substance that dissolves in a solvent to form a solution
  - The solvent is often water
- A **concentrated** solution is a solution that has a **high** concentration of solute
- A **dilute** solution is a solution with a **low** concentration of solute
- Concentration is usually expressed in one of three ways:
  - moles per unit volume
  - mass per unit volume
  - parts per million

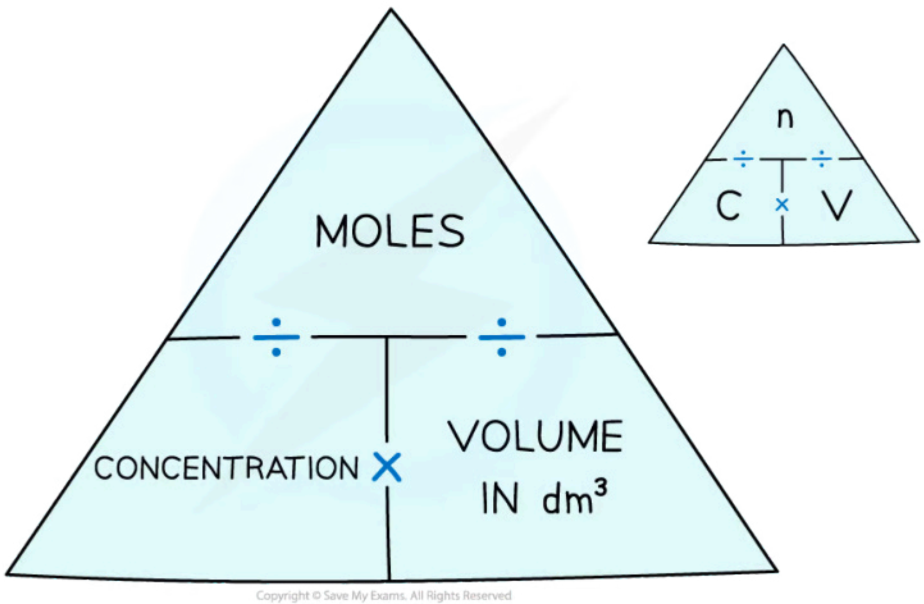
Moles per unit volume

- The formula for expressing concentration using moles is:

$$concentration(mol\ dm^{-3}) = \frac{number\ of\ moles\ of\ solute\ (mol)}{volume\ of\ solution\ (dm^3)}$$

- You must make sure you change cm^3 to dm^3 (by dividing by 1000)
- The relationships can be expressed using this formula triangle:

Concentration moles formula triangle diagram



To use the concentration formula triangle cover the one you want to find out about with your finger and follow the instructions

Worked example

Calculate the mass of sodium hydroxide, NaOH, required to prepare 250 cm^3 of a 0.200 mol dm^{-3} solution.

Answer:

- **Step 1**: Use the formula triangle to find the number of moles of NaOH needed
  number of moles = concentration (mol dm^{-3}) x volume (dm^3)

  n = 0.200 *mol dm^{-3}* x 0.250 *dm^3*

  n = **0.0500 mol**

- **Step 2**: Find the molar mass of NaOH
  M = 22.99 + 16.00 + 1.01 = 40.00 g mol^{-1}

- **Step 3**: Calculate the mass required
  mass = moles x molar mass

  mass = 0.0500 mol x 40.00 g mol^{-1} = **2.00 g**

Mass per unit volume

- Sometimes it is more convenient to express concentration in terms of mass per unit volume

Your notes

- The formula is:

$$concentration\ (g\ dm^{-3}) = \frac{mass\ of\ solute\ (g)}{volume\ of\ solution\ (dm^3)}$$

To change a concentration from mol dm^{-3} to g dm^{-3}

- Multiply the moles of solute by its molar mass

mass of solute (g) = number of moles (mol) x molar mass (g mol^{-1})

Parts per million

- When expressing extremely low concentrations a unit that can be used is **parts per million** or **ppm**
- This is useful when giving the concentration of a pollutant in water or the air when the absolute amount is tiny compared to the volume of water or air
- **1 ppm** is defined as
  - A mass of **1 mg** dissolved in **1 dm^3** of water
- Since 1 dm^3 weighs 1 kg we can also say it is
  - A mass of **1 mg** dissolved in **1 kg** of water, or 10^{-3} g in 10^3 g which is the same as saying the concentration is **1 in 10^6** or **1 in a million**

Worked example

The concentration of chlorine in a swimming pool should between between 1 and 3 ppm. Calculate the maximum mass, in kg, of chlorine that should be present in an Olympic swimming pool of size 2.5 million litres.

Answer:

- **Step 1**: calculate the total mass in mg assuming 3ppm(1 litre is the same as 1 dm^3)

  $3 \times 2.5 \times 10^6 = 7.5 \times 10^6$ mg

- **Step 2**: convert the mass into kilograms (1 mg = 10^{-6} kg)

  $7.5 \times 10^6 \times 10^{-6}$ kg = **7.5 kg**

Your notes

Avogadro's Law

Avogadro's Law

Volumes of gases

- In 1811 the Italian scientist Amedeo **Avogadro** developed a theory about the volume of gases
- **Avogadro's law** (also called **Avogadro's hypothesis)** enables the mole ratio of reacting gases to be determined from volumes of the gases
- **Avogadro** deduced that equal volumes of gases must contain the same number of molecules
- At standard temperature and pressure(**STP**) **one mole** of any gas has a volume of **22.7 dm^3**
- The units are normally written as **dm^3 mol^{-1}**(since it is 'per mole')
- The conditions of **STP** are
  - a temperature of **0° C** (**273 K**)
  - pressure of **100 kPa**

Stoichiometric relationships

- The stoichiometry of a reaction and **Avogadro's Law** can be used to deduce the **exact volumes** of gaseous reactants and products
  - Eg. in the **combustion** of 50 cm^3 of propane, the volume of oxygen needed is (5 x 50) 250 cm^3, and (3 x 50) 150 cm^3 of carbon dioxide is formed, using the ratio of propane: oxygen: carbon dioxide, which is 1: 5: 3 respectively, as seen in the equation

$$C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(l)$$

- Remember that if the gas volumes are not in the same ratio as the coefficients then the amount of product is determined by the limiting reactant so it is essential to identify it first

Your notes

Worked example

What is the total volume of gases remaining when 70 cm^3 of ammonia is combusted completely with 50 cm^3 of oxygen according to the equation shown?

$$4NH_3\,(g) + 5O_2\,(g) \rightarrow 4NO\,(g) + 6H_2O\,(l)$$

Answer:

- **Step 1**: From the equation deduce the molar ratio of the gases, which is NH_3 :O_2 :NO or 4:5:4 (water is not included as it is in the liquid state)
- **Step 2**: We can see that oxygen will run out first (the **limiting reactant**) and so 50 cm^3 of O_2 requires 4/5 x 50 cm^3 of NH_3 to react = 40 cm^3
- **Step 3**: Using Avogadro's Law, we can say 40 cm^3 of NO will be produced
- **Step 4**: There will be of 70–40 = 30 cm^3 of NH_3 left over

Therefore **the total remaining volume will be 40 + 30 = 70 cm^3 of gases**

Examiner Tip

Since gas volumes work in the same way as moles, we can use the '**lowest is limiting**' technique in limiting reactant problems involving gas volumes. This can be handy if you are unable to spot which gas reactant is going to run out first. Divide the volumes of the gases by the cofficients and whichever gives the lowest number is the **limiting reactant**

- E.g. in the previous problem we can see that
  - For NH_3 70/4 gives 17.5
  - For O_2 50/5 gives 10, so **oxygen is limiting**