



# SL IB Chemistry



Your notes

## Counting Particles by Mass: The Mole

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- \* Avogadro's Law

## The Mole Unit



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### 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  $6.02 \times 10^{23} \text{ mol}^{-1}$
- 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  $^{12}\text{C}$
  - If you had  $6.02 \times 10^{23}$  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  $(2 \times 1.01 + 16.00) = 18.02 \text{ g}$



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### Worked example

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

1. Na
2. H<sub>2</sub>
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<sup>-1</sup>
- 1 mol of Na will contain **6.02 x 10<sup>23</sup> 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<sub>2</sub>, the mass of 1 mol of H<sub>2</sub> is (2 x 1.01) 2.02 g mol<sup>-1</sup>
- 1 mol of H<sub>2</sub> will contain **6.02 x 10<sup>23</sup> molecules of H<sub>2</sub>**
- However, since there are 2 H atoms in each molecule of H<sub>2</sub>, 1 mol of H<sub>2</sub> molecules will contain **1.204 x 10<sup>24</sup> 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<sup>-1</sup>
- 1 mol of NaCl will contain **6.02 x 10<sup>23</sup> 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<sup>24</sup> atoms** in total

#### Summary:

1 mole of	Number of atoms	Number of molecules/ formula units	Relative mass
Na	6.02 x 10 <sup>23</sup>	-	23.99
H <sub>2</sub>	1.204 x 10 <sup>24</sup>	6.02 x 10 <sup>23</sup>	2.02
NaCl	1.204 x 10 <sup>24</sup>	6.02 x 10 <sup>23</sup>	58.44



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## 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

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

Table of Relative Molecular Mass Calculations

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

### 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



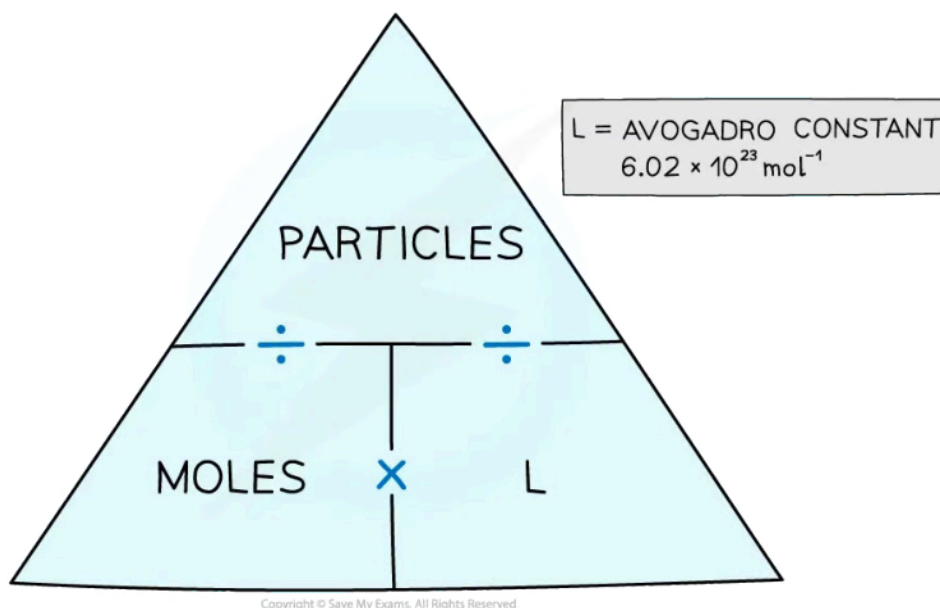
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## Molar Mass

### 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*



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### Worked example

How many hydrogen atoms are in 0.010 moles of  $\text{CH}_3\text{CHO}$ ?

**Answer:**

- There are 4 H atoms in 1 molecule of  $\text{CH}_3\text{CHO}$
- So, there are 0.040 moles of H atoms in 0.010 moles of  $\text{CH}_3\text{CHO}$
- The number of H atoms is the **amount in moles  $\times$  L**
- This comes to  $0.040 \times (6.02 \times 10^{23}) = \mathbf{2.4 \times 10^{22} \text{ atoms}}$

### Worked example

How many moles of hydrogen atoms are in  $3.612 \times 10^{23}$  molecules of  $\text{H}_2\text{O}_2$ ?

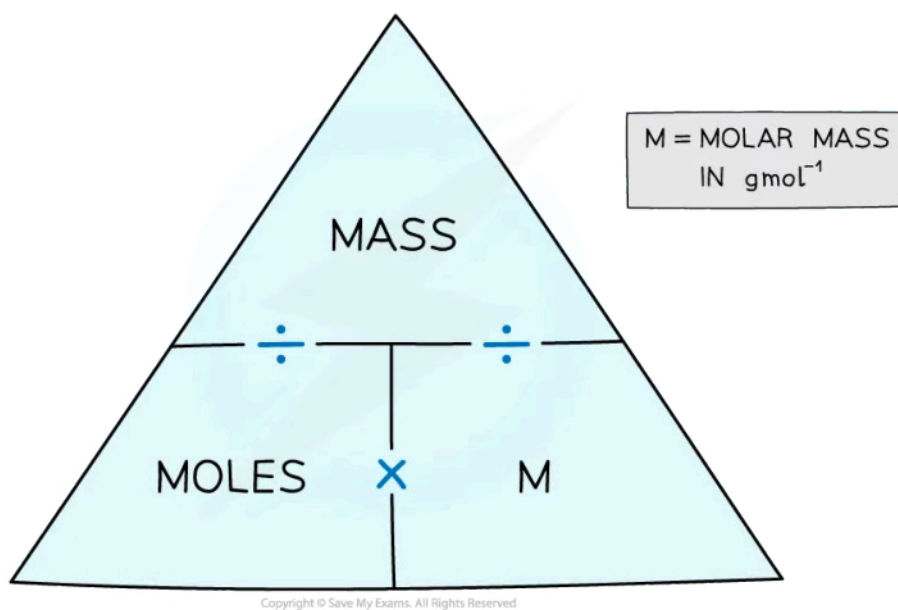
**Answer:**

- In  $3.612 \times 10^{23}$  molecules of  $\text{H}_2\text{O}_2$  there are  $2 \times (3.612 \times 10^{23})$  atoms of H
- So, there are  $7.224 \times 10^{23}$  atoms of H
- The number of moles of H atoms is the **number of particles  $\div$  L**
- This comes to  $7.224 \times 10^{23} \div (6.02 \times 10^{23}) = \mathbf{1.20 \text{ moles of H atoms}}$

## 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,  $A_r$ , or its relative formula mass,  $M_r$ , expressed in grams
- Molar mass has the units  **$\text{g mol}^{-1}$**

**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<sup>-1</sup>
- The mass is calculated by **moles x molar mass**
- This comes to 0.250 mol x 65.38 g mol<sup>-1</sup> = **16.3 g**

### Worked example

How many moles are in 2.64 g of sucrose, C<sub>12</sub>H<sub>22</sub>O<sub>11</sub> (M<sub>r</sub> = 342.3)?

**Answer:**

- The molar mass of sucrose is 342.3 g mol<sup>-1</sup>
- The number of moles is found by **mass ÷ molar mass**
- This comes to 2.64 g ÷ 342.3 g mol<sup>-1</sup> = **7.71 x 10<sup>-3</sup> 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.



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## Empirical Formula

### 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	$H_2O$	



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### 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 <sub>2</sub>	

### 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



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### 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 =  $(C \times 4) + (H \times 10) + (S \times 1)$   
Relative empirical mass =  $(12.01 \times 4) + (1.01 \times 10) + (32.07 \times 1)$   
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  
 $2 \times C_4H_{10}S = C_8H_{20}S_2$

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



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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<sup>3</sup> 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

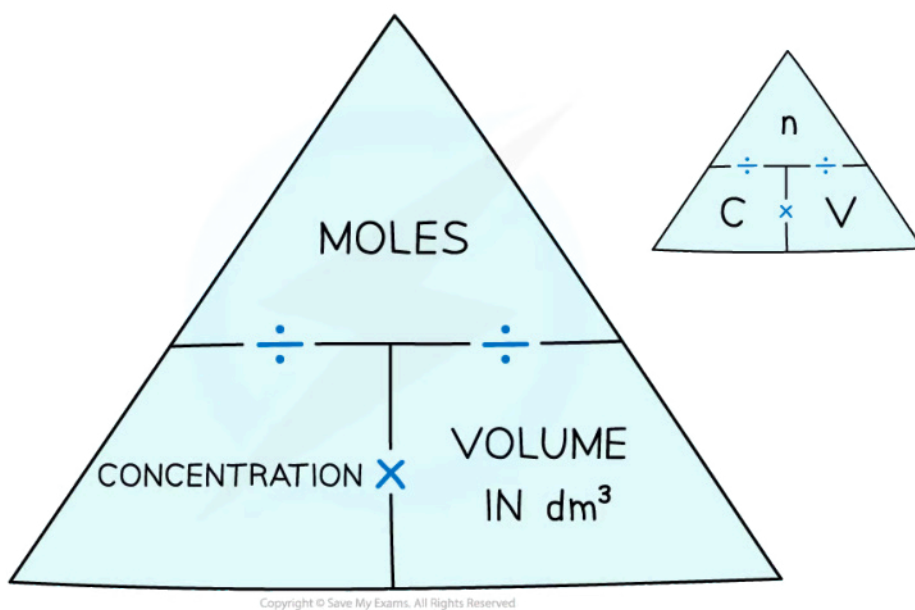
#### Moles per unit volume

- The formula for expressing concentration using moles is:

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

- You must make sure you change cm<sup>3</sup> to dm<sup>3</sup> (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<sup>3</sup> of a 0.200 mol dm<sup>-3</sup> solution.

**Answer:**

- **Step 1:** Use the formula triangle to find the number of moles of NaOH needed

$$\text{number of moles} = \text{concentration (mol dm}^{-3}\text{)} \times \text{volume (dm}^3\text{)}$$

$$n = 0.200 \text{ mol dm}^{-3} \times 0.250 \text{ dm}^3$$

$$n = \mathbf{0.0500 \text{ mol}}$$

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

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

$$\text{mass} = 0.0500 \text{ mol} \times 40.00 \text{ g mol}^{-1} = \mathbf{2.00 \text{ g}}$$

### Mass per unit volume

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

- The formula is:

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

To change a concentration from mol dm<sup>-3</sup> to g dm<sup>-3</sup>

- Multiply the moles of solute by its molar mass

$$\text{mass of solute (g)} = \text{number of moles (mol)} \times \text{molar mass (g mol}^{-1}\text{)}$$

## 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<sup>3</sup>** of water
- Since 1 dm<sup>3</sup> weighs 1 kg we can also say it is
  - A mass of **1 mg** dissolved in **1 kg** of water, or 10<sup>-3</sup> g in 10<sup>3</sup> g which is the same as saying the concentration is **1 in 10<sup>6</sup>** or **1 in a million**

### Worked example

The concentration of chlorine in a swimming pool should be 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<sup>3</sup>)  
 $3 \times 2.5 \times 10^6 = 7.5 \times 10^6 \text{ mg}$
- **Step 2:** convert the mass into kilograms (1 mg = 10<sup>-6</sup> kg)  
 $7.5 \times 10^6 \times 10^{-6} \text{ kg} = \mathbf{7.5 \text{ kg}}$



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## Avogadro's Law

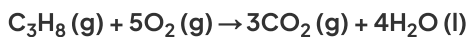
### Avogadro's Law

#### 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<sup>3</sup>**
- The units are normally written as **dm<sup>3</sup> mol<sup>-1</sup>** (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<sup>3</sup> of propane, the volume of oxygen needed is (5 x 50) 250 cm<sup>3</sup>, and (3 x 50) 150 cm<sup>3</sup> of carbon dioxide is formed, using the ratio of propane: oxygen: carbon dioxide, which is 1: 5: 3 respectively, as seen in the equation



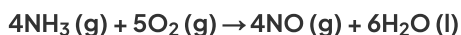
- Remember that if the gas volumes are not in the same ratio as the coefficients then the amount of product is determined by the limiting reactant so it is essential to identify it first



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### Worked example

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



**Answer:**

- **Step 1:** From the equation deduce the molar ratio of the gases, which is NH<sub>3</sub>:O<sub>2</sub>: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<sup>3</sup> of O<sub>2</sub> requires  $\frac{4}{5} \times 50 \text{ cm}^3$  of NH<sub>3</sub> to react = 40 cm<sup>3</sup>
- **Step 3:** Using Avogadro's Law, we can say 40 cm<sup>3</sup> of NO will be produced
- **Step 4:** There will be of 70 - 40 = 30 cm<sup>3</sup> of NH<sub>3</sub> left over

Therefore **the total remaining volume will be 40 + 30 = 70 cm<sup>3</sup> 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 coefficients and whichever gives the lowest number is the **limiting reactant**

- E.g. in the previous problem we can see that
  - For NH<sub>3</sub> 70/4 gives 17.5
  - For O<sub>2</sub> 50/5 gives 10, so **oxygen is limiting**