DP IB Chemistry: SL

Your notes

1.1 Matter, Chemical Change & the Mole Concept

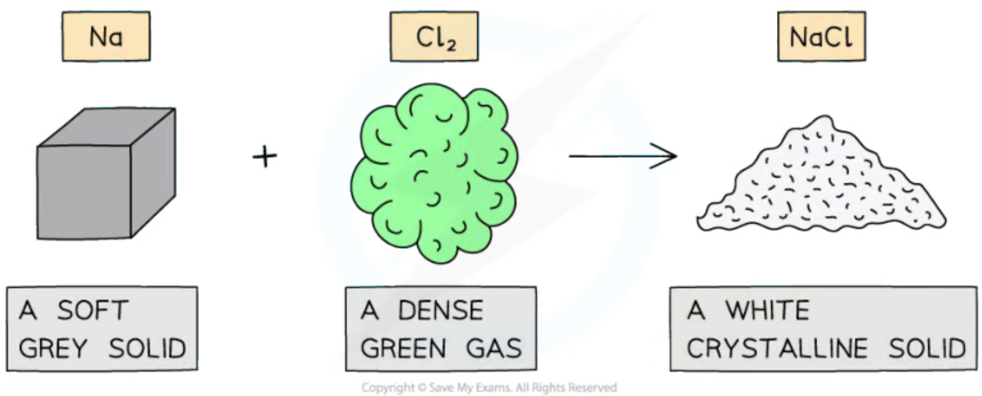
Contents

1.1.1 Elements, Compounds & Mixtures

Your notes

Elements & Compounds

- Elements are substances made from one kind of atom
- Compounds are made from two or more elements **chemically combined**
- Elements take part in chemical reactions in which new substances are made in processes that most often involve an energy change
- In these reactions, atoms combine together in **fixed ratios** that will give them full **outer shells** of electrons, producing **compounds**
- The properties of compounds can be quite different from the elements that form them



The properties of sodium chloride are quite different from sodium and chlorine

Mixtures

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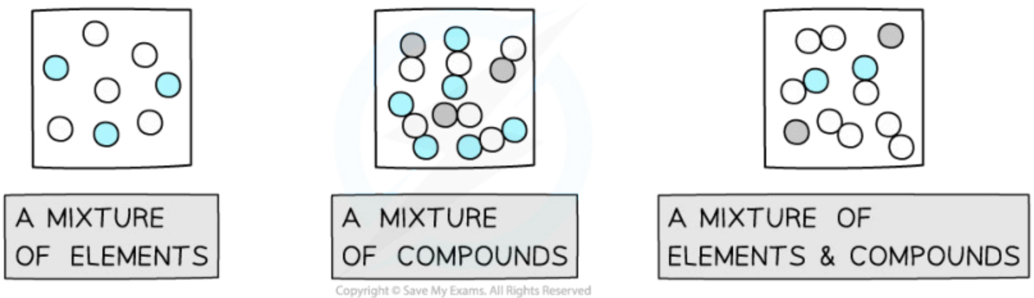
- In a mixture, elements and compounds are interspersed with each other, but are **not** chemically combined
- This means the components of a mixture retain the **same** characteristic properties as when they are in their pure form
- So, for example, the gases nitrogen and oxygen when mixed in air, retain the same characteristic properties as they would have if they were separate
- Substances will burn in air because the oxygen present in the air supports **combustion**



Mixtures at the molecular level

Homogeneous or heterogeneous

- A **homogeneous** mixture has uniform composition and properties throughout
- A **heterogeneous** mixture has non-uniform composition, so its properties are not the same throughout
- It is often possible to see the separate components in a **heterogeneous mixture**, but not in a **homogeneous mixture**

Types of Mixtures

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Mixture	Homogeneous or heterogeneous
Air	Homogeneous
Bronze (an alloy)	Homogeneous
Concrete	Heterogeneous
Orange juice with pulp	Heterogeneous

Separating Mixtures

- The components retain their individual properties in a mixture and we can often separate them relatively easily. The technique we choose to achieve this will take advantage of a suitable difference in the physical properties of the components

Mixtures & Separation Techniques

Mixture	What technique can be used to separate the components	The property that is different in the components
Air	Fractional distillation (of liquid air)	Boiling points
Salt and sand	Solution and filtration	Solubility in water
Pigments in food colours	Paper chromatography	Adsorption (on cellulose)
Sulfur and iron	Use a magnet	Magnetism

Your notes

1.1.2 Equations

Balancing Equations

- A **symbol** equation is a shorthand way of describing a chemical reaction using **chemical symbols** to show the number and type of each atom in the reactants and products
- A **word** equation is a longer way of describing a chemical reaction using only **words** to show the reactants and products

Balancing equations

- During chemical reactions, atoms cannot be **created** or **destroyed**
- The number of each atom on each side of the reaction must therefore be the **same**
  - E.g. the reaction needs to be **balanced**

- When balancing equations remember:
  - Not to change any of the formulae
  - To put the numbers used to balance the equation **in front** of the formulae
  - To balance firstly the carbon, then the hydrogen and finally the oxygen in **combustion reactions** of organic compounds

- When balancing equations follow the following the steps:
  - Write the formulae of the reactants and products
  - Count the numbers of atoms in each reactant and product
  - Balance the atoms one at a time until all the atoms are balanced
  - Use appropriate state symbols in the equation

- The **physical state** of reactants and products in a chemical reaction is specified by using **state symbols**
  - **(s)** solid
  - **(l)** liquid
  - **(g)** gas
  - **(aq)** aqueous

Ionic equations

- In aqueous solutions ionic compounds **dissociate** into their ions
- Many chemical reactions in aqueous solutions involve ionic compounds, however only some of the ions in solution take part in the reactions
- The ions that do **not** take part in the reaction are called **spectator ions**
- An **ionic equation** shows **only** the ions or other particles taking part in a reaction, without showing the spectator ions

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Worked example

Balance the following equation:

Magnesium + Oxygen ⟶ Magnesium oxide

Answer:

Step 1: Write out the symbol equation showing reactants and products

$$Mg + O_2 \rightarrow MgO$$

Step 2: Count the numbers of atoms in each reactant and product

	Mg	O
Reactants	1	2
Products	1	1

Step 3: Balance the atoms one at a time until all the atoms are balanced

$$2Mg + O_2 \rightarrow 2MgO$$

This is now showing that 2 moles of magnesium react with 1 mole of oxygen to form 2 moles of magnesium oxide

Step 4: Use appropriate **state symbols** in the fully balanced equation

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Worked example

1. Balance the following equation

   Zinc + copper(II) sulfate ⟶ Zinc sulfate + copper
2. Write down the ionic equation for the above reaction

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Answer 1:

Step 1: To balance the equation, write out the symbol equation showing reactants and products

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

Step 2: Count the numbers of atoms in each reactant and product. The equation is already balanced

	Zn	Cu	S	O
Reactants	1	1	1	4
Products	1	1	1	4

Step 3: Use appropriate **state symbols** in the equation

$$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$$

Answer 2:

Step 1: The full chemical equation for the reaction is

$$Zn\ (s) + CuSO_4\ (aq) \rightarrow ZnSO_4\ (aq) + Cu\ (s)$$

Step 2: Break down reactants into their respective ions

$$Zn\ (s) + Cu^{2+}\ SO_4^{2-}\ (aq) \rightarrow Zn^{2+}SO_4^{2-}\ (aq) + Cu\ (s)$$

Step 3: Cancel the spectator ions on both sides to give the ionic equation

$$Zn\ (s) + Cu^{2+}\cancel{SO_4^{2-}}\ (aq) \rightarrow Zn^{2+}\cancel{SO_4^{2-}}\ (aq) + Cu\ (s)$$

$$Zn\ (s) + Cu^{2+}(aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)$$

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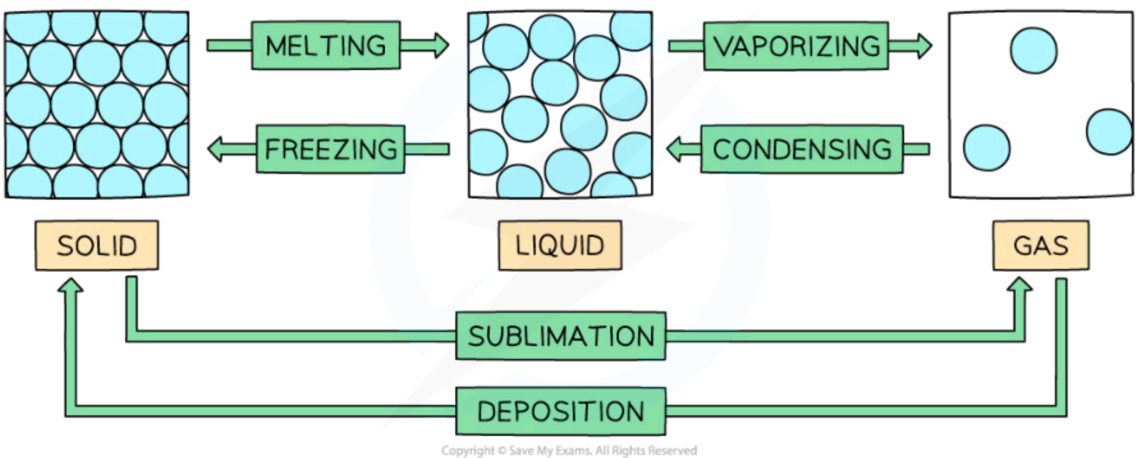
1.1.3 State Changes

State Changes

- Changes of state are **physical changes** that are reversible
- These changes do not change the chemical properties or chemical makeup of the substances involved
- **Vaporisation** includes **evaporation** and **boiling**
- **Evaporation** involves the change of liquid to gas, but unlike boiling, **evaporation** occurs only at the surface and takes place at temperatures below the **boiling point**
- **Boiling** occurs at a specific temperature and takes place when the **vapour pressure** reaches the external atmospheric pressure



State Changes

- The relationship between temperature and energy during state changes can be represented graphically

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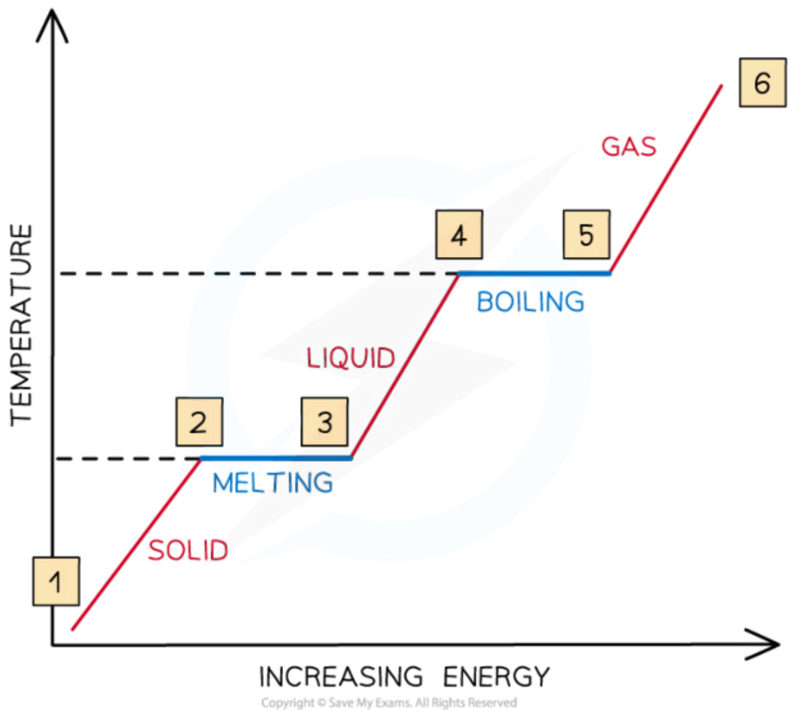


The relationship between temperature and energy during state changes

- Between 1 & 2, the particles are vibrating and gaining **kinetic energy** and the temperature rises
- Between 2 & 3, all the energy goes into breaking bonds – there is **no** increase in **kinetic energy** or **temperature**
- Between 3 & 4, the particles are moving around and gaining in **kinetic energy**
- Between 4 & 5, the substance is boiling, so bonds are breaking and there is **no** increase in **kinetic energy** or **temperature**
- From 5 & 6, the particles are moving around rapidly and increasing in **kinetic energy**

Examiner Tip

Be careful to match the bond breaking or bond making processes to the flow of energy during state changes.

Remember: To **break** bonds, energy is always **needed** to overcome the **forces of attraction** between the particles

1.1.4 The Mole Concept

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} g 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 x 1.01 + 16.00) = 18.02 g

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 **6.02×10^{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 **6.02×10^{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 **1.204×10^{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 **6.02×10^{23}** formula units of NaCl
- Since there is both an Na and a Cl atom in NaCl, 1 mol of NaCl will contain **1.204×10^{24} atoms** in total

1 mole of	Number of atoms	Number of molecules/ formula units	Relative mass
Na	6.02×10^{23}	–	22.99
H_2	1.204×10^{24}	6.02×10^{23}	2.02
NaCl	1.204×10^{24}	6.02×10^{23}	58.44

Your notes

Your notes

Relative 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

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

Relative isotopic mass

- The **relative isotopic mass is** the mass of a particular atom of an **isotope** compared to one twelfth the mass of a carbon-12 atom
- Atoms of the same element with a different number of neutrons are called **isotopes**
- **Isotopes** are represented by writing the **mass number** as ^{20}Ne, or neon-20 or Ne-20
  - To calculate the average atomic mass of an element the **percentage abundance** is taken into account
  - Multiply the atomic mass by the percentage abundance for each isotope and add them all together
  - Divide by 100 to get average relative atomic mass
  - This is known as the **weighted average** of the masses of the isotopes

$$Relative\ atomic\ mass = \frac{\Sigma(isotope\ abundance \times relative\ isotopic\ mass)}{100}$$

Relative molecular mass, M_r

- The **relative molecular mass** (M_r) is the weighted average mass of a molecule compared to one twelfth the mass of a carbon-12 atom
- The M_r has **no units**

$$M_r = \frac{weighted\ average\ mass\ of\ one\ molecule\ of\ a\ compound}{\frac{1}{12} mass\ of\ one\ atom\ of\ carbon{-}12}$$

- The M_r can be found by adding up the **relative atomic masses** of all atoms present in one molecule

Your notes

- When calculating the M_r the **simplest formula** for the compound is used, also known as the **formula unit**
  - E.g. Silicon dioxide has a giant covalent structure, but the simplest formula (the **formula unit**) is SiO_2

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

Relative formula mass, M_r

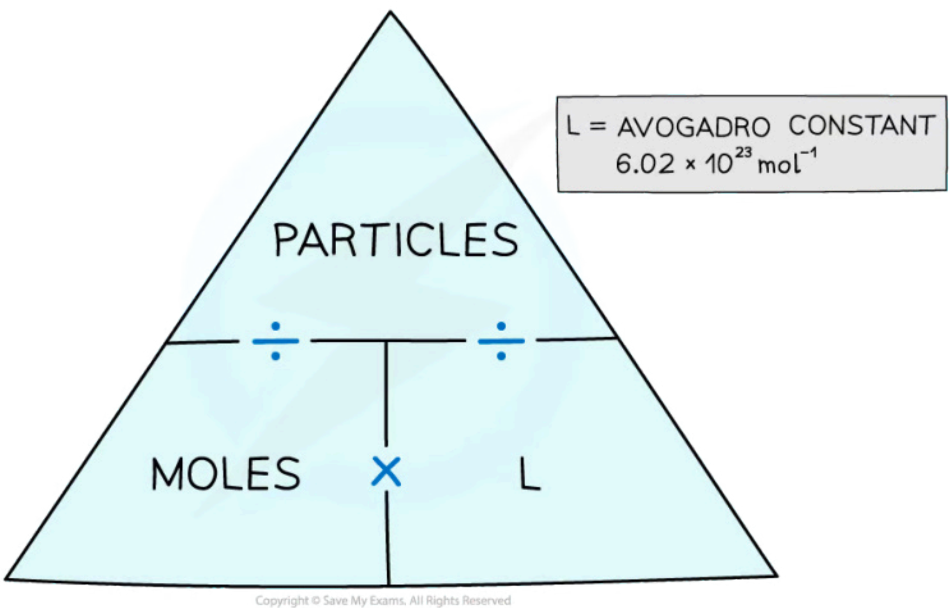
- The **relative formula mass** (M_r) is used for compounds containing **ions**
- It has the same units and is calculated in the same way as the **relative molecular mass**
- In the table above, the M_r for potassium carbonate, calcium hydroxide and ammonium sulfate are relative formula masses

1.1.5 Moles-Mass Problems

Your notes

Moles, Particles & Masses

- 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



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

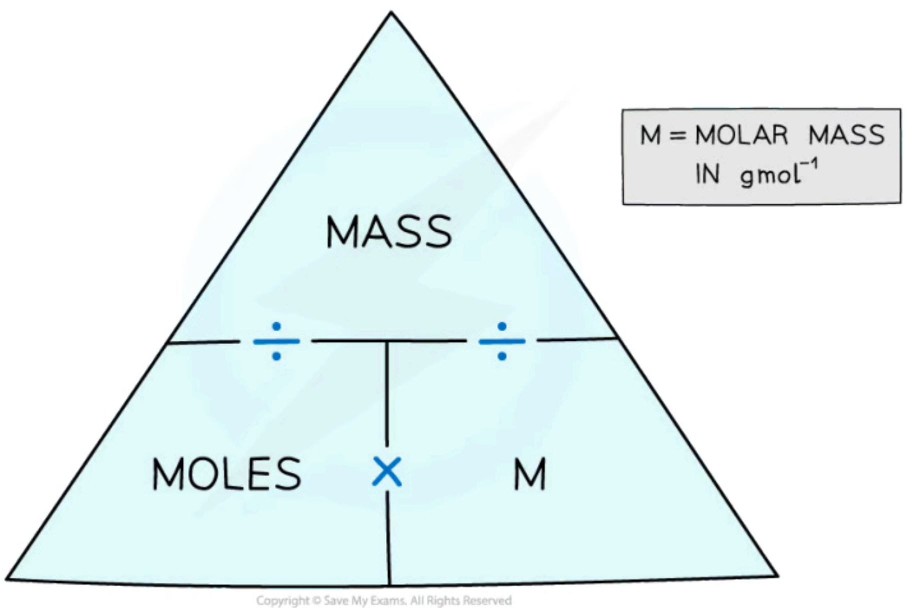
How many moles of hydrogen atoms are in 3.612 x 10^{23} molecules of H_2O_2?

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 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?

Your notes

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.

1.1.6 Empirical Formulae

Your notes

Empirical & Molecular Formulae

- 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
- **Organic molecules** often have **different** empirical and molecular formulae
- The formula of an **ionic compound** is always an **empirical formula**

Empirical Formula Calculations

Your notes

Empirical formula

- 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
- It is calculated from a knowledge of the masses of each element in a sample of the compound
- It can also be deduced from data that give the **percentage composition by mass** of the elements in a compound

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	$= \frac{80}{16.00}$ = 5.0 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	

Worked example

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

Answer:

Your notes

	Carbon	Hydrogen
Note the % by mass of each element	85.7	14.3
Divide the % 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

$2 \times C_4H_{10}S \;=\; C_8H_{20}S_2$

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

Your notes