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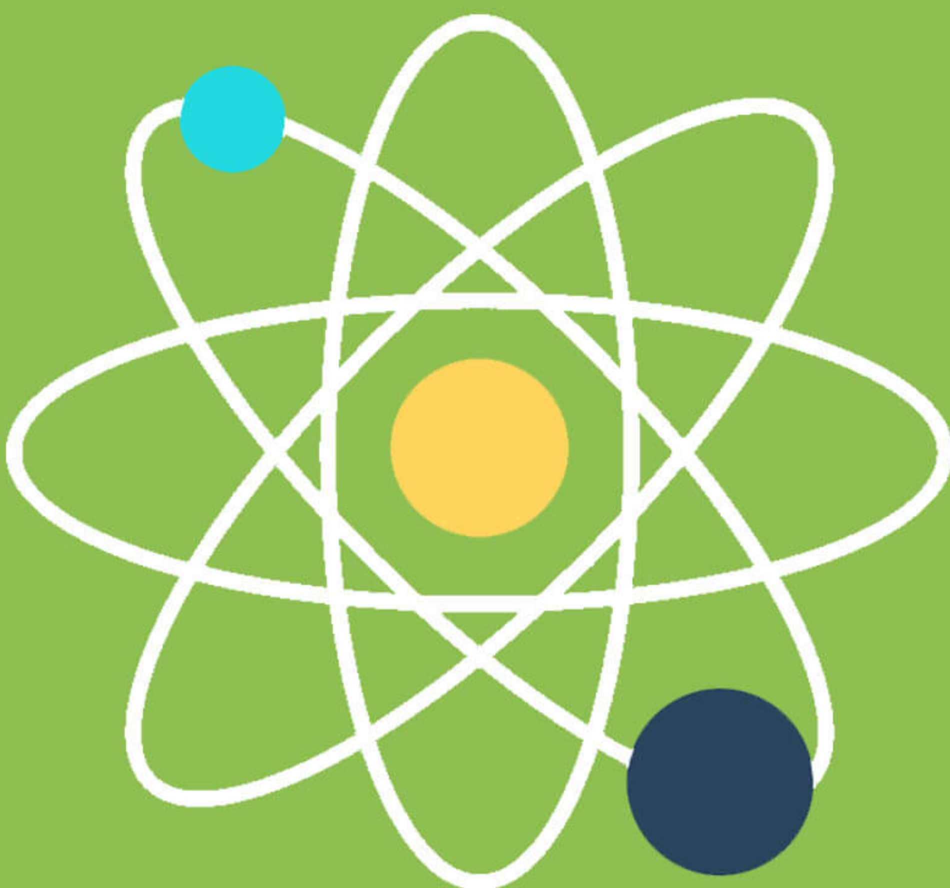
1.1 Matter, Chemical Change & the Mole Concept



IB Chemistry - Revision Notes

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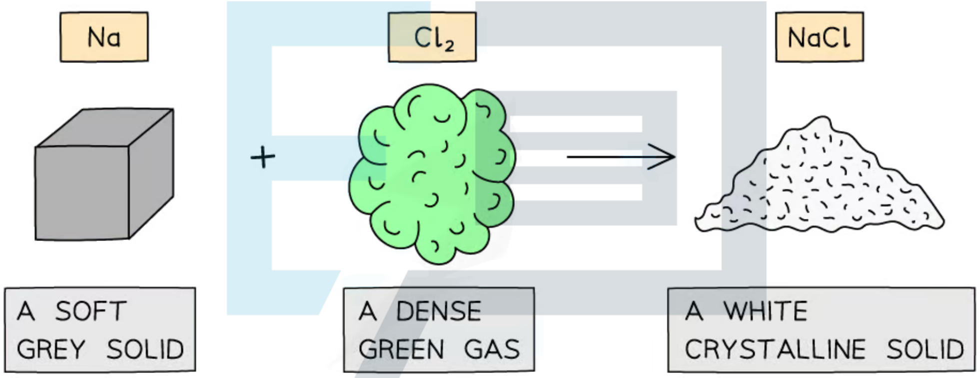
1.1.1 Elements, Compounds & Mixtures

Elements & Compounds

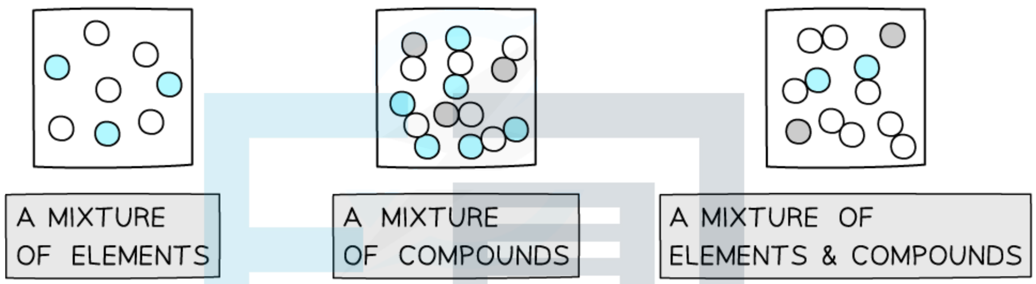
- Elements are substances made from one kind of atom
- Compounds are made from two or more elements **chemically combined**
- Elements take part in chemical reactions in which new substances are made in processes that most often involve an energy change
- In these reactions, atoms combine together in **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

- 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**

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

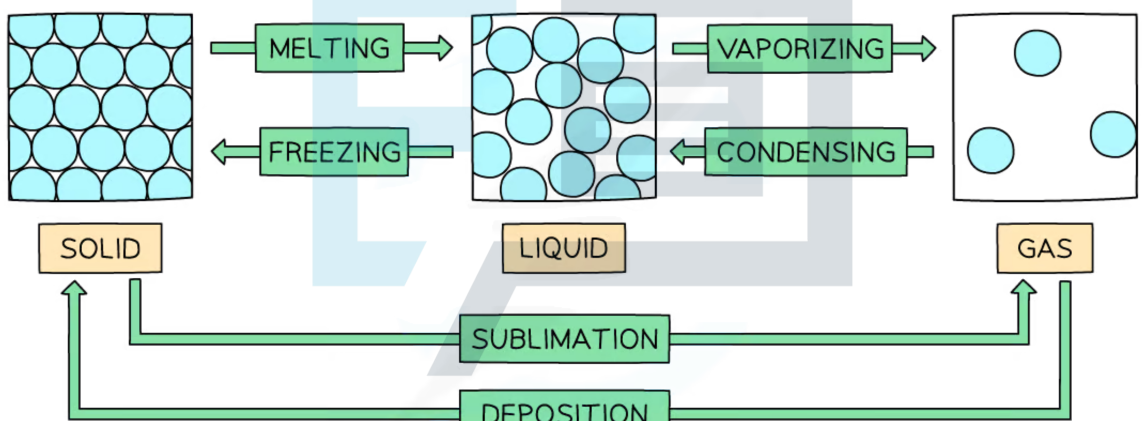
Ionic equations

- 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

1.1.3 State Changes

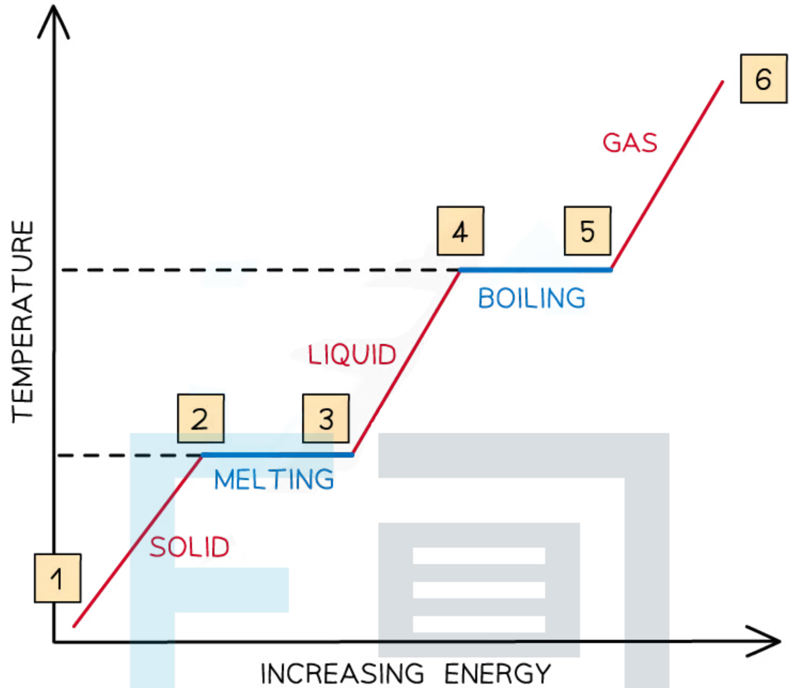
State Changes

- 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



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**

Exam 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

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

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
- 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

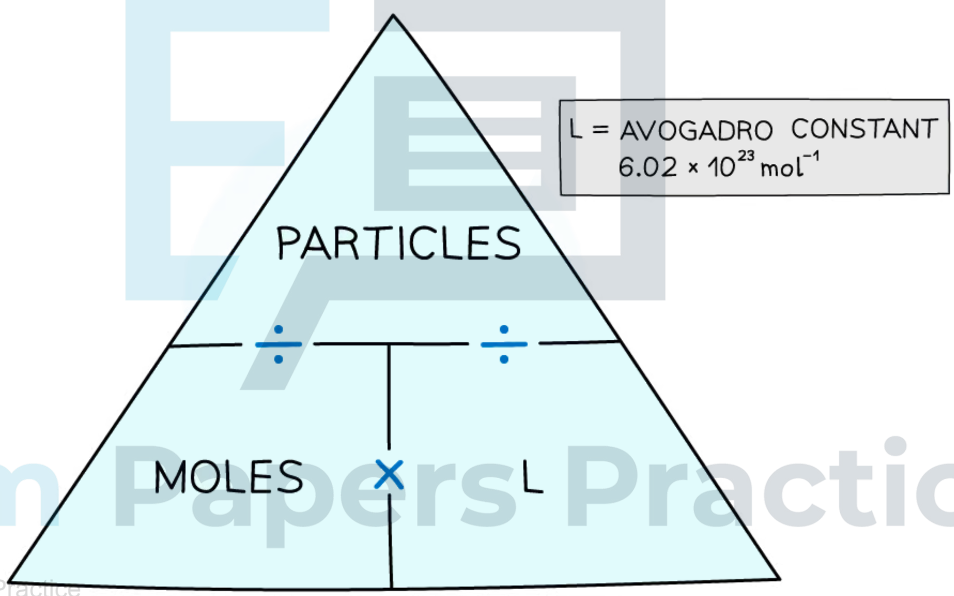
Relative formula mass, M_r

- 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

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 (6.02×10^{23}) = **2.4×10^{22} atoms**

Worked example

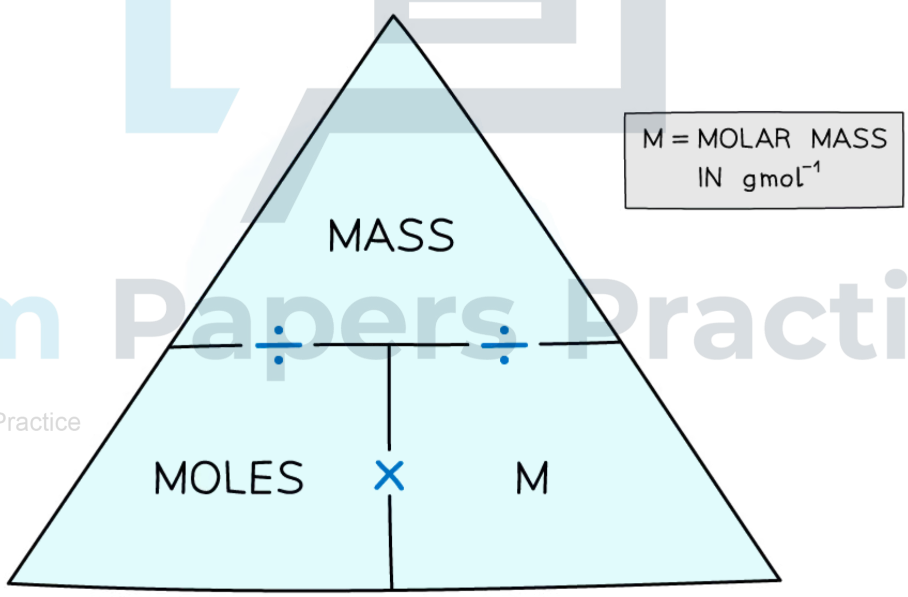
How many moles of hydrogen atoms are in 3.612×10^{23} molecules of H_2O_2?

Answer:

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

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**

Exam 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.