

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

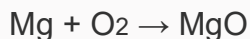
Worked example

Balance the following equation:



Answer:

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

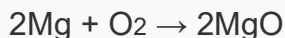


Step 2: Count the numbers of atoms in each reactant and product

	Mg	O
Reactants	1	2
Products	1	1

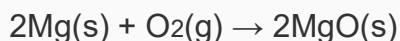
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Step 3: Balance the atoms one at a time until all the atoms are balanced



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



Worked example

- Balance the following equation
Zinc + copper(II) sulfate \rightarrow Zinc sulfate + copper
- Write down the ionic equation for the above reaction

Answer 1:

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

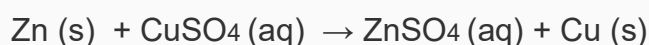


Step 2: Count the numbers of atoms in each reactant and product. The equation is already balanced

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

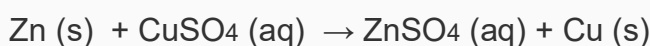
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Step 3: Use appropriate state symbols in the equation

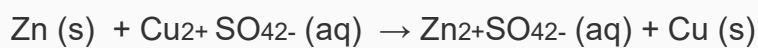


Answer 2:

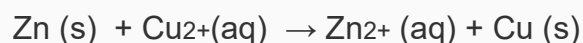
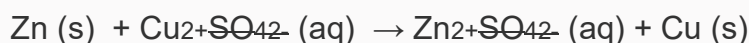
Step 1: The full chemical equation for the reaction is



Step 2: Break down reactants into their respective ions



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



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}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 $(2 \times 1.01 + 16.00) = 18.02 \text{ g}$

Worked example

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

1. Na
2. H₂
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 H₂, the mass of 1 mol of H₂ is (2×1.01) 2.02 g mol⁻¹
- 1 mol of H₂ will contain 6.02×10^{23} molecules of H₂
- However, since there are 2 H atoms in each molecule of H₂, 1 mol of H₂ 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

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

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

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

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

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

- 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	M_r
Hydrogen (H_2)	$2 \times \text{H}$	$(2 \times 1.01) = 2.02$
Water (H_2O)	$(2 \times \text{H}) + (1 \times \text{O})$	$(2 \times 1.01) + 16.00 = 18.02$
Potassium Carbonate (K_2CO_3)	$(2 \times \text{K}) + (1 \times \text{C}) + (3 \times \text{O})$	$(2 \times 39.10) + 12.01 + (3 \times 16.00) = 138.21$
Calcium Hydroxide ($\text{Ca}(\text{OH})_2$)	$(1 \times \text{Ca}) + (2 \times \text{O}) + (2 \times \text{H})$	$40.08 + (2 \times 16.00) + (2 \times 1.01) = 74.10$
Ammonium Sulfate ($(\text{NH}_4)_2\text{SO}_4$)	$(2 \times \text{N}) + (8 \times \text{H}) + (1 \times \text{S}) + (4 \times \text{O})$	$(2 \times 14.01) + (8 \times 1.01) + 32.07 + (4 \times 16.00) = 132.17$

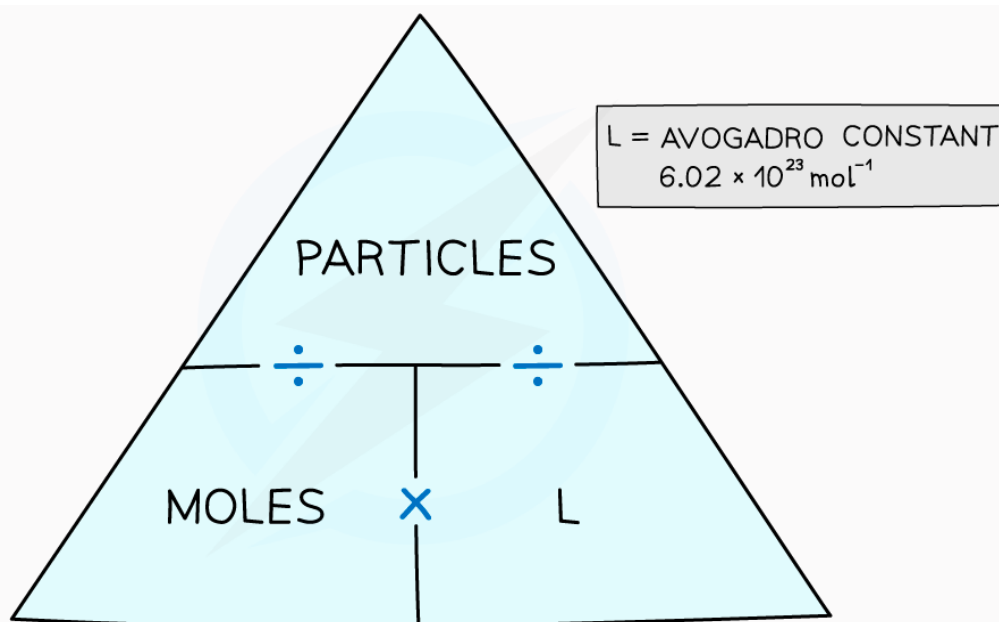
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Relative formula mass, M_r

- 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

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 \times L
- This comes to $0.040 \times (6.02 \times 10^{23}) = 2.4 \times 10^{22}$ atoms

Worked example

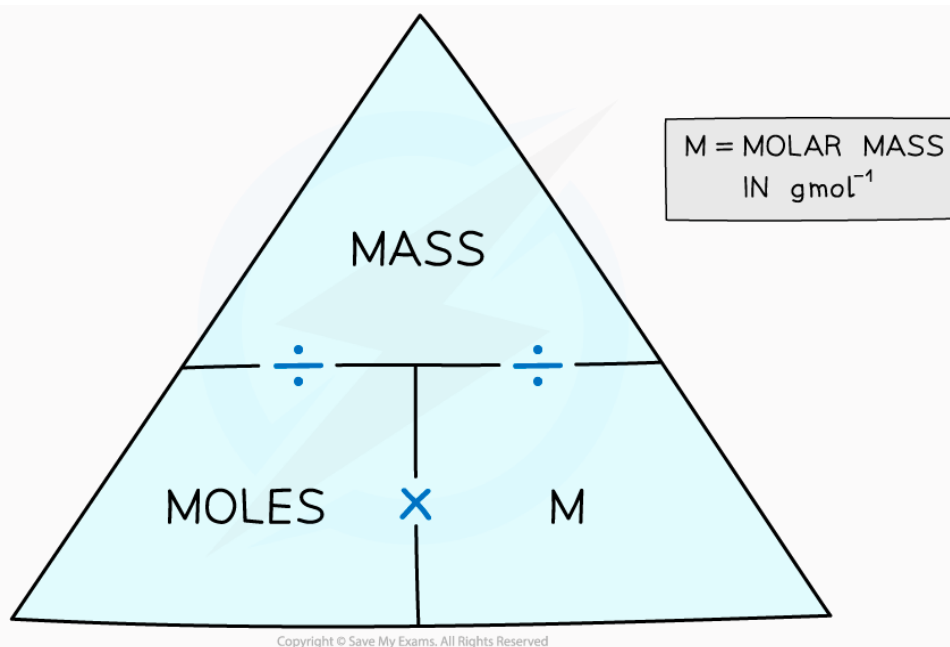
How many moles of hydrogen atoms are in 3.612×10^{23} molecules of H_2O_2 ?

Answer:

- In 3.612×10^{23} molecules of H_2O_2 there are $2 \times (3.612 \times 10^{23})$ 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 \div L
- This comes to $7.224 \times 10^{23} \div (6.02 \times 10^{23}) = 1.20$ 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, C₁₂H₂₂O₁₁ (*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 it's easier to check for errors and you may pick up credit if you get the final answer wrong.

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 C₂H₄O₂
- 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 CH₂O
- Organic molecules often have different empirical and molecular formulae
- The formula of an ionic compound is always an empirical formula

Empirical Formula Calculations

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 ₂ O	

Worked example

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

Answer:

	Carbon	Hydrogen
Note the % by mass of each element	85.7	14.3
Divide the % by atomic masses	$= \frac{85.7}{12.01}$ $= 7.14 \text{ mol}$	$= \frac{14.3}{1.01}$ $= 14.2 \text{ mol}$
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	

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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 $\text{C}_4\text{H}_{10}\text{S}$ 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 = $(\text{C} \times 4) + (\text{H} \times 10) + (\text{S} \times 1)$

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



The molecular formula of X is $\text{C}_8\text{H}_{20}\text{S}_2$