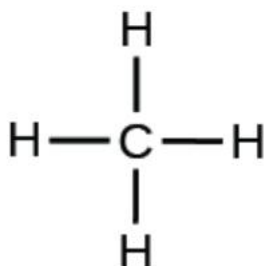


3 Stoichiometry

3.1 Formulae

1 State the formulae of the elements and compounds named in the subject content AND 3 Deduce the formula of a simple compound from the relative numbers of atoms present in a model or a diagrammatic representation



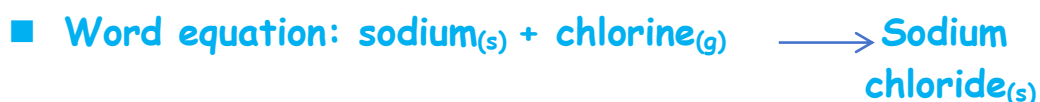
- This is methane, it has the elements carbon(C) and hydrogen(H). The formula of methane is CH₄.

2 Define the molecular formula of a compound as the number and type of different atoms in one molecule

- Example: the water molecule has 2 hydrogen atoms and one oxygen atom. The molecular formula is H₂O

4 Construct word equations and symbol equations to show how reactants form products, including state symbols AND 7 Construct symbol equations with state symbols, including ionic equations AND 8 Deduce the symbol equation with state symbols for a chemical reaction, given relevant information

- (s)=solid, (l)=liquid, (g)=gas, (aq)= aqueous solution
- Example: The reaction of chlorine with and sodium.



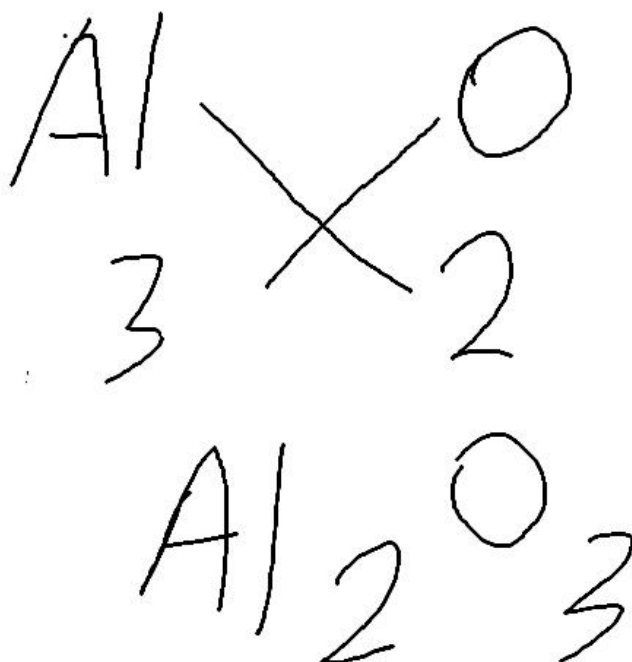


5 Define the empirical formula of a compound as the simplest whole number ratio of the different atoms or ions in a compound

- Example: The molecular formula of glucose is $\text{C}_6\text{H}_{12}\text{O}_6$, so the empirical formula will be CH_2O .

6 Deduce the formula of an ionic compound from the relative numbers of the ions present in a model or a diagrammatic representation or from the charges on the ions

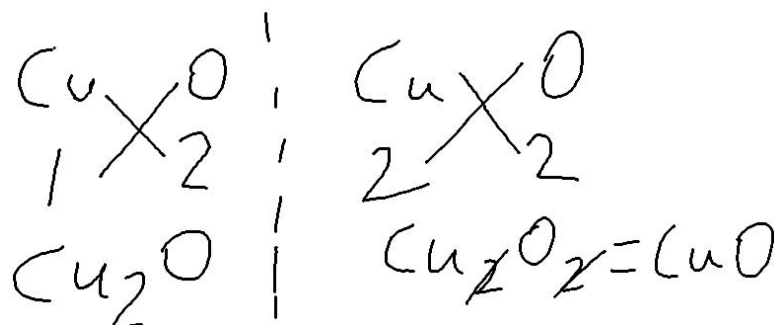
- Example: The aluminium ion has a charge of $3+$ and the oxygen ion has a charge of $2-$ so the formula of aluminium oxide is Al_2O_3 .
- Working:



- The charge is obtained from the group number of the atom:

Group	Elements (examples)	Typical Ion Charge
Group 1 (Alkali metals)	Li, Na, K	+1
Group 2 (Alkaline earth metals)	Mg, Ca, Ba	+2
Group 3 (Boron group)	B, Al	+3
Group 4 (Carbon group)	C, Si, Sn	± 4 (can be +4 or -4 depending on compound)
Group 5 (Nitrogen group)	N, P	-3 (non-metals), sometimes +3 or +5 in covalent compounds
Group 6 (Oxygen group / Chalcogens)	O, S	-2 (non-metals), sometimes +4 or +6 in covalent compounds
Group 7 (Halogens)	F, Cl, Br, I	-1
Group 8 (Noble gases)	He, Ne, Ar	0 (generally do not form ions)
hydrogen	H	+1 when it loses electrons and -1 when it gains electrons

- When writing the formulae or name of compounds, start with the cation then the anion; sodium chloride is NaCl as sodium is the cation.
- For transition elements, you must include the charge in the name as they could have multiple charges. Copper has two charges(valencies), +1 and +2. CuO is copper(ii) oxide and Cu₂O is copper(I) oxide. Below is the working.



3.2 Relative masses of atoms and molecules

1 Describe relative atomic mass, A_r , as the average mass of the isotopes of an element compared to $1/12$ th of the mass of an atom of ^{12}C AND 2 Define relative molecular mass, M_r , as the sum of the relative atomic masses. Relative formula mass, M_r , will be used for ionic compounds

- The A_r of sodium is 23 and the periodic table of chlorine is 35.5. The M_r of sodium chloride is 58.5(23+35.5)

3 Calculate reacting masses in simple proportions.
Calculations will not involve the mole concept

- Example: formation of water:
 - $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ (4 g H_2 + 32 g $\text{O}_2 \rightarrow$ 36 g H_2O)
 - There are 2 hydrogen molecules. Each molecule has 2 atoms and each atom has an A_r of 1. The total molar mass of hydrogen is 4g(1*4 atoms).
 - There is one oxygen molecule with 2 atoms. Each atom has an A_r of 16. the total mass of oxygen is 32g(2 atoms*16)
 - The total mass of water is 36 grams.

NOTE: The total mass of reactants is always equal to the total mass of products.

3.3 The mole and the Avogadro constant

1 **State that** concentration can be measured in g/dm^3 or mol/dm^3

2 **State that** the mole, mol, is the unit of amount of substance and that one mole contains 6.02×10^{23} particles, e.g. atoms, ions, molecules; this number is the Avogadro constant

- One mole of hydrogen atoms = 6.02×10^{23} hydrogen atoms
- One mole of hydrogen molecules = 6.02×10^{23} hydrogen molecules AND 1.204×10^{24} hydrogen ATOMS ($6.02 \times 10^{23} \times 2$ hydrogen atoms). each molecule of hydrogen has 2 hydrogen atoms.
- 1 mole of sodium chloride = 6.02×10^{23} sodium chloride particles and 6.02×10^{23} sodium ions AND 6.02×10^{23} chloride ions AND a total of 1.204×10^{24} ions.

3 **Use the relationship** amount of substance (mol) = mass (g)/molar mass (g/mol) to calculate:

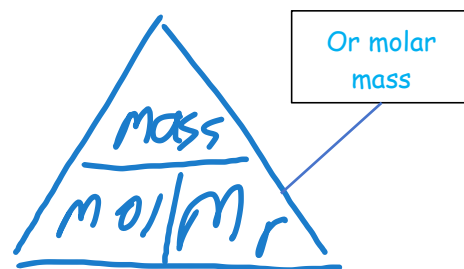
(a) amount of substance[mol]

(b) Mass[g]

(c) molar mass[g/mol]

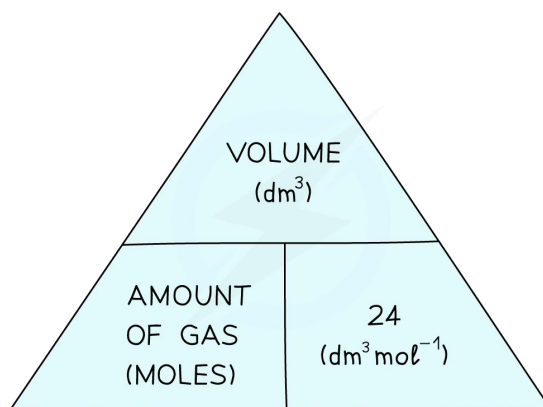
(d) relative atomic mass or relative molecular/formula mass[no unit]

(e) number of particles, using the value of the Avogadro constant



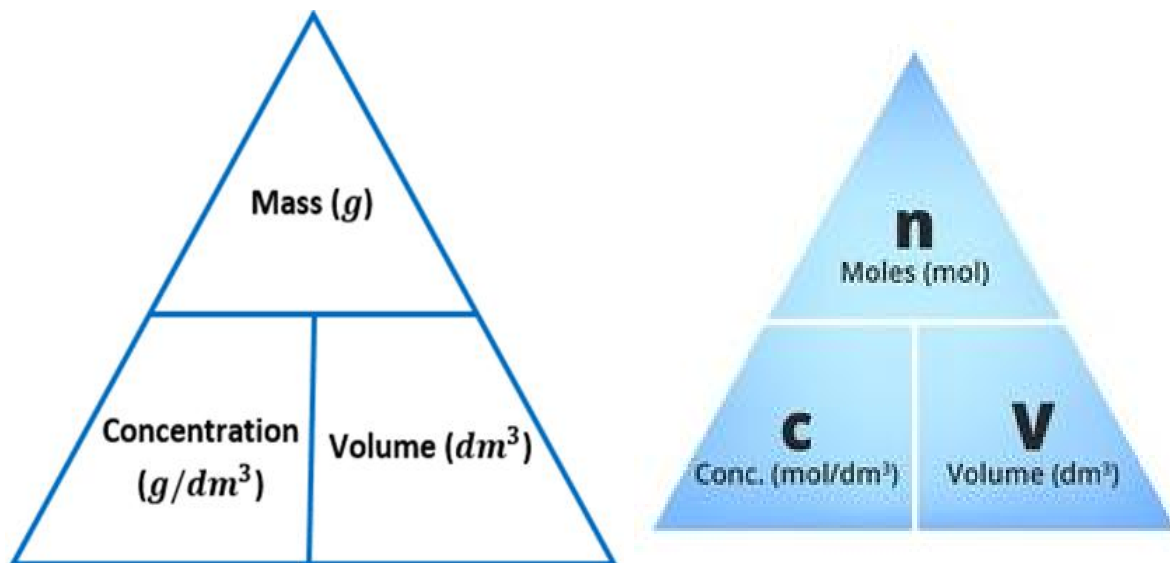
- **Example:** calculate the number of particles in 117g of NaCl:
 - **Step 1:** get A_r and the M_r . The M_r of NaCl is 58.5
 - **Step 2:** get the molar mass(or use the M_r directly). The molar mass is 58.5g/mol
 - **Step 3:** use the molar mass(or M_r) AND the mass to get the number of moles. $117\text{g}/58.5\text{g/mol} = 2 \text{ mols}$.
 - **Step 4:** calculate the number of particles in one mol. 2 moles(117g) of NaCl has 1.204×10^{24} particles.($6.02 \times 10^{23} \times 2 \text{ moles}$).

4 Use the molar gas volume, taken as 24dm^3 at room temperature and pressure, r.t.p., **in calculations involving gases AND 5 Calculate** stoichiometric reacting masses, limiting reactants, volumes of gases at r.t.p., volumes of solutions and concentrations of solutions expressed in g/dm^3 and mol/dm^3 , including conversion between cm^3 and dm^3



- **Example 1:** Calculate the amount of gas in 240cm^3
 - **Step 1:** convert cm^3 or dm^3 . $240/1000 = 0.24\text{dm}^3$ ($1\text{dm}^3 = 1000\text{cm}^3$)

- Step 2: divide the volume by $24 \text{ dm}^3/\text{mol}$. $0.24/24 = 0.01$ moles of gas.



- Example 2: calculate the concentration(g/dm^3) of 150cm^3 of 0.75mol/dm^3 of dilute sulfuric acid(H_2SO_4):
 - Step 1: find the volume in dm^3 . $150/1000 = 0.15\text{dm}^3$.
 - Step 2: find the number of moles.
 $0.15\text{dm}^3/0.75\text{mol/dm}^3 = 0.2$ moles
 - Step 3: get the mass. 0.2 moles and $58.5 = 11.7\text{g}$
 - Step 4: get the concentration in g/dm^3 .
 $11.7\text{g}/0.15\text{dm}^3 = 78\text{g/dm}^3$
- Example 3: 2g of hydrogen react with 32g of oxygen to give one mole of water. Which is the limiting reactant.
 - Step 1: write the equation. $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
 - Step 2: calculate the number of moles of water and hydrogen. This gives us one mole of hydrogen gas ($2/2$) and 1 mole of oxygen($32/32$).

- **Step 3: Conclusion.** Hydrogen is the limiting reactant as only one mole is provided. According to the equation 2 moles of hydrogen is needed to react completely with 1 mole of oxygen. Molar ratio is 2 hydrogen: 1 oxygen.

6 Use experimental data from a titration to calculate the moles of solute, or the concentration or volume of a solution

Titration experiments are done to find the volume of a base that is needed to neutralise all the acid.

- **Example :** Calculate the volume of 0.25mol/dm³ of NaOH needed to neutralise 50cm³ of 1.5mol/dm³ of HCl.
 - **Step 1:** write down the reaction equation(if not provided): $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$
 - **step 2:** get the volume of HCl in dm³. $50/1000 = 0.05\text{dm}^3$.
 - **Step 3:** get the number of moles of HCl. $0.05\text{dm}^3 \times 1.5\text{mol/dm}^3 = 0.075 \text{ moles}$
 - **Step 4:** get the mole ratio of HCl and NaOH. The ratio is 1:1(if there is nothing written next to the chemical then there is one mole so it is HCl instead of 1HCl):
 - **Step 5:** get the number of moles of NaOH. Since the ratio is 1:1 then the number of NaOH is also 0.075 moles
 - **step 6:** get the volume of NaOH. $0.075 \text{ mols} / 0.25\text{mol/dm}^3 = 0.3 \text{ dm}^3$.

7 Calculate empirical formulae and molecular formulae, given appropriate data

- Example: A compound contains 40% carbon, 6.7% hydrogen, and 53.3% oxygen. Find the empirical formula and the molecular formula if the Mr is 60
- step 1: get the simplest ratio of moles by dividing the % by the Ar. Here is the working.

	C	H	O
%	40%	6.7%	53.3%
Ar	12	1	16
	3.33	6.7	3.33
ratio	1	2	1

- Step 2: use the ratio to get the empirical formula and then the Mr of the empirical formula which is CH_2O . The Mr will be $30(12+2(1)+16)$.
- Step 3: divide the Mr of molecular formula by Mr of empirical formula to get the multiplier. $60/30 = 2$
- Step 4: multiply the multiplier by the empirical formula to get the molecular formula. $2(\text{CH}_2\text{O}) = \text{C}_2\text{H}_4\text{O}_2$.

8 Calculate percentage yield, percentage composition by mass and percentage purity, given appropriate data

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

AND %purity

- Example 2: 0.8g of hydrogen is produced from 12g of magnesium when it reacts with HCl. Calculate the % yield.
 - Step 1: write the equation. $\text{Mg} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$.
 - Step 2: get the number of moles of magnesium.
 $12/24(\text{Ar}) = 0.5$ moles.
 - Step 3: get the moles of hydrogen using mole ration. For every mole of magnesium, one mole of hydrogen is produced. The ratio is 1:1. This gives 0.5 moles of hydrogen.
 - Step 4: get the theoretical yield in grams. Divide 0.5 moles by 2(Mr). this gives a mass of 1g.
 - Get the % yield. $(0.8\text{g}/1.0\text{g}) \times 100$ which is 80% yield.
- Example 2: find the % composition of hydrogen and oxygen in water.
 - Step 1: Get the Mr of water and the Ar hydrogen and oxygen. $\text{Mr}(\text{H}_2\text{O}) = 2(1) + 16 = 18$
 - Step 2: divide the Ar by the Mr and *100 to get the %. $(2/18) \times 100$ this gives 11.1% of hydrogen. $(16/18) \times 100$ gives 89.9% of oxygen.
- Example 3: find % purity of a 10g sample of calcium carbonate which gives 3.3g of carbon dioxide when heated.
 - Step 1: Write the equation: $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
 - step 2: get the molar masses of calcium and carbon dioxide. For calcium it is 100g/mol(40+12+3(16)) and for CO_2 it is 44g/mol(12+2(16))

- Step 3: find the simplest ratio of moles. 100g(1mol) of calcium carbonate gives 44g(1mol) of carbon dioxide. So 10g of calcium carbonate should give 4.4g of carbon dioxide.
- Step 4: use the actual yield(4.4g) to get the % purity. $(3.3\text{g}/4.4\text{g}) \times 100 = 75\%$ purity.
- Example 3: A sample of pure NaCl weighs 9g. Another sample has impurities weighs 12g. calculate the % purity of the impure sample
- Use the formula:

$$\% \text{ purity} = \frac{\text{mass of pure substance}}{\text{mass of impure substance}} \times 100$$

- $(9/12) \times 100$ gives 75% purity.

END OF CHAPTER