Your notes

GCSE Chemistry: revision notes

Relative formula mass and moles

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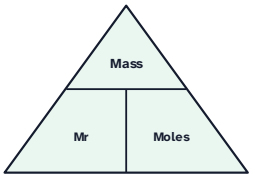
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Start here: counting atoms by weighing them (Higher)

- Atoms are far too small to count one by one, so chemists count them in huge batches called **moles**.
- One **mole** of any substance contains the same number of particles: 6.02×10^{23}, the **Avogadro constant**.
- One **mole** has a mass equal to its relative formula mass in grams: a **mole** of carbon (Ar 12) is 12 g, and a **mole** of water (Mr 18) is 18 g. So moles = mass ÷ Mr.

The mole (Higher)

- Chemical amounts are measured in **moles** (mol).
- One **mole** of a substance contains the same number of particles as one **mole** of any other substance: the **Avogadro constant**, 6.02×10^{23} per mole.
- The mass of one **mole** of a substance in grams equals its relative formula mass (Mr) or relative atomic mass (Ar). For example, one **mole** of CO_2 has a mass of 44 g.
- Number of **moles** = mass (g) ÷ Mr.



Mass = Mr × Moles

Mr = Mass ÷ Moles

Moles = Mass ÷ Mr

Cover the one you want: what's left shows how to work it out.

Reacting masses (Higher)

- A balanced equation shows how many **moles** react. $2Mg + O_2 \rightarrow 2MgO$ means 2 mol of Mg react with 1 mol of O_2 to make 2 mol of MgO.
- To find a mass: work out **moles** of the known substance, use the equation's ratio to find **moles** of the other, then multiply by its Mr.
- Example: 12 g of Mg is 12 ÷ 24 = 0.5 mol, which makes 0.5 mol of MgO = 0.5 × 40 = 20 g.

Balancing equations from masses (Higher)

- Convert the masses of reactants and products to **moles**, then divide by the smallest number of **moles** to get the simplest whole-number ratio.
- Example: 4.8 g Mg, 3.2 g O_2 and 8.0 g MgO are 0.2, 0.1 and 0.2 mol, a ratio of 2 : 1 : 2, so $2Mg + O_2 \rightarrow 2MgO$.

Limiting reactants (Higher)

- In a reaction, one reactant may be completely used up. It is the **limiting reactant**, because it limits the amount of product that can form.
- The other reactant is **in excess**: some of it is left over.
- The mass of product depends on the amount of the **limiting reactant**.

Volumes of gases (Higher, separate sciences)

- One **mole** of any gas occupies 24 dm^3 at room temperature and pressure (20 °C and 1 atmosphere).
- Volume of gas (dm^3) = **moles** × 24. For example, 0.25 mol of any gas occupies 6 dm^3 at RTP.

Empirical formulae (Edexcel and OCR)

- The **empirical formula** is the simplest whole-number ratio of atoms of each element in a compound. For example, the **empirical formula** of glucose, $C_6H_{12}O_6$, is CH_2O.
- To find it from masses: divide the mass of each element by its Ar to get **moles**, then divide by the smallest number of **moles** to get the simplest whole-number ratio.
- Example: 2.4 g of magnesium reacts with 1.6 g of oxygen. Mg: 2.4 ÷ 24 = 0.1 mol; O: 1.6 ÷ 16 = 0.1 mol. The ratio is 1 : 1, so the **empirical formula** is MgO.
- (Edexcel) Core practical: heat a weighed piece of magnesium ribbon in a crucible, lifting the lid now and then, until it stops glowing. Reweigh to find the mass of oxygen that combined, then work out the **empirical formula**.
- The molecular formula is a whole-number multiple of the **empirical formula**: divide the Mr by the mass of the **empirical formula** to find the multiple.

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Exam tip

- **Moles** = mass ÷ relative formula mass. Show every step, and go through **moles** for any reacting mass question.
- Use the ratio in the balanced equation before you turn **moles** back into a mass.

Common mistakes

- Using one atom's mass rather than the whole formula's.
- Forgetting to divide by 1000 to turn cm^3 into dm^3.
- Skipping the ratio from the equation.

Worked examples

Find the relative formula mass (Mr) of water, H_2O. (Ar: H = 1, O = 16)

Add up the relative atomic masses of every atom in the formula. **2 × 1 + 16**

Answer: 18

How many moles are in 22 g of carbon dioxide, CO_2? (Mr = 44)

Use moles = mass ÷ Mr. **n = 22 ÷ 44**

Answer: 0.5 mol

What is the mass of 0.25 mol of sodium hydroxide, NaOH? (Mr = 40)

Rearrange: mass = moles × Mr. **m = 0.25 × 40**

Answer: 10 g

Key terms

Mole	The unit for amount of substance; one mole contains 6.02×10^{23} particles.
Avogadro constant	The number of particles in one mole: 6.02×10^{23}.
Molar mass	The mass of one mole of a substance in grams, equal to its Mr.
Limiting reactant	The reactant that is completely used up, limiting the amount of product.
In excess	More than enough of a reactant; some is left over.
RTP	Room temperature and pressure: 20 °C and 1 atmosphere.
Molar gas volume	The volume of one mole of any gas at RTP: 24 dm^3.
Empirical formula	The simplest whole-number ratio of atoms of each element in a compound. (Edexcel and OCR)

Your notes

Quick questions

Cover the answers and test yourself: remembering is what makes it stick.

1. What is the Avogadro constant?
2. How many moles are in 36 g of water (Mr 18)?
3. What is the mass of 0.5 mol of carbon dioxide (Mr 44)?
4. How many moles are in 4 g of sodium hydroxide, NaOH (Mr 40)?
5. In $2Mg + O_2 \rightarrow 2MgO$, how many moles of MgO form from 1 mol of Mg?
6. What is a limiting reactant?
7. What volume does 2 mol of any gas occupy at RTP?
8. How many moles of oxygen are in 6 dm^3 of the gas at RTP?

Answers

1. 6.02×10^{23}
2. 2 mol
3. 22 g
4. 0.1 mol
5. 1 mol
6. The reactant that is completely used up
7. 48 dm^3
8. 0.25 mol

What to do next

Practise 13 questions on this topic

Try 8 exam questions, with mark schemes

Read the lesson, one step at a time

Free revision for GCSE and KS3 at lycolearn.co.uk. Some notes are for one exam board only, and say so in brackets.

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