

CIE Chemistry A Level

1 : Atoms, Molecules and Stoichiometry

Notes



Relative atomic masses of atoms and molecules

- **Relative atomic mass:** the weighted average mass of an atom of an element compared with 1/12 of the mass of an atom of carbon-12.
- **Relative isotopic mass:** the mass of an atom of an isotope as compared to 1/12 of the mass of an atom of carbon-12.
- **Relative molecular mass:** the weighted average mass of one molecule of an element or compound compared with 1/12 of the mass of an atom of carbon-12.
- **Relative formula mass:** the weighted average mass of one unit of a substance compared with 1/12 of the mass of an atom of carbon-12.

The mole and the Avogadro constant

One mole is the **amount** of any substance containing as many particles as there are carbon atoms in **exactly 12 g of carbon-12**. One mole of any substance will always contain exactly the same number of particles. **Avogadro's constant** is the **number of particles in one mole** of any substance. This is equal to 6.02×10^{23} .

Therefore:

$$\text{Number of particles/atoms/molecules} = \text{Avogadro's constant} \times \text{moles}$$

Example

How many atoms are there in 0.250 moles of carbon?

$$0.25 \times 6.02 \times 10^{23} = 1.51 \times 10^{23} \text{ atoms}$$

Relative atomic mass

Mass spectrometry is an analytical technique which gives information about the abundance of different elements on a mass spectra.

Below is the mass spectrum of boron. Each line represents a **different isotope** of boron so this shows 2 isotopes were detected. Isotopes have the same number of protons and electrons but a different number of neutrons. A **mass to charge ratio (m/z)** of 10 shows that a +1 ion has a mass of 10. The relative heights of the peaks show the **relative abundance**.

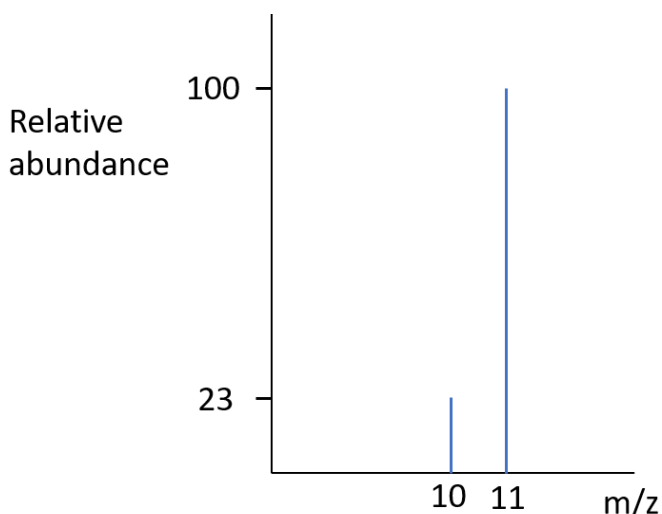
This graph can be used to calculate the isotopic abundance of ^{10}B :

$$23 \div (100 + 23) = 0.187$$

To calculate the relative atomic mass of an atom of boron:

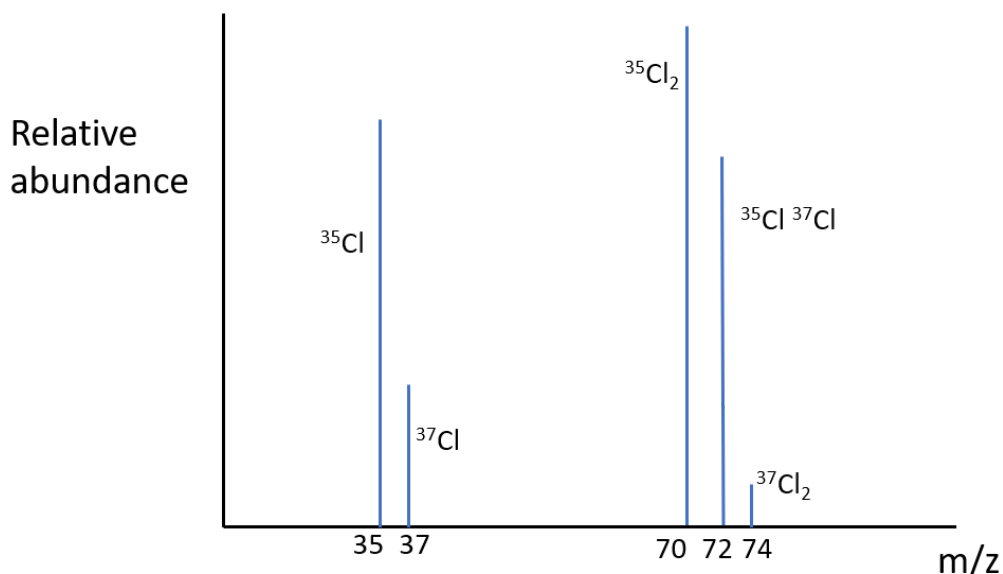
$$(10 \times 23) + (11 \times 100) = 1330$$

$$1330 \div (23 + 100) = 10.8$$



Diatomic elements

Chlorine exists as molecules of Cl_2 rather than as individual atoms. When chlorine is ionised during mass spectrometry, a **molecular ion** is formed (Cl_2^+). As this ion is quite unstable, **fragmentation** may occur, forming Cl and Cl^+ . The Cl atom isn't detected by the mass spectrometer, but Cl_2^+ and Cl^+ are detected. There are 2 isotopes of chlorine so Cl_2^+ may contain ^{35}Cl , ^{37}Cl or both. The mass spectrum will appear like this:



Empirical and molecular formulae

- **Empirical formula:** the simplest whole number ratio of atoms of each element in a compound.
- **Molecular formula:** the actual number of atoms of each element in a compound.

Calculating empirical formula from masses

1. Divide the mass (or percentage by mass) of each element by its molar mass to calculate the molar ratio.
2. Divide each number in the ratio by the smallest number to get the simplest ratio of elements.
3. If the ratio contains decimal numbers, multiply it as appropriate to obtain whole numbers.

Calculating empirical formula from combustion data

If the compound contains **only carbon and hydrogen**:

1. Divide the mass of CO_2 produced during combustion by 44 to find the number of moles of CO_2 . This is equal to the amount of carbon in the compound.
2. Divide the mass of water produced during combustion by 18 to find the number of moles of water. One mole of H_2O contains 2 moles of hydrogen atoms so double the number of moles of water to find the number of moles of hydrogen in the compound.



3. Divide the number of moles of hydrogen and carbon by the smallest value to find the simplest molar ratio. This is the empirical formula.

Example

A sample is burned in oxygen. 0.069 g of CO_2 and 0.0113 g of H_2O is produced. What is the empirical formula of the sample burned?

Moles of carbon in the sample = $(0.069/44) = 0.00157$

Moles of hydrogen in the sample = $(0.0113/18) \times 2 = 0.00126$

Ratio of carbon to hydrogen = 1.25 : 1

Empirical formula ratio = 5 : 4 so the empirical formula is C_5H_4

If the compound contains **other elements** as well as carbon and hydrogen:

1. Use the mass of carbon dioxide to calculate the number of moles of carbon atoms then use this to calculate the mass of carbon present.
2. Use the mass of water to calculate the number of moles of hydrogen atoms then use this to calculate the mass of hydrogen present.
3. Add the masses of carbon and hydrogen and compare to the initial mass of the compound. If there is no difference, the compound is a hydrocarbon so continue with step 4. If there is a difference, **assume that this is due to oxygen** (unless otherwise told). Calculate the mass of oxygen present and use this to calculate the number of moles of oxygen atoms.
4. Divide the moles of each element by the smallest value to obtain the empirical formula.

Calculating molecular formulae

Calculate the relative formula mass of the empirical formula then divide the relative formula mass of the compound by this calculated value. This will show how many times larger the molecular formula is than the empirical formula.

Example

A compound has the empirical formula HO and the relative formula mass is 34. What is the molecular formula?

Empirical formula mass = $16 + 1 = 17$

$34/17 = 2$ so the molecular formula is H_2O_2 .

Reacting masses and volumes

When completing calculations in chemistry, quantities should be given to the **same number of significant figures** as the **least accurate measured quantity**. So if the values are given to 2 significant figures then the answer should also be given to 2 significant figures.

Calculations involving reacting masses:

Molar mass is the mass per mole of a substance and can be calculated by **adding the relative atomic masses** of all the atoms in a formula. To calculate the number of moles using mass and molar mass:

$$\text{Number of moles} = \text{mass (g)} \div \text{molar mass}$$



Calculations involving volumes of gases:

At the **same temperature and pressure**, one mole of any gas will occupy the **same volume**. At room temperature and pressure, this is 24 dm³. This does not mean that these gases will have the same mass. The number of moles of a gas at room temperature and pressure can be calculated, where V is volume:

$$\text{Moles of gas} = \text{Volume (dm}^3\text{)} \div 24$$

Calculations involving concentrations and volumes of solutions:

Concentration is the **amount of solute** dissolved in a **given volume of solution**. The number of moles can be calculated from concentration (in mol dm⁻³) and volume (in dm³):

$$\text{Number of moles} = \text{concentration (mol dm}^{-3}\text{)} \times \text{volume (dm}^3\text{)}$$

Stoichiometry

Stoichiometry is the **proportions of reactants and products**. A balanced chemical equation can be used to find the stoichiometry of a reaction.

