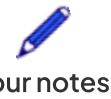




# Cambridge (CIE) A Level Chemistry



Your notes

## Relative Masses of Atoms & Molecules

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# Relative Masses

## Atomic Mass Unit

- The mass of a single atom is so small that it is impossible to weigh it directly
- Atomic masses are therefore defined in terms of a **standard** atom which is called the **unified atomic mass unit**
- This unified atomic mass is defined as **one-twelfth** of the mass of a carbon-12 isotope
- The symbol for the unified atomic mass is ***u*** (often *Da*, Dalton, is used as well)
- $1u = 1.66 \times 10^{-27} \text{ kg}$

## Relative atomic mass, $A_r$

- The **relative atomic mass** ( $A_r$ ) of an element is the **ratio** of the average mass of the atoms of an element to the **unified atomic mass unit**
- The relative atomic mass is determined by using the **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

$$\text{Relative atomic mass of } X = \frac{\text{average mass of one atom of } X}{\text{one twelfth of the mass of one carbon - 12 atom}}$$

## Relative isotopic mass

- The **relative isotopic mass** is the mass of a particular atom of an **isotope** compared to the value of the **unified atomic mass unit**
- Atoms of the same element with a different number of neutrons are called **isotopes**
- Isotopes** are represented by writing the **mass number** as  $^{20}\text{Ne}$ , 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

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

## Relative molecular mass, $M_r$



Your notes

- The **relative molecular mass** ( $M_r$ ) is the **ratio** of weighted average mass of a molecule of a molecular compound to the **unified atomic mass unit**
- The  $M_r$  has **no units**

$$M_r = \frac{\text{weighted average mass of molecules in a given sample of a molecular compound}}{\text{unified atomic mass unit}}$$

- 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**
  - Eg. silicon dioxide has a giant covalent structure, however the simplest formula (the **formula unit**) is  $\text{SiO}_2$

## Example $M_r$ calculations

Substance	Atoms present	$M_r$
Hydrogen ( $\text{H}_2$ )	$2 \times \text{H}$	$(2 \times 1.0) = 2.0$
Water ( $\text{H}_2\text{O}$ )	$(2 \times \text{H}) + (1 \times \text{O})$	$(2 \times 1.0) + (1 \times 16.0) = 18.0$
Potassium carbonate ( $\text{K}_2\text{CO}_3$ )	$(2 \times \text{K}) + (1 \times \text{C}) + (3 \times \text{O})$	$(2 \times 39.1) + (1 \times 12.0) + (3 \times 16.0) = 138.2$
Calcium hydroxide ( $\text{Ca}(\text{OH})_2$ )	$(1 \times \text{Ca}) + (2 \times \text{O}) + (2 \times \text{H})$	$(1 \times 40.1) + (2 \times 16.0) + (2 \times 1.0) = 74.1$
Ammonium sulfate ( $(\text{NH}_4)_2\text{SO}_4$ )	$(2 \times \text{N}) + (8 \times \text{H}) + (1 \times \text{S}) + (4 \times \text{O})$	$(2 \times 14.0) + (8 \times 1.0) + (1 \times 32.1) + (4 \times 16.0) = 132.1$

## 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 sulfates are relative formula masses