1 Moles and formulas – Part 1

Name: _____ Class: _____ Date: _____

Vocabulary

Write each word three times.

English	**Write it 3 times** (English)		
moles			
Avogadro's number			
molar mass			
Isotopes			
relative abundance			
empirical formula			

Recall

You have already met the building blocks of chemistry:

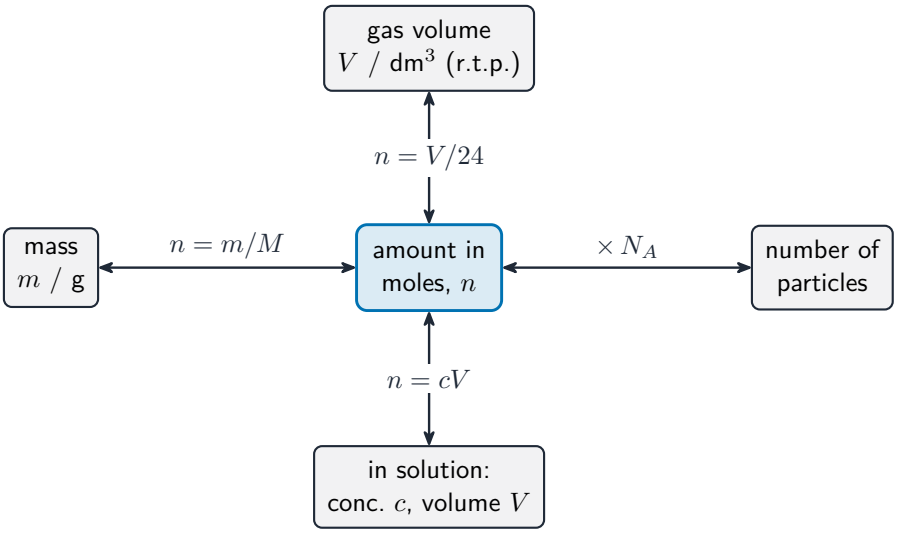
- Atoms are far too small to count one by one.
- Every element has a **relative atomic mass** written on the periodic table.
- You have written formulas like H_2O and CO_2.

This sheet lines up the "counting" tool chemists use – the mole. Each idea has a worked example.

New ideas

1 The mole

Because atoms are so tiny, chemists count them in **moles** . One mole is **Avogadro's number** of particles, $N_A = 6.022 \times 10^{23}$.



2 Molar mass

The **molar mass** M (grams per mole) comes straight from the periodic table. It converts between mass and moles:

$$n = \frac{m}{M}.$$

Worked example. 36.0 g of water ($M = 18.0$ g/mol) is $n = \dfrac{m}{M} = \dfrac{36.0}{18.0} = 2.0$ mol, which is $2.0 \times 6.022 \times 10^{23} = 1.2 \times 10^{24}$ molecules.

3 Isotopes and relative atomic mass

Isotopes are atoms of the same element with different numbers of neutrons. The relative atomic mass is the **weighted average** of the isotope masses, using their **relative abundance** .

Worked example. Chlorine is 75.8% chlorine-35 and 24.2% chlorine-37: $A_r = (35)(0.758) + (37)(0.242) = 35.5$ – closer to 35 because it is more abundant.

4 Empirical formula

The **empirical formula** is the simplest whole-number ratio of atoms. Find it by turning each element's mass into moles and dividing by the smallest.

Worked example. A compound of 40.0% C, 6.7% H, 53.3% O: C $= \frac{40.0}{12} = 3.33$, H $= \frac{6.7}{1} = 6.7$, O $= \frac{53.3}{16} = 3.33$; dividing by 3.33 gives 1 : 2 : 1, so CH_2O.

Watch out: the relative atomic mass is a **weighted** average – it sits closer to the more abundant isotope, not exactly halfway between the two.

exercises.lol exercises.com.cn

Practice

Try these in order. The methods are all above.

2.1 How many moles are in 22.0 g of carbon dioxide ($M = 44.0$ g/mol)? [2]

2.2 How many molecules are in 0.50 mol of a gas? ($N_A = 6.022 \times 10^{23}$) [2]

2.3 An element is 60% of an isotope of mass 20 and 40% of an isotope of mass 22. Find its relative atomic mass. [3]

2.4 Find the mass of 3.0 mol of water ($M = 18.0$ g/mol). [2]

2.5 A compound is 50% sulfur and 50% oxygen by mass ($S = 32$, $O = 16$). Find its empirical formula. [3]