UNIVERSITY OF LEICESTER

DEPARTMENT OF CHEMISTRY

REVISION AID

Moles

Physical Constants

Name	Common Symbol	Value
Avogadro constant	N_A	6.022142×10^{23} mol^{-1}
Boltzmann constant	k, k_B	1.380650×10^{-23} J K^{-1}
Elementary charge	e	1.602176×10^{-19} C
Faraday constant	F	9.648533×10^{4} C mol^{-1}
Gas constant	R	8.314472 J K^{-1} mol^{-1}
Mass of an electron	m_e	9.109382×10^{-31} kg
Planck constant	h	6.626069×10^{-34} J s
Speed of light	c	2.99792458×10^{8} m s^{-1}
Vacuum permittivity	ε_o	8.854188×10^{-12} C^2 J^{-1} m^{-1}

University of Leicester
Department of Chemistry
University Road
Leicester, LE1 7RH, UK
t: +44 (0)116 252 2100
e: chemadmin@le.ac.uk
w: www.le.ac.uk/chemistry

/Leicesterchem
uniofleicester
@Leicesterchem

Atomic Mass

The modern system of atomic masses is based on ^{12}C (carbon 12). In this system, one atomic mass unit is assigned a mass of exactly 1/12th of ^{12}C.

The relative atomic mass, A_r of an element is the average mass of the naturally occurring isotopes of the element relative to the mass of an atom of ^{12}C.

The relative molecular mass, M_r, is the mass of a molecule relative to the mass of an atom of ^{12}C.

e.g. A_r (H) = 1, A_r (O) = 16, M_r (CH_4) = 16, M_r (NH_3) = 17

Relative masses have NO UNITS.

The Mole

Real samples contain very large numbers of atoms, molecules or ions. To simplify matters, a special unit is defined called **the mole**. This is defined as the amount of substance that contains the same number of particles as there are atoms in 12 g of ^{12}C.

The number of atoms in 12 g of ^{12}C is known as the **Avogadro number (N_A)**.

$N_A = 6.022142 \times 10^{23}$ mol^{-1}

Thus one mole of any substance contains 6.022142×10^{23} separate atoms, molecules or formula units of that substance.

e.g. one mole of O_2 contains 6.022142×10^{23} oxygen molecules

one mole of Al contains 6.022142×10^{23} aluminium atoms

Never forget that the mole is just a convenient unit, like 'a dozen'.

A dozen eggs is 12 eggs
A mole of eggs is 6.022142×10^{23} eggs

The mole is the S.I. unit for the amount of a substance and has the symbol **mol**.

Molar Mass

By definition, the mass of one mole of particles is equal to the relative molecular mass of the substance in grams. Therefore, we can define a new quantity, Molar Mass, which is equal to the mass of one mole of a substance in grams (units: g mol^{-1}).

Empirical and Molecular Formulae

The molecular formula of a compound shows the number of atoms of each element that are present in one molecule of the compound.

e.g. molecular formula of methane = CH_4
molecular formula of ethanoic acid = $C_2H_4O_2$

The empirical formula of a compound is the formula that shows the simplest ratio of atoms of each element in that compound.

e.g. empirical formula of methane = CH_4
empirical formula of ethanoic acid = CH_2O

If the masses or percentage mass of each element present in the compound are known, the empirical formula of the compound can be calculated. If the relative molecular mass is also known, the molecular formula can be calculated.

Calculating an empirical formula

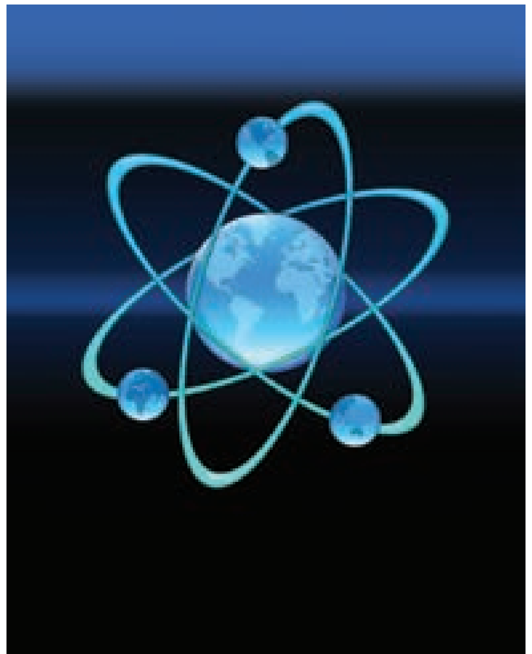
1) Draw a table with the following headings:

1	2	3	4	5	6
Element	Mass or % mass	Relative atomic mass (A_r)	Mass or % mass / A_r	Test ratio	Simplest ratio

2) Enter the data given into the first two columns.
3) Find the relative atomic masses for the elements and enter them in the third column.
4) Perform the calculation in the fourth column (i.e. divide the value in column 2 by that in column 3).
5) To determine the test ratio, take the smallest number in column 4 and divide each of the column 4 values by this number – write the answers in column 5.
6) Look at the numbers in column 5. Check that they are all whole numbers. If so, copy them in to column 6. If any of them are not whole numbers, multiply all of them by the lowest number needed to make them all whole numbers and write these in column 6.
7) Write out the empirical formula by using the numbers in column 6.
Example – A hydrocarbon was found to contain 83.72% carbon by mass. Calculate the empirical formula.

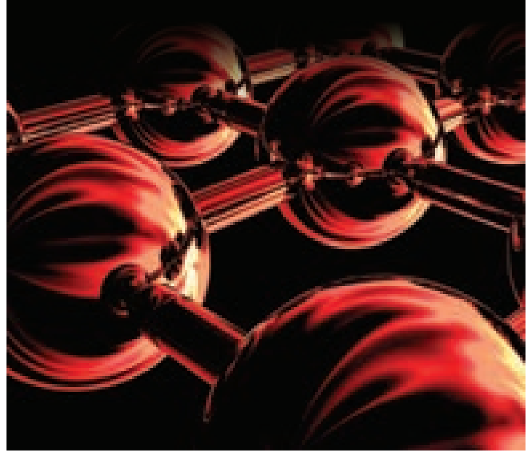
1	2	3	4	5	6
Element	Mass or % mass	Relative atomic mass (A_r)	Mass or % mass / A_r	Test ratio	Simplest ratio
C	83.72	12	6.977	1	3
H	16.28	1	16.28	2.333	7

Thus the empirical formula is C_3H_7.



Calculating a molecular formula from an empirical formula and a relative mass.

1) Calculate the mass of your empirical formula (e.g. C_3H_7 = 36+7 = 43)
2) Divide the relative mass given by the mass calculated in 1). (e.g. if the relative molecular mass is 86, divide 86 by 43 – giving 2)
3) The number calculated gives the scaling factor for each atom in the formula (thus for the example the molecular formula would be $C_{2x3}H_{2x7}$ = C_6H_{14}).



Reacting Masses

When chemical reactions take place, molecules react in the ratios given by the chemical equation. For instance, in the reaction:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Two molecules of hydrogen react with one molecule of oxygen to make two molecules of water. As a mole contains exactly the Avogadro number of molecules, it is also true that two moles of hydrogen react with one mole of oxygen to make two moles of water.

In any situation, the number of moles of a substance can be calculated from its mass and molar mass, and vice versa.

Therefore, the masses of a substance that react with each other can be calculated.

Example

7 g of hydrogen is burnt completely in oxygen. What mass of oxygen is used up and what mass of water is produced?

The following steps allow the answer to be easily calculated.

1) Calculate the number of moles of hydrogen (H_2) using the given mass and the molar mass.

$$\text{Number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{7\text{ g}}{2\text{ g mol}^{-1}} = 3.5\text{ mol}$$

2) Using the chemical equation, determine the corresponding numbers of moles of oxygen used and water produced.

From the chemical equation, two moles of hydrogen react with one mole of oxygen to give two moles of water. Therefore, 3.5 mol of hydrogen must react with 1.75 mol of oxygen to give 3.5 mol of water.

3) Use the number of moles from 2) along with the molar mass to calculate the mass of oxygen used and the mass of water produced.

Oxygen:

Number of moles = 1.75 — Molar mass = 2 x 16 g mol^{-1} = 32 g mol^{-1}

Mass of oxygen used = Number of moles x Molar mass = 1.75 mol x 32 g mol^{-1} = 56 g

Water:

Number of moles = 3.5 — Molar mass = 2 x 1 g mol^{-1} + 16 g mol^{-1} = 18 g mol^{-1}

Mass of water produced = Number of moles x Molar mass = 3.5 mol x 18 g mol^{-1} = 63 g

Therefore, 7 g of hydrogen reacts with 56 g of oxygen to make 63 g of water.

Solutions

When dealing with solutions, a certain amount of solute (i.e. a substance) is dissolved in a certain amount of solvent. Dissolving more solute or using less solvent result in a more concentrated solution. Imagine making a coffee. A normal cup of coffee might use one spoon of coffee granules in one cup of hot water. If two spoons of coffee granules are used, the coffee will be twice as strong. In addition, if half a cup of water is used with one spoon of coffee granules, this coffee will also be twice as strong.

As usual, it is helpful to work with moles, so in chemistry the concentration of a solution is called the MOLARITY and tells us the number of moles of solute per dm^3 (= litre) of solution.

$$\text{Molarity} = \frac{\text{number of moles of solute}}{\text{volume of solution (in dm}^3\text{)}}$$

Example

If 2.8 g of potassium hydroxide (KOH) is dissolved in 100 cm^3 of water, what is the molarity of the solution?

1) Calculate the number of moles of KOH.

$$\text{Number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{2.8\text{ g}}{56\text{ g mol}^{-1}} = 0.05\text{ mol}$$

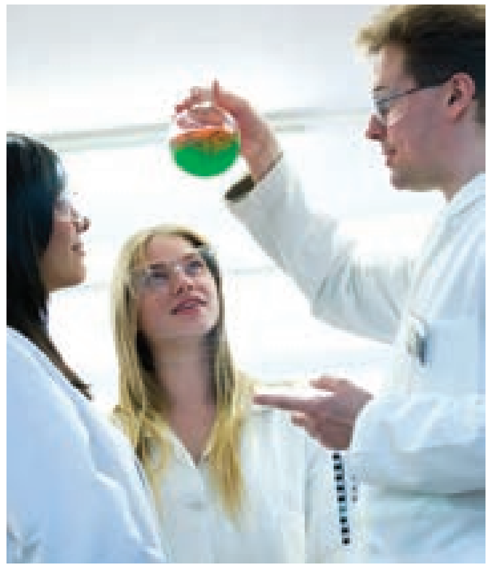
2) Write the volume of the solution in dm^3. Remember that 1 dm^3 = 1000 cm^3.

Volume of solution = 100 cm^3 = 0.1 dm^3

3) Use the equation to calculate the molarity.

$$\text{Molarity} = \frac{\text{number of moles of solute}}{\text{volume of solution (in dm}^3\text{)}} = \frac{0.05\text{ mol}}{0.1\text{ dm}^3} = 0.5\text{ mol dm-3 (or M) mol}$$

The concentration of the KOH solution is 0.5 mol dm^{-3} or 0.5 M.



Reacting Volumes

When working with gases, it is much easier to measure their volumes than their masses. As long as the pressure and temperature are kept the same, equal volumes of gases contain equal numbers of molecules. **This is Avogadro's Law.**

In the laboratory, two sets of conditions are commonly used:

Standard Temperature and Pressure (STP)
Temperature = 273 K (0°C)
Pressure = 101325 Pa

Room Temperature and Pressure (RTP)
Temperature = 298 K (25°C)
Pressure = 101325 Pa

One mole of ANY gas at STP has a volume of 22.4 dm^3
One mole of ANY gas at RTP has a volume of 24.0 dm^3

We can use this to calculate the volume of gases that react together.

Note: The IUPAC definition of standard pressure and room pressure is 100000 Pa.

Example

What reacting volumes of nitrogen and hydrogen are needed to produce 4.8 dm^3 of ammonia at RTP?

1) Write down the balanced chemical equation.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

2) Calculate the number of moles of ammonia in 4.8 dm^3 at RTP.

24.0 dm^3 of ammonia at RTP contains one mole of ammonia molecules. Therefore, 4.8 dm^3 of ammonia at RTP contains {4.8 dm^3}/{24.0 dm^3 mol^{-1}} = 0.2 mol of ammonia molecules.

3) Use the equation to calculate the number of moles of nitrogen and hydrogen required.

From the equation, one mole of nitrogen reacts with three moles of hydrogen to make two moles of ammonia. Therefore, 0.1 mol of nitrogen reacts with 0.3 mol of hydrogen to make 0.2 mol of ammonia.

4) Calculate the volumes of nitrogen and hydrogen required at RTP.

One mole of nitrogen has a volume of 24.0 dm^3 at RTP.

Therefore, 0.1 mol of nitrogen has a volume of 0.1 mol × 24.0 dm^3 mol^{-1} = 2.4 dm^3.

One mole of hydrogen has a volume of 24.0 dm^3 at RTP.

Therefore, 0.3 mol of hydrogen has a volume of 0.3 mol × 24.0 dm^3 mol^{-1} = 7.2 dm^3.

2.4 dm^3 of nitrogen and 7.2 dm^3 of hydrogen are needed to produce 4.8 dm^3 of ammonia at RTP.

Titrations

In titrations, we use a solution of known concentration **(A)** to find the unknown concentration of another solution **(B)** based on a chemical reaction between the two solutions. During the titration, the volumes of **A** and **B** that react together are accurately determined. One is fixed, usually by using a pipette and the other is varied, using a burette.

Example

25.00 cm^3 of a sodium hydroxide solution of unknown concentration are titrated with 0.100 M (mol dm^{-3}) sulfuric acid solution. The endpoint occurs when 23.95 cm^3 of the acid have been added. Determine the concentration of the sodium hydroxide solution.

Procedure

1) Summarise the details of the titration.

Solution **A**: Sulfuric acid, $H_2SO_4(aq)$: 23.95 cm^3 of a 0.100 M solution

Solution **B**: Sodium hydroxide, NaOH(aq) : 25.00 cm^3 of a solution of unknown concentration

2) Write down the balanced chemical equation for the neutralisation reaction.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

3) Calculate the number of moles of solution A.

Number of moles of A = volume of solution A (in dm^3) x molarity of solution A (mol dm^{-3})

Number of moles of A = 0.02395 dm^3 x 0.100 mol dm^{-3}
= 2.395 x 10^{-3} mol

4) Use the equation to relate the number of moles of A to the number of moles of B.

From the chemical equation, one mole of H_2SO_4 **(A)** reacts with two moles of NaOH **(B)**.

Therefore, 2.395 × 10^{-3} mol of H_2SO_4 (A) reacts with 4.790 × 10^{-3} mol of NaOH (B).

5) Use the number of moles of B along with its volume to calculate the concentration of B.

25.00 cm^3 = 0.025 dm^3

0.025 dm^3 of **B** contains 4.790 × 10^{-3} mol of NaOH.
The concentration of **B** is 4.790 × 10^{-3} mol / 0.025 dm^3 = 0.1916 mol dm^{-3}.

In the question, the lowest precision number is the concentration of sulfuric acid (3 significant figures) so this precision should be used for the final answer.

The concentration of the solution of NaOH is 0.192 mol dm^{-3} or 0.192 M.