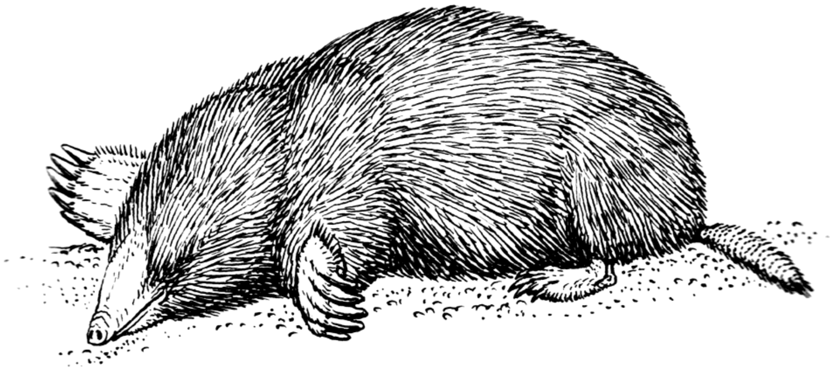
UNIT 3

AMOUNT OF SUBSTANCE

PART 2 – MEASURING THE AMOUNT OF SUBSTANCE



Contents

Key words: mole, Avogadro's number, molar mass, molar concentration, mass concentration, Avogadro's Law, ideal gas equation, molar volume of gas, empirical formula

Units which must be completed before this unit can be attempted:
Unit 1 – Atomic Structure and the Periodic Table
Unit 2 – Particles, Structure and Bonding

1) Introducing the Mole

(a) Avogadro's number

Atoms and molecules are very small – far too small to count individually! Since they are so small, any sensible laboratory quantity of substance must contain a huge number of individual particles:
1 litre of water contains 3.3 x 10^{25} molecules.
1 gram of magnesium contains 2.5 x 10^{22} atoms.
100 cm^3 of oxygen contains 2.5 x 1021 molecules.

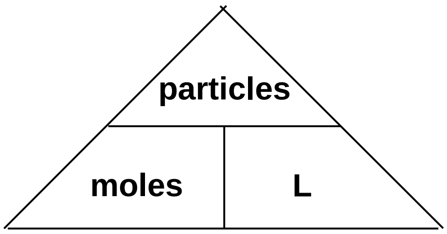
Such numbers are not convenient to work with, but it very important in Chemistry to be able to measure the amount of substance present. It is therefore necessary to find a unit of "amount" which corresponds better to the sort of quantities of substance normally being measured. The unit chosen for this purpose is the **mole**.

The number of particles in one mole of a substance is 6.02 x 10^{23}. This is known as **Avogadro's number, L**.

When we need to know the number of particles of a substance, we usually count the number of moles. It is much easier than counting the number of particles.

The number of particles can be calculated by multiplying the number of moles by Avogadro's number. The number of moles can be calculated by dividing the number of particles by Avogadro's number.

(Number of particles) = (number of moles) x L



Test Your Progress: Topic 3 Part 2 Exercise 1

a) If you have 2.5 x 10^{21} atoms of magnesium, how many moles do you have?
b) If you have 0.25 moles of carbon dioxide, how many molecules do you have?
c) How many moles are present in 3 x 10^{22} molecules of nitrogen?
d) How many atoms of carbon are present in 0.02 moles?

(b) Relative molecular and formula masses

In giant metallic, giant covalent and simple atomic elements, the formula of the substance is simply the symbol of the atom is made from (eg magnesium = Mg; silicon = Si; carbon = C; argon = Ar)

Atoms have a relative atomic mass, which is given in the Periodic Table. It is defined as "the average mass of one atom of that element compared to 1/12th of the mass of one atom of carbon-12".

Simple molecular substances have a **molecular formula**, which is the number atoms of each element in one molecule of the substance (eg water = H_2O; carbon dioxide = CO_2 ; chlorine = Cl_2).

Simple molecular substances therefore also have a **relative molecular mass**, which is the sum of the relative atomic masses of all the atoms in one molecule of that substance. It is defined as "the average mass of one molecule of that substance compared to 1/12th of the mass of one atom of carbon-12".

Giant covalent and giant ionic compounds have a **unit formula**, which is the simplest whole number ratio of atoms (or ions) of each type in the compound (eg sodium chloride = NaCl; magnesium hydroxide = $Mg(OH)_2$; silicon dioxide = SiO_2).

Giant covalent and giant ionic compounds therefore also have a **relative formula mass**, which is the sum of the relative atomic masses of all the atoms in one formula unit of that substance. It is defined as "the average mass of one formula unit of that substance compared to 1/12th of the mass of one atom of carbon-12".

Test Your Progress: Topic 3 Part 2 Exercise 2

1. Use the Periodic Table of Elements to write down the relative atomic masses of:
   (a) Carbon (b) Oxygen (c) Chlorine (d) Sodium (e) Hydrogen (f) Magnesium
2. Use your answers from question (1) to deduce the relative molecular masses of:
   (a) O_2 (b) CO_2 (c) Cl_2 (d) HCl (e) CH_4 (f) H_2O
3. Use your answers from question (1) to deduce the relative formula masses of:
   (a) NaCl (b) Na_2CO_3 (c) MgO (d) $MgCl_2$ (e) $Mg(OH)_2$

(c) Definition of a mole and molar masses

The number of particles in a mole (Avogadro's number) is chosen so that 1 mole of a substance corresponds to its relative mass measured in grams.

One mole of carbon has a mass of 12.0 g.
One mole of hydrogen atoms has a mass of 1.0 g.
One mole of hydrogen molecules has a mass of 2.0 g.
One mole of sodium chloride has a mass of 58.5 g.

A mole is thus defined as follows:

A mole of a substance is the amount of that substance that contains the same number of elementary particles or formula units as there are carbon atoms in 12.00000 grams of carbon-12.

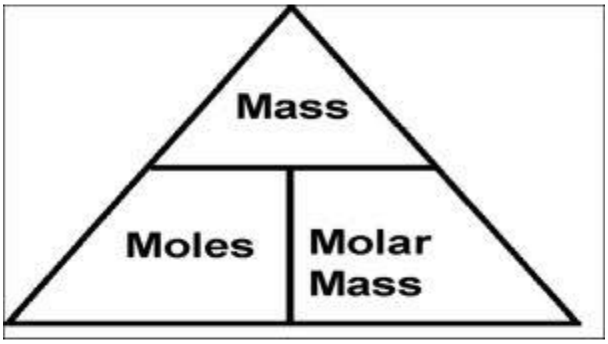
The mass of one mole of a substance is known as its **molar mass**, and has units of $gmol^{-1}$. The symbol for the molar mass of compounds or molecular elements is m_r. The symbol for molar mass of atoms is a_r.

Relative atomic, molecular and formula masses have the same numerical value but have no units.

You can find the number of moles of a substance if you are given its **mass** and you know its **molar mass**:

number of moles = mass/molar mass

$$n = m/m_r$$



Test Your Progress: Topic 3 Part 2 Exercise 3		
1. Calculate the number of moles present in:	2. Calculate the mass of:	3. Calculate the molar mass of the following substances:
a) 2.3 g of Na	a) 0.05 moles of Cl_2	a) 0.015 moles, 0.42 g
b) 2.5 g of O_2	b) 0.125 moles of KBr	b) 0.0125 moles, 0.50 g
c) 240 kg of CO_2	c) 0.075 moles of $Ca(OH)_2$	c) 0.55 moles, 88 g
d) 12.5 g of $Al(OH)_3$	d) 250 moles of Fe_2O_3	d) 2.25 moles, 63 g
e) 5.2 g of PbO_2	e) 0.02 moles of $Al_2(SO_4)_3$	e) 0.00125 moles, 0.312 g

2) Amount of Substance in Solution

(a) Moles, volume and concentration

The amount of substance present in a given volume of solution is known as the concentration of the solution. Concentration is usually measured in moles per cubic decimetre ($moldm^{-3}$). This type of concentration measurement is also known as the **molar concentration** or **molarity** of the solution.

Concentration can also be measured in grams per cubic decimetre (gdm^{-3}). This type of concentration measurement is also known as the **mass concentration** of the solution.

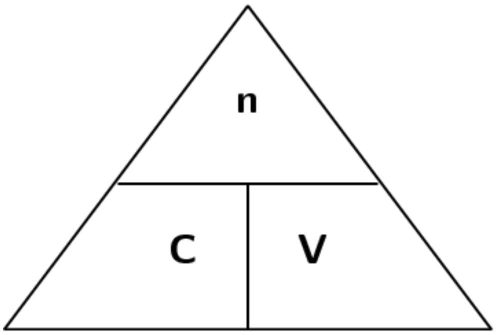
A solution with a relatively high concentration is said to be concentrated. A solution with a relatively low concentration is said to be dilute.

Eg a typical solution of hydrochloric acid used in the laboratory will have a molarity of 0.1 $moldm^{-3}$. Concentrated hydrochloric acid has a molarity of 12 $moldm^{-3}$.

You can find the number of moles of a substance dissolved in water (aqueous) if you are given the **volume** of solution and you know its **molar concentration**:

number of moles = aqueous volume x molar concentration

n = V x C



Aqueous volume MUST be measured in dm^3
concentration has units of $moldm^{-3}$

If you know the molar mass of the substance, you can convert the molar concentration into a mass concentration:

Molar concentration ($moldm^{-3}$) x m_r = mass concentration (gdm^{-3})

Test Your Progress: Topic 3 Part 2 Exercise 4		
1. Calculate the number of moles of substance present in each of the following solutions:	2. Calculate the molar concentration and the mass concentration of the following solutions:	3. Calculate the molar concentration and the mass concentration of the following solutions:
a) 25 cm^3 of 0.1 $moldm^{-3}$ HCl	a) 0.05 moles of HCl in 20 cm^3	a) 35 g of NaCl in 100 cm^3
b) 40 cm^3 of 0.2 $moldm^{-3}$ HNO_3	b) 0.01 moles of NaOH in 25 cm^3	b) 20 g of $CuSO_4$ in 200 cm^3
c) 10 cm^3 of 1.5 $moldm^{-3}$ NaCl	c) 0.002 moles of H_2SO_4 in 16.5 cm^3	c) 5 g of HCl in 50 cm^3
d) 5 cm^3 of 0.5 $moldm^{-3}$ $AgNO_3$	d) 0.02 moles of $CuSO_4$ in 200 cm^3	d) 8 g of NaOH in 250 cm^3
e) 50 cm^3 of 0.1 $moldm^{-3}$ H_2SO_4	e) 0.1 moles of NH_3 in 50 cm^3	e) 2.5 g of NH_3 in 50 cm^3

(b) Standard solutions

A standard solution is a solution of precisely known concentration. Standard solutions can be prepared by dissolving a measured mass of solute to make a specific volume of solution. Standard solutions are important in chemical analysis.

(i) Preparing standard solutions:

A standard solution can be prepared using the following steps:
- weigh out the required mass of solute
- pour the solid into a beaker
- add enough water to dissolve the solid
- transfer the solution to a volumetric flask
- make up to the mark with distilled water

The required mass is calculated as follows:
- choose the concentration and volume of standard solution to be prepared, and hence calculate the moles of solute required
  eg If you want to prepare 250 cm^3 of 0.2 $moldm^{-3}$ NaOH, you will need 250/1000 x 0.2 = 0.05 moles of NaOH
- use the moles required to calculate the mass required
  eg mass of NaOH required = moles x molar mass = 0.05 x 40 = 2.0 g
  So the required mass of NaOH would be 2.0 g

Practical: prepare 250 cm^3 each of 0.1 $moldm^{-3}$ standard solutions of NaOH, NaCl, Na_2CO_3, sugar ($C_{12}H_{22}O_{11}$) and hydrated ethanedioic acid (m_r = 118) (CAUTION – NaOH is highly corrosive)
1) Deduce the mass of each solid required to prepare 250 cm^3 of a 0.1 $moldm^{-3}$ solution
2) Weigh out the required amount of each solid and pour it into a beaker
3) Add enough water to completely dissolve the solid, and then transfer the solution into a volumetric flask
4) Add distilled water until the base of the meniscus lies on the graduated mark on the volumetric flask, shaking well

You can also prepare standard solutions from concentrated solutions, if the percentage by mass of solute in the concentrated solution is known. For example, concentrated HCl is known to contain 36% HCl by mass (so 100 g of concentrated HCl contains 36 g of HCl).

If required to prepare 250 cm^3 of 0.1 $moldm^{-3}$ HCl, the moles required is 250/1000 x 0.1 = 0.025 and so the mass of HCl required is 0.025 x 36.5 = 0.9125 g
The mass of concentrated HCl which contains this mass of pure HCl = 0.9125 x 100/36 = 2.53 g.
So the required mass of concentrated HCl would be 2.53 g.

Practical: prepare 250 cm^3 of a 0.1 $moldm^{-3}$ standard solution of HCl
(CAUTION – HCl is highly corrosive)
1) Weigh out 2.53 g of concentrated HCl
2) Add 100 cm^3 of water to a beaker, and then add the HCl and stir
3) Transfer the solution into a volumetric flask and add water until the base of the meniscus lies on the graduated mark on the volumetric flask, shaking well

(ii) Diluting standard solutions:

Sometimes, a standard solution is available but its concentration is too high to be used in a particular experiment. The standard solution therefore needs to be diluted.

There are three steps to diluting standard solutions:
- Choose the volume and concentration of the diluted solution required
- Calculate the **dilution factor** (the ratio of the high concentration to the low concentration)
- Calculate the volume of the concentrated solution which needs to be diluted and made up to the desired volume

Eg a 2.0 moldm^{-3} solution of hydrogen peroxide (H_2O_2) needs to be diluted to make 250 cm^3 of a 0.05 moldm^{-3} solution.

The dilution factor is 2.0/0.05 = 40.
The desired volume of 0.05 moldm^{-3} is H_2O_2 is 250 cm^3, so the volume of 2.0 moldm^{-3} solution of H_2O_2 needed is 250/40 = 6.25 cm^3.
So 6.25 cm^3 of H_2O_2 should be diluted to a volume of 250 cm^3 in order to achieve this dilution.

Practical: prepare 250 cm^3 of a 0.1 moldm^{-3} solution of hydrogen peroxide by diluting a 2.0 moldm^{-3} solution (dilution factor = 20 so volume to be diluted = 250/20 = 12.5 cm^3)
1) Measure 12.5 cm of 2.0 moldm^{-3} H_2O_2 as accurately as possible into a measuring cylinder
2) Add 100 cm^3 of water to a beaker, and then add the H_2O_2 and stir
3) Transfer the solution into a volumetric flask and add water until the base of the meniscus lies on the graduated mark on the volumetric flask, shaking well

3) Avogadro's Law and the Ideal Gas Equation

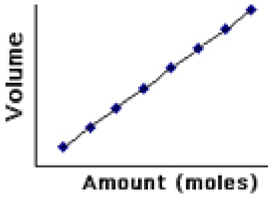
(a) Avogadro's Law

All gases follow the general gas law PV/T = k, at least approximately. A fixed amount of any gas at the same temperature and pressure will have the same volume.

The greater the number of particles, the greater the volume that the gas will occupy at a given temperature and pressure. In fact the volume occupied by a gas is directly proportional to the number of particles at a given temperature and pressure. This is known as **Avogadro's Law**:

Mathematically, $\frac{V}{n} = k$ or $V_2/V_1 = n_2/n_1$

Graphically:



Example: If 0.02 moles of a gas occupy a volume of 0.4 dm^3, what volume will 0.03 moles of gas occupy at the same temperature and pressure?

Answer: $V_1/n_1 = 0.4/0.02 = 20 = V_2/n_2$, so $V_2 = 20 \times 0.03 =$ **0.6 dm^3**

(b) The Ideal Gas Equation

Avogadro's Law can be combined with the general gas law to produce the following:

PV/nT = k.

If P is the pressure measured in pascals (Pa), V is the volume in m^3, T is the temperature measured in Kelvin (K), and n is the number of moles, then the value of the constant is 8.31 $Jmol^{-1}K^{-1}$. It is given the symbol R and is known as the molar gas constant.

The pressure, temperature, volume and amount of gas can be related by a simple equation known as the **ideal gas equation**:

$$\boxed{\mathbf{PV = nRT}}$$

Remember: volume must be measured in m^3: 1 m^3 = 1000 dm^3 = 10^6 cm^3.
Remember: temperature must be measured in Kelvin (K). 0 ^oC = 273 K

Using this equation, it is possible to calculate the volume occupied by a gas at standard temperature (273 K) and pressure (1 atm = 101 kPa) (**stp** = standard temperature and pressure):

V = nRT/P = 1 x 8.31 x 273/101,000 = **0.0224 m^3 = 22.4 dm^3 = 22,400 cm^3**.
This volume is known as the **molar volume of a gas at stp**.

It is also possible to calculate the volume occupied by a gas at room temperature (298 K) and pressure (1 atm = 101 kPa) (**rtp** = room temperature and pressure):

V = nRT/P = 1 x 8.31 x 298/101,000 = **0.0244 m^3 = 24.4 dm^3 = 24,400 cm^3**.
This volume is known as the **molar volume of a gas at rtp**.

Test Your Progress: Topic 3 Part 2 Exercise 5		
1. Calculate the number of moles present in:	2. Calculate the volume of gas occupied by:	3. Calculate the mass of the following gas samples:
a) 48 dm^3 of O_2 at 298 K and 100 kPa	a) 0.05 moles of Cl_2 at 298 K and 100 kPa	a) 48 dm^3 of O_2 at 298 K and 100 kPa
b) 1.2 dm^3 of CO_2 at 298 K and 100 kPa	b) 0.25 moles of CO_2 at 298 K and 100 kPa	b) 1.2 dm^3 of CO_2 at 298 K and 100 kPa
c) 200 cm^3 of N_2 at 273 K and 250 kPa	c) 28 g of N_2 at 273 K and 250 kPa	c) 200 cm^3 of N_2 at 273 K and 250 kPa
d) 100 dm^3 of Cl_2 at 30 ^oC at 100 kPa	d) 3.2 g of O_2 at 30 ^oC at 100 kPa	d) 100 dm^3 of Cl_2 at 30 ^oC at 100 kPa
e) 60 cm^3 of NO_2 at 25 ^oC and 100 kPa	e) 20 g of NO_2 at 25 ^oC and 100 kPa	e) 60 cm^3 of NO_2 at 25 ^oC and 100 kPa

4) Empirical Formulae

(a) Definition of empirical formula

The **empirical formula** of a compound is the formula which shows the simplest whole-number ratio in which the atoms in that compound exist.

The empirical formula is not the same as the **unit formula** of a compound, which is the simplest whole-number ratio in which the particles in that compound exist. In giant covalent substances, the empirical formula and unit formula are always the same, but in ionic compounds they can be different:

Name of ionic compound	Unit formula	Empirical Formula
Sodium oxide	Na_2O (2 Na^+ ions per O^{2-} ion)	Na_2O
Sodium peroxide	Na_2O_2 (2 Na^+ ions per O_2^{2-} ion)	NaO
Aluminium hydroxide	$Al(OH)_3$ (3 OH^- ions per Al^{3+} ion)	AlO_3H_3
Ammonium Nitrate	NH_4NO_3 1 NH_4^+ ion per NO_3^- ion	$N_2H_4O_3$
Magnesium nitrate	$Mg(NO_3)_2$ 2 NO_3^- ions per Mg^{2+} ion	MgN_2O_6

The empirical formula is also not the same as the **molecular formula** of a substance, which is the actual number of atoms of each element in one molecule of that substance:

Name of molecule	Molecular formula	Empirical Formula
Chlorine	Cl_2	Cl
Carbon dioxide	CO_2	CO_2
Ethane	C_2H_6	CH_3
Ethene	C_2H_4	CH_2
Propene	C_3H_6	CH_2

The empirical formula does not uniquely identify a substance, because different substances can have the same empirical formula (eg C_2H_4 and C_3H_6, or NO_2 and N_2O_4).

(b) Determining empirical formulae from mass composition data

If the percentage composition by mass of each element in a compound is known, or the actual mass of each element in a known mass of compound is known, then its empirical formula can be determined by:

- dividing the mass of each element by its molar mass
- dividing each answer by the smallest of the answers
- multiplying all answers by the smallest factor required to ensure that all numbers are whole numbers

Eg If a compound contains 85.8% carbon and 14.2% hydrogen, what is its empirical formula?

Answer: Mole ratio = 85.8/12: 14.2/1 = 7.15:14.2 = 1:2 so **empirical formula = CH_2**

The molecular formula is always a simple multiple of the empirical formula, and can be deduced if the empirical formula and the relative molecular mass are known.

Eg If a compound has the empirical formula CH_2 and a relative molecular mass of 56, what is its molecular formula?

Answer: relative empirical formula mass = 12 + 2 = 14, so there must be 56/14 = 4 empirical formula units in the molecular formula, so the **molecular formula** = $(CH_2)_4$ **= C_4H_8**

Test Your Progress: Topic 3 Part 2 Exercise 6

1. A compound contains C 62.08%, H 10.34% and O 27.58% by mass. Find its empirical formula and its molecular formula given that its relative molecular mass is 58.
2. Find the empirical formula of the compound containing C 22.02%, H 4.59% and Br 73.39% by mass.
3. A compound containing 84.21% carbon and 15.79% hydrogen by mass has a relative molecular mass of 114. Find its molecular formula.
4. Analysis of a hydrocarbon showed that 7.8 g of the hydrocarbon contained 0.6 g of hydrogen and that the relative molecular mass was 78. Find the molecular formula of the hydrocarbon.
5. 3.36 g of iron join with 1.44 g of oxygen in an oxide of iron. What is the empirical formula of the oxide?