Topic 7: Answers
Chemical Calculations

SYLLABUS RELEVANCE & TEXTBOOK CHAPTERS		
O-LEVEL PURE (5072)	✓	Chapters 9 & 10
O-LEVEL SCIENCE (5116)	✓	Chapters 8 & 9
N-LEVEL SCIENCE (5155)	✗	

Lesson Package & Accompanying Slides Designed by Alex Lee (2007)
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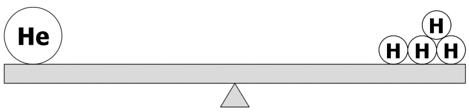
1. What is "Relative Mass"?

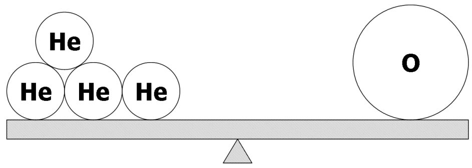
When we look at atomic mass (A_r) and molecular mass (M_r), we realize that there are no physical units, e.g. grams, kilograms, pounds, to these masses! This is because the values are not absolute, but relative to ('in comparison with') each other.

For example, an atom of oxygen-16 has <u>four times</u> the mass of an atom of helium-4.

$A_r = 16$ --- *four times of* ---> $A_r = 4$



A helium atom (A_r = 4) has four times the mass of a hydrogen atom (A_r = 1)



An oxygen atom (A_r = 16) has four times the mass of a helium atom (A_r = 4)

By definition, the **carbon-12 isotope** is used as the basis of comparison.

Relative atomic mass, represented by the symbol ……………, is defined as the ……………………………
mass of ………………………………………… of a particular element as compared with ………………………………………………………
…….

Relative molecular mass, represented by the symbol ……………, is defined as the ……………………………
mass of ………………………………………… of a particular substance as compared with ……………………………………………
…….

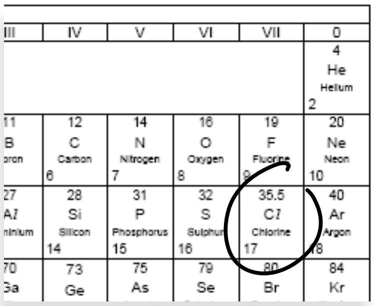
Sometimes, we use the term '**relative formula mass**' instead for ionic compounds (e.g. sodium chloride) since technically, ionic compounds do not form molecules! The chemical symbol for relative formula mass is also M_r.

2. **Why 'Average' Mass?**

In the previous page, we observed that we use the term *average* mass in our definitions. Why is this so significant? Let us recall the concept of *isotopes*.

Since the same element can have different isotopes – e.g. hydrogen can occur as ^{1}H or ^{2}H – it is important to use the average mass for our calculations.

(a) Let us take a look at chlorine, which is circled in the abstract from the Periodic Table below.



III	IV	V	VI	VII	0
					4 He Helium 2
11 B oron	12 C Carbon 6	14 N Nitrogen 7	16 O Oxygen 8	19 F Fluorine 9	20 Ne Neon 10
27 Al inium	28 Si Silicon 14	31 P Phosphorus 15	32 S Sulphur 16	35.5 Cl Chlorine 17	40 Ar Argon 18
70 3a	73 Ge	75 As	79 Se	80 Br	84 Kr

Explain why the mass of chlorine, as stated in the periodic table, is not a whole number.

Chlorine exists as 2 isotopes, chlorine-35 (75%) and chlorine-37 (25%). Hence the mass stated in the periodic table is the average mass of the various isotopes.

Chlorine is not alone when it comes to having a relative mass that is not a whole number. In a more precise version of the Periodic Table, we would see that almost all the elements have decimal places in their relative masses. The first five elements are shown below:

Element	Hydrogen	Helium	Lithium	Beryllium	Boron
Relative Atomic Mass	1.008	4.003	6.941	9.012	10.81

(b) Calculate the relative atomic mass of the following elements to two decimal places, given the abundance of the various isotopes as shown:

Isotope	**Abundance**
bromine-79	50.69 %
bromine-81	49.31 %

A_r = $0.5069 \times 79 + 0.4931 \times 81$
= 79.9862
= 79.99 (2 d.p.)

Isotope	**Abundance**
neon-20	90.48 %
neon-22	9.52 %

A_r = $0.9048 \times 20 + 0.0952 \times 22$
= 20.1904
= 20.19 (2 d.p.)

3. Relative Atomic Mass vs Relative Formula Mass

When we refer to a substance "nitrogen", there are two things that we could be referring to. One is the **atom**, which is straight from the Periodic Table. The other is naturally occurring nitrogen, which is a **molecule** of formula N_2. The same applies for many other elements.

Hence it is important to note that when we ask for the A_r of nitrogen, we are asking for the mass of a nitrogen atom (which is 14). Similarly, when we ask for the M_r of nitrogen, we are asking for the mass of a nitrogen molecule (which is 28).

Substance	Symbol	A_r	Substance	Symbol	M_r
Nitrogen Atom	N	14	Nitrogen Molecule	N_2	28
Chlorine Atom	Cl	35.5	Chlorine Molecule	Cl_2	71
Fluorine Atom	F	19	Fluorine Molecule	F_2	38
Hydrogen Atom	H	1	Hydrogen Molecule	H_2	2
Oxygen Atom	O	16	Oxygen Molecule	O_2	32
Phosphorus Atom	P	31	Phosphorus Molecule	P_4	124
Sulfur Atom	S	32	Sulfur Molecule	S_8	256
Carbon Atom	C	12	Diamond	C	12
Sodium Atom	Na	23	Sodium Metal	Na	23

4. Calculating Relative Formula Mass

Calculate the relative formula mass (M_r) of the substances below.

(a) C_2H_6

M_r = 30

(b) $Ca_3(PO_4)_2$

M_r = 310

(c) CH_3COOH

M_r = 60

(d) Cl_2O

M_r = 87

(e) CO_2

M_r = 44

(f) H_2O

M_r = 18

(g) H_3PO_4

M_r = 98

(h) HCl

M_r = 36.5

(i) $Mg(NO_3)_2$

M_r = 148

(j) NaOH

M_r = 40

(k) $(NH_4)_2SO_4$

M_r = 132

(l) $Zn(OH)_2$

M_r = 99

5. The Mole & Avogadro's Constant

One day, we go to the stationary store to buy A4 paper. We found that they are sold in reams, and each ream of paper contains 500 sheets. Why do they do so? This is because it is too troublesome to sell plain paper sheet by sheet, and thus using a larger quantity such as 500 sheets makes the transaction much simpler.

The same concept can be applied to chemical mass. When we handle chemicals in a laboratory, counting the number of atoms is impossible. So we use a larger denomination, **the mole**, to facilitate calculations.

> **Avogadro's constant (L)** has the value of ……………………………, and represents the number of particles (could be atoms, molecules, ions or even electrons) in one mole of a substances.

The symbol for mole is simply 'mol'.

How many moles of atoms are there in 4.0 mol of nitrogen gas, N_2?	8.0 mol
How many moles of atoms are there in 0.5 mol of methane, CH_4?	2.5 mol
How many moles of ions are there in 3.0 mol of barium chloride, $BaCl_2$?	9.0 mol
How many moles of molecules can be formed from 0.50 mol of oxygen atoms?	0.25 mol
How many moles of molecules can be formed from 2.0 mol of chlorine atoms?	1.0 mol

Why such a number? This is because it was found that 6×10^{23} nucleons (protons or neutrons) have a mass of 1 gram. This allows for simple number-to-mass conversions –

Since: 1 mole of nucleons	*= 1 gram,*	
Thus: 1 mole of oxygen atoms	*= 16 moles of nucleons* *= 16 grams*	*[A_r of oxygen = 16]*
And: 1 mole of oxygen molecules	*= 32 moles of nucleons* *= 32 grams*	*[M_r of oxygen = 32]*

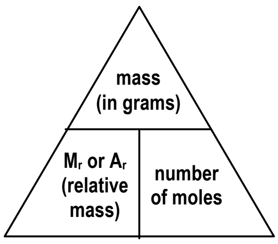
In other words, **the mass of one mole of substance is simply its relative mass**, in grams!

This gives rises to a simple equation:

Mass (in grams) = Relative Mass (M_r or A_r) × No. of Moles

or

$$\textbf{No. of Moles} = \frac{\textbf{Mass (in grams)}}{\textbf{Relative Mass } (M_r \textbf{ or } A_r)}$$



(a) Find the **mass** of

(i) 1.0 mol of carbon dioxide

M_r of CO_2 = 12 + 16 + 16
= 44

Mass of CO_2 = M_r × No. of Moles
= 44 × 1.0
= 44.0 grams (3 s.f.)

(ii) 0.60 mol of ammonia

M_r of NH_3 = 14 + 1 + 1 + 1
= 17

Mass of NH_3 = M_r × No. of Moles
= 17 × 0.60
= 10.2 grams (3 s.f.)

(iii) 2.5 mol of water

M_r of H_2O = 1 + 1 + 16
= 18

Mass of H_2O = M_r × No. of Moles
= 18 × 2.5
= 45.0 grams (3 s.f.)

(iv) 0.020 mol of chlorine gas

M_r of Cl_2 = 35.5 × 2
= 71

Mass of Cl_2 = M_r × No. of Moles
= 71 × 0.020
= 1.42 grams (3 s.f.)

(b) Find the **number of moles** of substance present in

(i) 33.3 grams of calcium chloride

M_r of $CaCl_2$ = 40 + 35.5 + 35.5
= 111

Mol of $CaCl_2$ = Mass ÷ M_r
= 33.3 ÷ 111
= 0.300 mol (3 s.f.)

(ii) 14.2 grams of sodium sulfate

M_r of Na_2SO_4 = 23×2 + 32 + 16×4
= 142

Mol of Na_2SO_4 = Mass ÷ M_r
= 14.2 × 142
= 0.100 mol (3 s.f.)

(iii) 0.45 grams of water vapour

M_r of H_2O = 1 + 1 + 16
= 18

Mol of H_2O = Mass ÷ M_r
= 0.45 ÷ 18
= 0.025 mol (3 s.f.)

(iv) 12.0 grams of oxygen gas

M_r of O_2 = 16 × 2
= 32

Mol of H_2O = Mass ÷ M_r
= 12.0 ÷ 32
= 0.375 mol (3 s.f.)

6. **Applying Mole Ratio**

Observe the balanced chemical equation below:

$$CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

We interpret this as "one methane molecule reacts with two oxygen molecules". Similarly, we can expect one mole of methane molecules to react with two moles of oxygen molecules – since mole is simply a direct multiplication of particles.

This is why balancing equations is so important – this enables us to find the **mole ratio** between various reactants!

(a) In the above equation, how many moles of oxygen gas will react with

(i) 2.0 mol of methane? 4.0 mol

(ii) 0.5 mol of methane? 1.0 mol

(b) In the above equation, how many moles of carbon dioxide will be produced with

(i) 4.0 mol of oxygen? 2.0 mol

(ii) 0.1 mol of oxygen? 0.05 mol

(c) Find the mass of water produced when 40 grams of methane is burnt in excess oxygen. Use the 'table method' below to help you.

	CH_4 (g)	+	2 O_2 (g)	$\longrightarrow$	CO_2 (g)	+	2 H_2O (g)
ratio	1		2		1		2
moles	2.5 mol						5.0 mol
M_r	16						18
mass	40 g						90 g

(d) Find the mass of methane needed to produce 15.4 grams of carbon dioxide.

	CH_4 (g)	+	2 O_2 (g)	$\longrightarrow$	CO_2 (g)	+	2 H_2O (g)
ratio	1		2		1		2
moles	0.35 mol				0.35 mol		
M_r	16				44		
mass	5.60 g				15.4 g		

In summary, chemical calculations require you to <u>convert all values</u> into the <u>number of moles</u> first, before proceeding with further steps.

7. **Review Questions**

(a) A 50 gram sample of calcium carbonate was added to excess dilute hydrochloric acid:

$$CaCO_3\,(s) + 2\,HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2\,(g)$$

Calculate the mass of calcium chloride produced.

Mol of $CaCO_3$ = Mass ÷ M_r
= 50 ÷ (40+12+3×16)
= 0.5 mol

Mol of $CaCl_2$ = Mol of $CaCO_3$
= 0.5 mol

Mass of $CaCl_2$ = No. of Mol × M_r
= 0.5 × (40+2×35.5)
= 55.5 g (3 s.f.)

(b) A 5.4 gram sample of aluminium was allowed to burn in excess oxygen:

$$4\,Al\,(s) + 3\,O_2\,(g) \longrightarrow 2\,Al_2O_3\,(s)$$

Calculate the mass of oxygen used.

Mol of Al = Mass ÷ A_r
= 5.4 ÷ 27
= 0.2 mol

Mol of O_2 = $\frac{3}{4}$ × Mol of Al
= $\frac{3}{4}$ × 0.2 mol
= 0.15 mol

Mass of O_2 = No. of Mol × M_r
= 0.15 × (2×16)
= 4.80 g (3 s.f.)

(c) 0.030 kg of magnesium metal was added to excess dilute nitric acid.

$$Mg\,(s) + 2\,HNO_3\,(aq) \longrightarrow Mg(NO_3)_2\,(aq) + H_2\,(g)$$

Calculate the mass of acid used up in the reaction.

Mol of Mg = Mass ÷ A_r
= 30 ÷ 24
= 1.25 mol

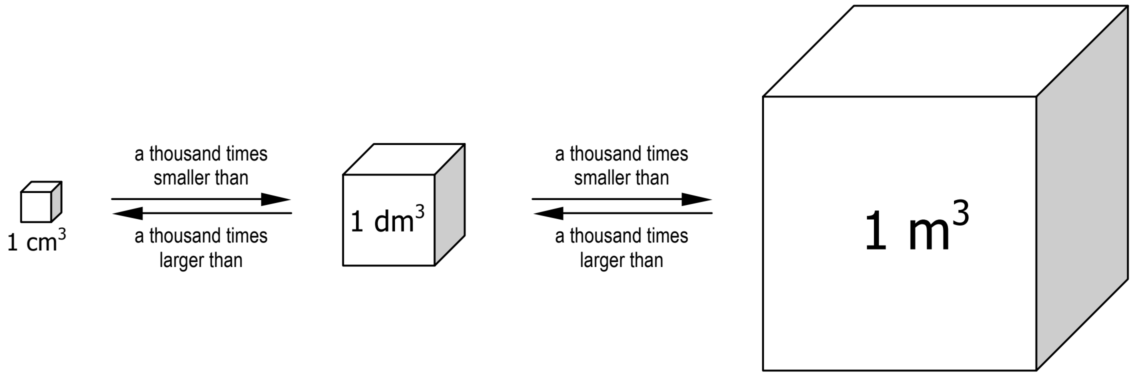
Mol of HNO_3 = 2 × Mol of Mg
= 2 × 1.25 mol
= 2.50 mol

Mass of HNO_3 = No. of Mol × M_r
= 2.50 × (1+14+3×16)
= 158 g (3 s.f.)

8. **Units of Volume**

In chemistry, the key unit we use for calculating volume is the **cubic decimetre** (dm^3), which is **equivalent to a volume of one litre**.

Many of our laboratory measurements, however, come in cubic centimetres (cm^3) or cubic metres (m^3); hence it is important to convert between the various units of volume.



(a) Convert the following volumes into dm^3.

(i) 25.5 cm^3	(iii) 0.350 m^3	(v) 5.00 litres
0.0255 dm^3	**350 dm^3**	**5.00 dm^3**
(ii) 1500 cm^3	(iv) 5.50 m^3	(vi) 300 millilitres
1.50 dm^3	**5500 dm^3**	**0.300 dm^3**

(b) Convert the following volumes into cm^3.

(i) 24 dm^3	(iii) 2.00 dm^3	(v) 0.250 litres
24 000 cm^3	**2000 cm^3**	**250 cm^3**
(ii) 0.0048 dm^3	(iv) 0.124 m^3	(vi) 500 millilitres
4.80 cm^3	**124 000 cm^3**	**500 cm^3**

Note that chemical calculations require you use grams (g) in calculations involving mass and cubic decimeters (dm^3) in calculations involving volume!

9. **<u>Concentration of Aqueous Solutions</u>**

Earlier, we have looked at calculating the number of moles of a particular substance by dividing its mass by its M_r. However, when it comes to aqueous solutions, it is difficult to find the mass of substance dissolved in the water. Instead, we use the concentration and volume of the solution.

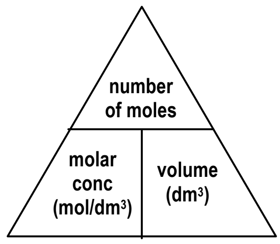
It is important to note that the term 'concentration' can refer to two things – **molar concentration** or **mass concentration**.

MOLAR CONCENTRATION – measured in ………………………

$$\textbf{Molar Concentration} = \frac{\textbf{No. of Moles}}{\textbf{Volume}}$$

or

$$\textbf{No. of Moles} = \textbf{Molar Concentration} \times \textbf{Volume}$$



(a) Find the number of moles present in

(i) 3.00 dm^3 of 0.500 mol/dm^3 HCl

Mol of HCl = Molar Conc × Vol
= 0.500 × 3.00
= 1.50 mol (3 s.f.)

(ii) 0.620 dm^3 of 2.40 mol/dm^3 NaOH

Mol of NaOH = Molar Conc × Vol
= 2.40 × 0.620
= 1.49 mol (3 s.f.)

(iii) 250 cm^3 of 2.00 mol/dm^3 H_2SO_4

Mol of H_2SO_4 = Molar Conc × Vol
= 2.00 × 0.250
= 0.500 mol (3 s.f.)

(iv) 20.0 cm^3 of 0.100 mol/dm^3 KOH

Mol of KOH = Molar Conc × Vol
= 0.100 × 0.0200
= 0.00200 mol (3 s.f.)

(b) Calculate the mass of salt present in a 25.0 cm^3 sample of 0.400 mol/dm^3 $CuSO_4$.

Mol of $CuSO_4$ = Molar Conc × Vol
= 0.400 × 0.025
= 0.1 mol

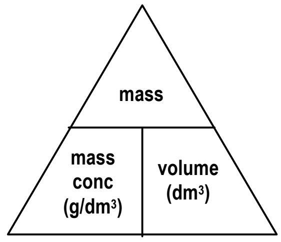
Mass of $CuSO_4$ = Mol × Mr
= 0.1 × (64+32+4×16)
= 0.1 × 160
= 16.0 grams (3 s.f.)

MASS CONCENTRATION – measured in ....

Mass Concentration = $\dfrac{\textbf{Mass}}{\textbf{Volume}}$

or

Mass = Mass Concentration × Volume



(c) Find the number of moles present in

(i) 2.00 dm^3 of 7.3 g/dm^3 HCl

Mass of HCl = Mass Conc × Vol
= 7.3 × 2.00
= 14.6 grams

Mol of HCl = Mass ÷ M_r
= 14.6 ÷ (1 + 35.5)
= 0.400 mol
(3 s.f.)

(ii) 0.150 dm^3 of 2.0 g/dm^3 NaOH

Mass of NaOH = Mass Conc × Vol
= 2.0 × 0.150
= 0.3 grams

Mol of NaOH = Mass ÷ M_r
= 0.3 ÷ (23+16+1)
= 0.00750 mol
(3 s.f.)

(iii) 480 cm^3 of 4.9 g/dm^3 H_2SO_4

Mass of H_2SO_4 = Mass Conc × Vol
= 4.9 × 0.480
= 2.352 grams

Mol of H_2SO_4 = Mass ÷ M_r
= 2.352 ÷ 98
= 0.0240 mol
(3 s.f.)

(iv) 8.00 cm^3 of 7.0 g/dm^3 KOH

Mass of KOH = Mass Conc × Vol
= 7.0 × 0.00800
= 0.056 grams

Mol of KOH = Mass ÷ M_r
= 0.056 ÷ 56
= 0.00100 mol
(3 s.f.)

(d) 3.8 grams of magnesium chloride is dissolved in 400 cm^3 of water. Calculate the molar concentration of the resulting solution.

Mol of $MgCl_2$ = Mass ÷ M_r
= 3.8 ÷ (24+2×35.5)
= 3.8 ÷ 95
= 0.04 mol

Molar Conc of $MgCl_2$ = Mol ÷ Vol
= 0.04 ÷ 0.400
= 0.100 mol/dm^3 (3 s.f.)

10. **Review Questions**

(a) A 25.0 cm^3 sample of 0.100 mol/dm^3 sulfuric acid was neutralized by adding a certain volume of dilute sodium hydroxide. Below shows the chemical equation for this reaction.

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

(i) Calculate the number of moles of sulfuric acid used.

Mol of H_2SO_4 = Molar Concentration × Volume
= 0.100 × 0.025
= 0.00250 mol

(ii) Calculate the volume of 0.125 mol/dm^3 sodium hydroxide needed to neutralize this volume of sulfuric acid.

Mol of NaOH = 2 × Mol of H_2SO_4
= 0.0050 mol

Vol of NaOH = Mol ÷ Molar Conc
= 0.0050 ÷ 0.125
= 0.040 dm^3 (or 40 cm^3)

(b) In a laboratory experiment, 19 g of anhydrous magnesium chloride was dissolved in 1 dm^3 of water. 20 cm^3 of the resulting solution was extracted and allowed to react with excess silver nitrate in the following precipitation reaction.

$$MgCl_2\ (aq) + 2\ AgNO_3\ (aq) \longrightarrow 2\ AgCl\ (s) + Mg(NO_3)_2\ (aq)$$

(i) Find the number of moles of magnesium chloride in 20 cm^3 of the solution.

Mass of $MgCl_2$ = Mass Conc × Volume
= 19 × 0.020
= 0.38 grams

Mol of $MgCl_2$ = Mass ÷ M_r
= 0.38 ÷ (24+35.5×2)
= 0.00400 mol

(ii) Hence calculate the mass of the precipitate produced.

Mol of AgCl = 2 × Mol of $MgCl_2$
= 2 × 0.00400
= 0.00800 mol

Mass of AgCl = Mol of AgCl × M_r
= 0.00800 × (108+35.5)
= 1.15 grams

(iii) Calculate the molar concentration of the original magnesium chloride solution.

Mol of $MgCl_2$ in 19 grams = Mass ÷ M_r
= 19 ÷ (24+35.5×2)
= 0.200 mol

Molar Conc of $MgCl_2$ = Mol of $MgCl_2$ ÷ Volume of Water Dissolved
= 0.200 ÷ 1
= 0.200 mol/dm^3

11. **Volume of Gases**

Like aqueous solutions, it is hard to obtain the mass of a gas in a laboratory. Hence, often we find the number of moles of a gas by using Avogadro's Law.

> **Avogadro's Law** states that ………………………… of any gas occupies a volume of …………………………,
>
> under room conditions of ………………………… and ………………………….

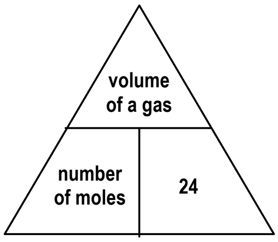
It is important to note that **this applies only for gases**, and not solids, liquids nor solutions.

VOLUME OF GASES

$$\textbf{Volume of Gas (in dm}^3\textbf{)} = \textbf{No. of Moles} \times 24$$

or

$$\textbf{No. of Moles} = \frac{\textbf{Volume of Gas (in dm}^3\textbf{)}}{\textbf{24}}$$



(a) State the volume (in dm^3) of the following gases, assuming room conditions:

(i) 1.0 mol of ethane, C_2H_6 …… 24.0 dm^3

(ii) 1.0 mol of sulfur dioxide, SO_2 …… 24.0 dm^3

(iii) 2.0 mol of carbon dioxide, CO_2 …… 48.0 dm^3

(iv) 0.5 mol of ammonia, NH_3 …… 12.0 dm^3

(v) 8.0 grams of oxygen gas, O_2 …… 6.00 dm^3

(vi) 4.0 grams of hydrogen, H_2 …… 48.0 dm^3

(b) State the number of moles of the following gases, assuming room conditions:

(i) 2.4 dm^3 of nitrogen dioxide, NO_2 …… 0.100 mol

(ii) 2.4 dm^3 of chlorine, Cl_2 …… 0.100 mol

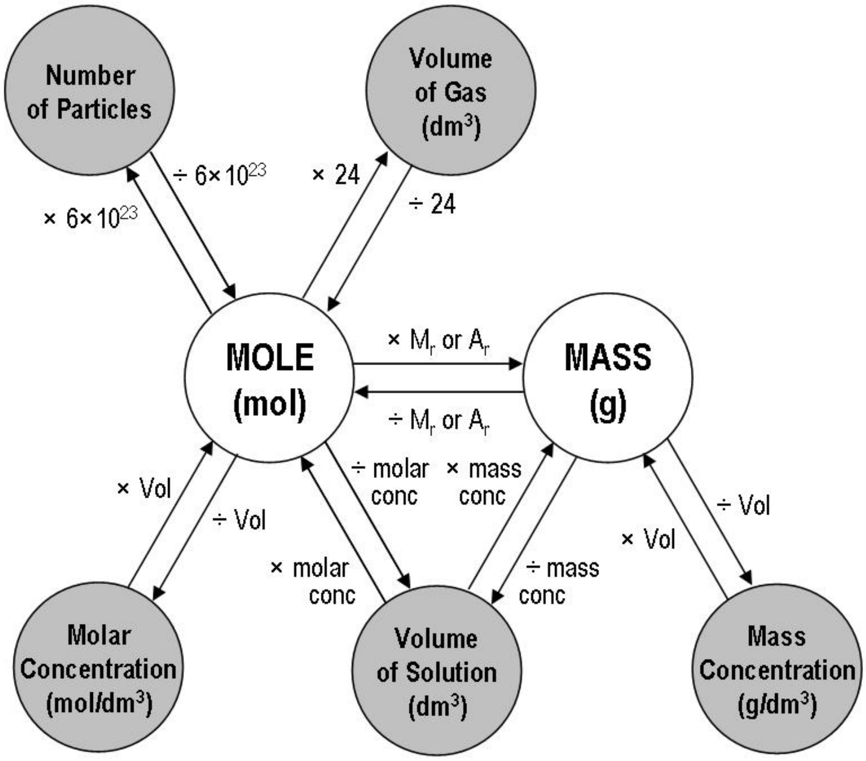
(iii) 360 cm^3 of nitrogen, N_2 …… 0.0150 mol

(iv) 0.48 m^3 of methane, CH_4 …… 20.0 mol

(v) 48 dm^3 of ozone, O_3 …… 2.00 mol

(vi) 48 g of ozone, O_3 …… 1.00 mol

12. **<u>Review of Mole Calculation Formulas</u>**



(a) 300 cm^3 of hydrogen gas, under room conditions, was burnt in excess oxygen, producing water as shown:

$$2\ H_2\,(g) + O_2\,(g) \longrightarrow 2\ H_2O\,(l)$$

(i) Find the number of moles of hydrogen gas combusted.

Mol of H_2 = Vol of H_2 ÷ 24
= 0.300 ÷ 24
= 0.0125 mol

(ii) Find the mass of water produced.

Mol of H_2O = Mol of H_2
= 0.0125 mol

Mass of H_2O = Mol of H_2O × M_r
= 0.0125 × (2+16)
= 0.225 grams

(iii) Calculate the volume of oxygen used up in this reaction.

Mol of O_2 = ½ × Mol of H_2
= ½ × 0.0125
= 0.00625 mol

Vol of O_2 = Mol of O_2 × 24
= 0.00625 × 24
= 0.150 dm^3 (or 150 cm^3)

(b) Jonathan prepares carbon dioxide in the lab by reacting 20.0 grams of solid calcium carbonate with excess of 0.100 mol dm^{-3} hydrochloric acid.

$$CaCO_3\,(s) + 2\,HCl\,(aq) \longrightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2\,(g)$$

(i) Find the number of moles in 20.0 grams of calcium carbonate.

Mol of $CaCO_3$ = Mass ÷ M_r
= 20 ÷ (40 + 12 + 3×16)
= 0.200 mol

(ii) Calculate the volume of carbon dioxide gas produced.

Mol of CO_2 = Mol of $CaCO_3$
= 0.200 mol

Vol of CO_2 = Mol of CO_2 × 24
= 0.200 × 24
= 4.80 dm^3

(iii) Calculate the mass of water produced.

Mol of H_2O = Mol of $CaCO_3$
= 0.200 mol

Mass of H_2O = Mol of H_2O × M_r
= 0.200 × (2+16)
= 3.60 grams

(iv) Calculate the minimum volume of 0.100 mol dm^{-3} hydrochloric acid needed.

Mol of HCl = 2 × Mol of $CaCO_3$
= 2 × 0.200 mol
= 0.400 mol

Vol of HCl = Mol of HCl ÷ Molar Conc
= 0.400 ÷ 0.100
= 4.00 dm^3

(v) Find the mass of calcium carbonate needed to produce 60 cm^3 of carbon dioxide instead.

Mol of CO_2 = Vol of CO_2 ÷ 24
= 0.060 ÷ 24
= 0.0025 mol

Mol of $CaCO_3$ = Mol of CO_2
= 0.0025 mol

M_r of $CaCO_3$ = 40 + 12 + 3×16
= 100

Mass of $CaCO_3$ = Mol of $CaCO_3$ × M_r
= 0.0025 × 100
= 0.250 grams

(c) An unknown metal, **M**, was allowed to react with dilute nitric acid as shown:

$$\mathbf{M\ (s) + 2\ HNO_3\ (aq) \longrightarrow M(NO_3)_2\ (aq) + H_2\ (g)}$$

It was found that a 8.2 gram sample of metal **M** required 40.0 cm^3 of 189 g/dm^3 nitric acid for complete reaction.

(i) Find the number of moles of nitric acid used.

Mass of HNO_3 = Mass Conc × Volume
= 189 × 0.040
= 7.56 grams

Mol of HNO_3 = Mass of HNO_3 ÷ M_r
= 7.56 ÷ (1+14+16×3)
= 0.120 mol

(ii) State the number of moles of metal **M** used.

Mol of M = $\frac{1}{2}$ × Mol of HNO_3
= $\frac{1}{2}$ × 0.12
= 0.0600 mol

(iii) Hence calculate the relative atomic mass (A_r) of metal **M**.

A_r of M = Mass of M ÷ Mol of M
= 8.2 ÷ 0.0600
= 136.67
= 137 (3 s.f.)

(iv) Calculate the volume of gas produced.

Mol of H_2 = $\frac{1}{2}$ × Mol of HNO_3
= $\frac{1}{2}$ × 0.12
= 0.0600 mol

Vol of H_2 = Mol of H_2 × 24
= 0.0600 × 24
= 1.44 dm^3

(v) Assuming that the total volume of the resulting solution remains at 40.0 cm^3, calculate the mass concentration of the salt formed.

Mol of $M(NO_3)_2$ = $\frac{1}{2}$ × Mol of HNO_3
= $\frac{1}{2}$ × 0.12
= 0.0600 mol

Mass of $M(NO_3)_2$ = Mol of HNO_3 × M_r
= 0.0600 × 260.67
= 15.6402 grams

M_r of $M(NO_3)_2$ = 136.67 + 2(14+16×3)
= 260.67

Mass Conc = Mass ÷ Volume
= 15.6402 ÷ 0.0400
= 391.005 g/dm^3
= 391 g/dm^3 (3 s.f.)

13. **Limiting & Excess Reactants**

Up till now, we have always used a single reactant to determine the amount of product formed, assuming that the remaining reactants are in excess? Let us use an analogy to understand this.

3 Scoops of Ice-Cream + 2 Bananas ⟶ 1 Banana Split

How many banana splits can I make if I have 12 scoops of ice-cream?

If you said "4" to the above question, you are making an assumption that you have enough bananas. What if you only had 6 bananas to go with these 12 scoops of ice-cream?

Naturally you can make only 3 banana splits. We are limited by the number of bananas that we have (we call this the **limiting reagent**), and we have extra ice-cream that is unused at the end of the process (we call this the **excess reagent**).

Similarly, when dealing with chemical calculations, it is important to determine which reactant is the limiting reagent – **we must choose the limiting reagent when calculating mole ratio.**

(a) Complete the table below for the reaction between nitrogen and hydrogen.

$$N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g)$$

Mol of N_2	Mole of H_2	Limiting Reagent	Mole of NH_3 Formed	Excess
1 mol	4 mol	N_2	2 mol	1 mol H_2
2 mol	3 mol	H_2	2 mol	1 mol N_2
6 mol	6 mol	H_2	4 mol	4 mol N_2
3 mol	11 mol	N_2	6 mol	2 mol H_2
2.5 mol	4.5 mol	H_2	3.0 mol	1.0 mol N_2
0.2 mol	1.0 mol	N_2	0.4 mol	0.4 mol H_2

(b) 36.0 dm^3 of methane is allowed to combust with 48.0 dm^3 of oxygen:

$$CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(l)$$

By showing your working carefully, identify the limiting reagent.

Mol of CH_4 = Vol ÷ 24
= 36.0 ÷ 24
= 1.5 mol

Mol of O_2 = Vol ÷ 24
= 48.0 ÷ 24
= 2.0 mol

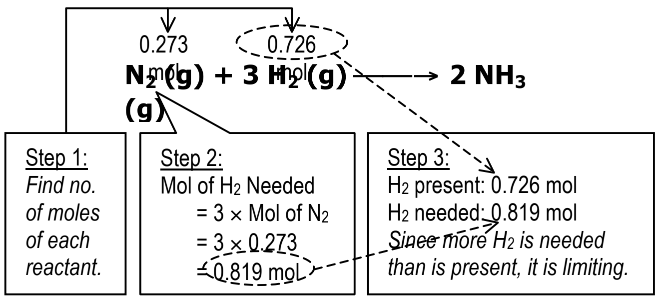
1.5 mol of CH_4 would require 3.0 mol of O_2 for complete reaction.
However there is insufficient O_2 (only 2.0 mol present).
∴ O_2 is limiting.

14. **Limiting & Excess Reactants – A Systematic Approach**

While sometimes we may determine the limiting reagent mentally, often it becomes hard to do so especially if there are many decimal places involved. We hence adopt a systematic approach.

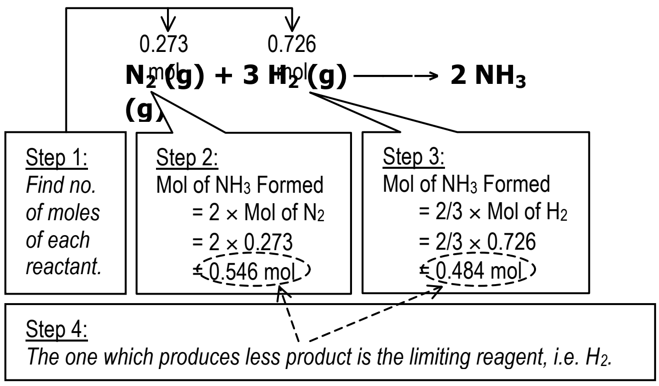
APPROACH #1 (Ratio of Reactants)

① Compute the number of moles of each reactant.
② Calculate the number of moles of reactant B necessary to completely react with reactant A.
③ Check if the value computed in Step 2 is sufficient – if it is, reactant B is in excess. If it is insufficient, reactant B is limiting.



APPROACH #2 (Lesser of Products)

① Compute the number of moles of each reactant.
② Calculate the number of moles of product formed, using only the amount of reactant A.
③ Calculate the number of moles of product formed, using only the amount of reactant B.
④ The reactant which produces a smaller amount of product is the limiting reagent.



(a) 20.0 grams of hydrogen was allowed to react with 20.0 grams of oxygen, producing water. Calculate the maximum mass of water that can be formed.

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O\,(l)$$

Mol of H_2 = Mass ÷ M_r = 20 ÷ 2
= 10.0 mol

Mol of O_2 = Mass ÷ M_r = 20 ÷ 32
= 0.625 mol

10.0 mol of H_2 requires 5.0 mol of O_2 for complete reaction.

However, there is only 0.625 mol of O_2 (i.e. insufficient). Hence O_2 is limiting.

We hence use the number of moles of the limiting reagent, O_2, to apply the mole ratio.

Mol of H_2O formed = 2 × Mol of O_2
= 2 × 0.625
= 1.25 mol

Mass of H_2O formed = Mol of H_2O × M_r
= 1.25 × 18
= 22.5 grams

(b) Find the maximum number of moles of salt that can be formed when 20.0 cm^3 of 0.200 M dilute sulfuric acid is allowed to react with 30.0 cm^3 of 0.300 M sodium hydroxide.

$H_2SO_4 + 2\ NaOH \longrightarrow Na_2SO_4 + 2\ H_2O$

Mol of H_2SO_4 = Molar Conc × Vol
= 0.200 × 0.0200
= 0.00400 mol

Mol of NaOH = Molar Conc × Vol
= 0.300 × 0.0300
= 0.00900 mol

0.00400 mol of H_2SO_4 requires 0.00800 mol of NaOH for complete reaction.
Since there is 0.00900 mol of NaOH present, NaOH must be in excess.

∴ H_2SO_4 is limiting. We apply the mole ratio using the mole of H_2SO_4 present.

Mol of Na_2SO_4 = Mol of H_2SO_4
= 0.00400 mol (3 s.f.)

(c) In the preparation of insoluble lead(II) iodide, 24.0 cm^3 of 2.00 mol/dm^3 lead(II) nitrate is mixed with 150 cm^3 of 0.150 mol/dm^3 potassium iodide. Calculate the mass of precipitate that is produced.

$Pb(NO_3)_2 + 2\ KI \longrightarrow PbI_2 + 2\ KNO_3$

Mol of $Pb(NO_3)_2$ = Molar Conc × Vol
= 0.024 × 1.00
= 0.024 mol

Mol of KI = Molar Conc × Vol
= 0.150 × 0.100
= 0.0225 mol

Since 0.024 mol of $Pb(NO_3)_2$ requires 0.048 mol of KI for complete reaction.
Since there is only 0.0225 mol of KI present, there is insufficient KI.

∴ KI is limiting. We apply mole ratio using the mole of KI present.

Mol of PbI_2 = $\frac{1}{2}$ × Mol of KI
= $\frac{1}{2}$ × 0.0225 mol
= 0.01125 mol

Mass of PbI_2 = Mol of PbI2 × M_r
= 0.01125 × (207 + 2 × 127)
= 0.01125 × 461
= 5.17 grams (3 s.f.)

Self-Designed Summary

Supplementary Questions

1. Observe the following reaction:

$$2\ C_2H_6\ (g) + 7\ O_2\ (g) \longrightarrow 4\ CO_2\ (g) + 6\ H_2O\ (l)$$

If 378 grams of water was produced in this reaction, find
(a) mass of carbon dioxide produced,
(b) mass of ethane (C_2H_6) required,
(c) mass of oxygen gas required.

2. Observe the following reaction:

$$2\ AgNO_3\ (aq) + MgCl_2\ (aq) \longrightarrow 2\ AgCl\ (s) + Mg(NO_3)_2\ (aq)$$

If 120 grams of $MgCl_2$ reacted, find the
(a) mass of magnesium nitrate produced,
(b) mass of silver chloride produced,
(c) mass of silver nitrate required.

3. Observe the following reaction:

$$Na_2CO_3\ (aq) + 2\ HCl\ (aq) \longrightarrow 2\ NaCl\ (aq) + CO_2\ (g) + H_2O\ (l)$$

In the above reaction, 40cm^3 of 0.200 mol/dm^3 HCl was used to react with sodium carbonate.
(a) Find the number of moles of HCl used.
(b) Find the number of moles of Na_2CO_3 needed.
(c) Find the volume of Na_2CO_3 required, given that its concentration is 0.150 mol/dm^3.
(d) Find the mass of water formed.

4. Observe the following reaction:

$$(NH_4)_2SO_4\ (aq) + 2\ NaOH\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2\ NH_3\ (g) + 2\ H_2O\ (l)$$

If 25 cm^3 of 0.200 mol/dm^3 sodium hydroxide was needed for complete reaction, find
(a) the number of moles of NaOH needed
(b) the volume of 0.100 mol/dm^3 $(NH_4)_2SO_4$ needed
(c) the volume of NH_3 produced
(d) the mass of salt formed

5. Observe the following reaction:

$$MgCO_3\ (s) + 2\ HCl\ (aq) \longrightarrow MgCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

In the above reaction, 21 g of carbonate was allowed to react with dilute hydrochloric acid. Find
(a) volume of 0.200 mol/dm^3 HCl needed,
(b) volume of carbon dioxide gas produced,
(c) mass of salt produced,
(d) concentration of the resulting $MgCl_2$ solution in mol/dm^3, using (a) as the total volume.

6. Observe the following reaction:

$$MgCO_3\ (s) + 2\ HNO_3\ (aq) \longrightarrow Mg(NO_3)_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

In the above reaction, 360 cm^3 of carbon dioxide was produced.
(a) Find the mass of water produced.
(b) Find the volume of 0.400 mol/dm^3 nitric acid needed.
(c) Calculate the mass of magnesium carbonate needed.
(d) Assuming the total volume of the resulting solution to be the same as your answer in (b), calculate the concentration of the salt produced in g/dm^3.

7. A beaker contains 300 grams of water, and reacts as follows:

$$2\ K\ (s) + 2\ H_2O\ (l) \longrightarrow 2\ KOH\ (aq) + H_2\ (g).$$

Find the maximum volume of hydrogen gas produced when 0.60 grams of potassium is added to this beaker of water.

8. Ethene combusts in oxygen as shown:

$$C_2H_4\ (g) + 3\ O_2\ (g) \longrightarrow 2\ CO_2\ (g) + 2\ H_2O\ (g)$$

Assuming that only complete combustion can occur, calculate the maximum volume of carbon dioxide that can be formed, measured at r.t.p., when 7.0 grams of ethene is allowed to combust with 12.0 dm^3 of oxygen.

9. Find the maximum mass of precipitate formed when 20.0 cm^3 of 0.500 M silver nitrate is mixed with 30.0 cm^3 of 0.200 M magnesium chloride.

10. Find the loss in mass of the reaction vessel when 5.0 grams of magnesium carbonate is allowed to react with 100 cm^3 of 1.00 mol/dm^3 nitric acid.

Supplementary Questions (Answers)

Question 1
(a) 616 g
(b) 210 g
(c) 784 g

Question 2
(a) 187 g
(b) 363 g
(c) 429 g

Question 3
(a) 0.00800 mol
(b) 0.00400 mol
(c) 26.7 cm^3 (or 0.0267 dm^3)
(d) 0.0720 g

Question 4
(a) 0.005 mol
(b) 25.0 cm^3 (or 0.0250 dm^3)
(c) 0.12 dm^3 (or 120 cm^3)
(d) 0.355 g

Question 5
(a) 2.50 dm^3
(b) 6.00 dm^3
(c) 23.8 g
(d) 0.100 mol/dm^3

Question 6
(a) 0.270 g
(b) 75.0 cm^3 (or 0.0750 dm^3)
(c) 1.26 g
(d) 29.6 g/dm^3

Question 7
0.185 dm^3

Question 8
8.00 dm^3

Question 9
1.44 g

Question 10
2.20 g

Lecture Slides

Chemical Calculations

Anglo-Chinese School (Barker Road)

CHEMISTRY
Prepared by Alex Lee

Chapter Overview

Textbook Reference: Chapters 9 & 10

1. Relative Mass
2. Mole Concept & Mole Ratio
3. Calculating Mole from Mass
4. Calculating Mole from Aqueous Concentration
5. Calculating Mole from Volume of Gas
6. Limiting & Excess Reagents

Relative Mass

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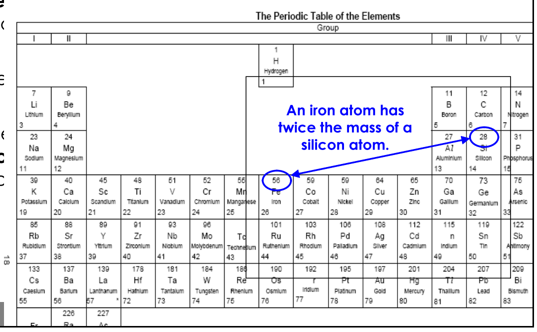

Relative Mass

Relative Atomic Mass (A_r)

"The <u>average</u> mass of <u>one atom</u> of an element, as compared with <u>1/12 the mass of a carbon-12 isotope</u>."

Why 'Average'?

Recall – each element is made up of several isotopes.

For example, chlorine has a relative atomic mass of 35.5. This is because chlorine is made up of two isotopes – chlorine-35 (75% abundance) and chlorine-37 (25% abundance).

Hence, A_r of chlorine = (75/100 × 35) + (25/100 × 37) = 35.5

To find A_r, we look for the 'top number' in the periodic table.
[e.g. A_r of Nitrogen = 14]

Relative Mass

Relative Molecular Mass (M_r)

"The <u>average</u> mass of <u>one molecule</u> of a substance, as compared with <u>1/12 the mass of a carbon-12 isotope</u>."

Relative Formula Mass (M_r)

Sometimes we use the term 'relative **formula** mass' instead of relative **molecular** mass. This is more accurate when describing ionic compounds – as ionic compounds do not form molecules!

However, both are represented by the same symbol, M_r, and have identical calculations.

To find M_r, we add A_r of the atoms within the compound.
[e.g. M_r of Nitrogen (N_2) = 2 × (A_r of Nitrogen) = 28]

Relative Mass

- In your version of the Periodic Table, chlorine has a relative mass of 35.5, while all other elements seem to have a whole number.
- This is actually a great simplification.
- Many other elements also have decimal places in their relative masses:
  - Hydrogen: 1.008
  - Helium: 4.003
  - Lithium: 6.941
  - Beryllium: 9.012
- This indicates that all elements do exist in isotopes!

Finding Relative Mass

- The **relative atomic mass** of any element can be found directly from the periodic table.
  e.g. relative mass of C = 12 *[from periodic table]*

- The **relative molecular mass**, on the other hand, needs to be calculated by adding up the various relative atomic masses of its constituent atoms.
  e.g. relative mass of CO_2 = mass of C + 2 × mass of O
  = 12 + (2 × 16)
  = 44

Finding Relative Mass

We have to be careful with non-metallic elements, such as nitrogen, oxygen, hydrogen and chlorine.

For example, the term "nitrogen" could refer to:

Nitrogen **Atoms**	Nitrogen **Molecules**
Chemical Formula: **N**	Chemical Formula: **N_2**
Relative Atomic Mass: **14**	Relative Molecular Mass: **28**

The Mole

- When we go to the supermarket, we see these prices:

Fresh Eggs	$1.50 / dozen
Minced Beef	$2.10 / 100 grams
White Sugar	$4.20 / packet

Why can't we purchase a single egg, a single gram of beef, or a single grain of sugar?

- The same applies to chemicals. It is impossible to measure out a single atom or molecule, and hence we use a larger quantity – a **mole** (or 'mol' for short) of atoms or molecules instead.

The Mole

- The number of atoms (or any other particles) in a mole is defined by **Avogadro's constant**:

$$\mathbf{6 \times 10^{23}}$$

(600,000,000,000,000,000,000,000)

"one mole of atoms" = 6×10^{23} atoms
"two moles of electrons" = 12×10^{23} electrons [1.2×10^{24}]
"half mole of molecules" = 3×10^{23} molecules

The Mole

- How many atoms are there in a molecule of CO_2?
  3 atoms
- How many atoms are there in a dozen CO_2?
  3 dozen atoms
- How many atoms are there in a mole of CO_2?
  3 moles of atoms
- How many atoms are there in two moles of CO_2?
  6 moles of atoms

The Mole

How did Avogardo's constant come about?

- It was found that a particle with a relative mass of "1" (e.g. a single hydrogen atom) had a mass of $\frac{1}{6\times10^{23}}$ grams.

- Therefore, 6×10^{23} particles ("one mole") with a relative mass of "1" should have a mass of one gram!

- This allows us to convert between relative mass (which we cannot measure) and real mass (which we can measure) much more simply.

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The Mole

- As a result, the mass of one mole of any substance is equal to its own relative mass, in grams!

What is the mass of **1 mole** of carbon?
A_r of C = 12 **Mass of 1 mol of C = 12 grams**

What is the mass of **1 mole** of sodium?
A_r of Na = 23 **Mass of 1 mol of Na = 23 grams**

What is the mass of **1 mole** of oxygen gas?
M_r of O_2 = 32 **Mass of 1 mol of O_2 = 32 grams**

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The Mole

- Similarly, if there are more or less than one more of the substance, we can predict its mass proportionately.

What is the mass of **2 moles** of carbon?
A_r of C = 12 **Mass of 2 mol of C = 24 grams**

What is the mass of **10 moles** of sodium?
A_r of Na = 23 **Mass of 10 mol of Na = 230 grams**

What is the mass of **0.25 moles** of oxygen gas?
M_r of O_2 = 32 **Mass of 0.25 mol of O_2 = 8 grams**

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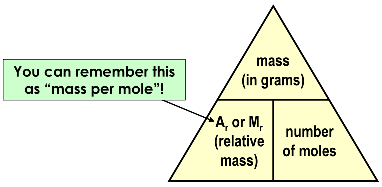
The Mole

- This can be summarized as a simple equation:

Mass = No. of Moles × Relative Mass



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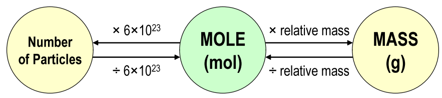
Concept Map

A summary of the two relationships learnt so far:



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Applying Mole Ratio

Observe the balanced chemical equation below:

$$CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

- From the equation, we see that one mole of methane will react with two moles of oxygen gas. This produces one mole of carbon dioxide and one mole of water.

- We apply this concept of mole ratio ("how many of A will react with how many of B") to solve chemical calculations.

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Applying Mole Ratio

Observe the balanced chemical equation below:

$$CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

How many moles of oxygen gas will react
(i) with 2.0 mol of methane? **4.0 mol**
(ii) with 0.5 mol of methane? **1.0 mol**

How many moles of carbon dioxide will be produced
(i) with 4.0 mol of oxygen? **2.0 mol**
(ii) with 0.1 mol of oxygen? **0.05 mol**

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Applying Mole Ratio

Observe the balanced chemical equation below:

$$CH_4\ (g) + 2\ O_2\ (g) \longrightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

• Find the mass of water produced when **40 grams** of methane is burnt in excess oxygen.

	CH_4 (g)	+ 2 O_2 (g) →	CO_2 (g) +	2 H_2O (g)
mole ratio	1	2	1	2
no of moles	2.5	5	2.5	5
M_r	16	32	44	18
mass (grams)	40 g	160 g	110 g	90 g

Applying Mole Ratio

Observe the balanced chemical equation below:

$$CH_4\ (g) + 2\ O_2\ (g) \longrightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

• Find the mass of methane needed to produce **15.4 grams** of carbon dioxide.

	CH_4 (g)	+ 2 O_2 (g) →	CO_2 (g) +	2 H_2O (g)
mole ratio	1	2	1	2
no of moles	0.35	0.70	0.35	0.70
M_r	16	32	44	18
mass (grams)	5.60 g	22.4 g	15.4 g	12.6 g

Applying Mole Ratio

Most important thing is to calculate the **number of moles** from the given data.

Then apply **mole ratio**, from a balanced chemical equation, to make calculations.

(You **may not** calculate the mass of product directly from the mass of reactant, or vice versa.)

Units of Volume

• In chemistry, we use the **cubic decimetre (dm^3)** as our unit of volume. This is the same as one litre.

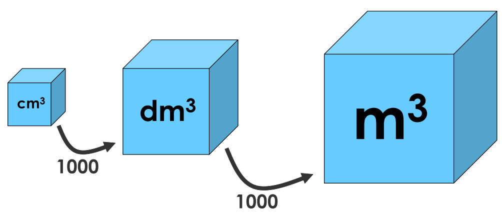
1 dm	= 0.1 m	1 dm	= 10 cm
$(1\ dm)^3$	= $(0.1\ m)^3$	$(1\ dm)^3$	= $(10\ cm)^3$
1 dm^3	= 0.001 m^3	1 dm^3	= 1000 cm^3
			(1 litre = 1000 millilitres)

• To convert between the various units:

$$m^3 \xrightarrow{\times 1000} dm^3 \qquad cm^3 \xrightarrow{\div 1000} dm^3$$

Units of Volume

• Just remember which is the largest and which is the smallest – to convert between each other, the difference is simply a thousand!



Which Units to Use?

• In chemistry, we always use the unit of **grams (g)** for calculations involving mass, and the unit of **cubic decimetres (dm^3)** for calculations involving volume.

• This is not to be confused with physics!

	Chemistry	Physics
mass	grams	kilograms
volume	dm^3	m^3

Aqueous Concentration

- Used for **aqueous solutions** only.
- Refers to the 'amount of chemical **per unit volume** of water'. This can mean two things:

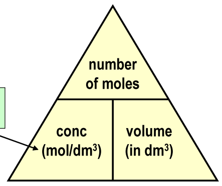
MOLAR CONCENTRATION – mol / dm^3
MASS CONCENTRATION – g / dm^3

Aqueous Concentration

Molar Concentration (mol/dm^3)

Molar Concentration = $\frac{\text{No. of Moles}}{\text{Volume}}$

You can identify molar concentration from its units.



Aqueous Concentration

Molar Concentration (mol/dm^3)

"mol/dm^3" can be shortened to "M"

"Find the number of moles in 0.500 dm^3 of 3.00 mol/dm^3 hydrochloric acid."

mol/dm^3 = molar concentration!

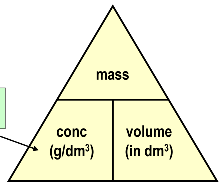
Mol of HCl = Molar Conc × Volume
= 3.00 × 0.500
= 1.50 mol

Aqueous Concentration

Mass Concentration (g/dm^3)

Mass Concentration = $\frac{\text{Mass (grams)}}{\text{Volume}}$

You can identify mass concentration from its units.



Aqueous Concentration

Mass Concentration (g/dm^3)

"Find the number of moles in 1.25 dm^3 of 0.250 g/dm^3 sodium hydroxide."

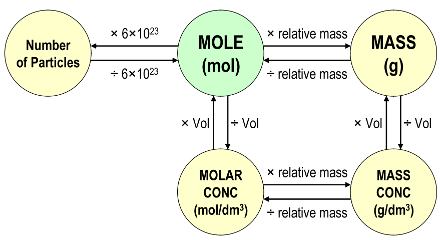
g/dm^3 = mass concentration!

Mass of NaOH = Mass Conc × Volume
= 0.250 × 1.25
= 0.3125 grams

Mol of NaOH = Mass ÷ M_r
= 0.3125 ÷ (23 + 16 + 1)
= 0.00781 mol

Concept Map

A summary of the four relationships learnt so far:



Volume of Gases

- Avogadro's Law states that **one mole** of any gas occupies a **volume of 24 dm³** at "room conditions":
  - temperature of **25 °C**
  - pressure of **1 atm**

- Two things to note when applying this rule:
  - applies to **gases only**
  - **regardless** of relative mass

Volume of Gases

- Avogadro's Law applies regardless of relative mass:

… At r.t.p., **1 mole** of carbon dioxide occupies **24 dm³**
… At r.t.p., **1 mole** of nitrogen occupies **24 dm³**
… At r.t.p., **1 mole** of hydrogen occupies **24 dm³**
… At r.t.p., **1 mole** of ammonia occupies **24 dm³**
… At r.t.p., **1 mole** of methane occupies **24 dm³**

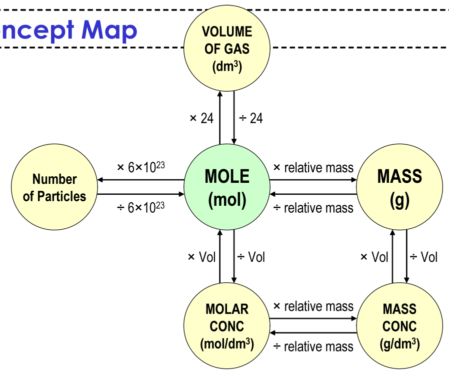
- All that matters is that the substance is a gas!

Volume of Gases

Remember the 23-24-25 Rule:

6×10^{23} particles of gas occupy a volume of 24 dm³ at r.t.p. of 25°C and 1 atm

Concept Map



Limiting & Excess Reagents

- So far we have assumed that, in all reactions, exact proportions of reactants are used, i.e. there is no leftover of any chemical.

- In reality, this is often not so. Usually, one reactant would be leftover while the other reactant is completely used up.

- The chemical that is leftover is known as the **excess reactant**, while the chemical that was completely used up is known as the **limiting reactant**.

Limiting & Excess Reagents

Observe the chemical reaction below:

$$X + Y \longrightarrow Z$$

What will happen if I have three moles of chemical **X** and four moles of chemical **Y**?

3 moles

X: 3 moles
Y: 3 moles

(i) How many moles of **Z** will be produced?
(ii) How many moles of **X** & **Y** will actually be used?
(iii) Which chemical is in excess? Which is limiting?
(iv) Calculate the amount of excess.

Y in excess
X is limiting

Y: 1 mole excess

Limiting & Excess Reagents

Observe the chemical reaction below:

A + 2 B ⟶ C

What will happen if I have three moles of chemical **A** and four moles of chemical **B**?

(i) How many moles of **C** will be produced?
(ii) How many moles of **A** & **B** will actually be used?
(iii) Which chemical is in excess? Which is limiting?
(iv) Calculate the amount of excess.

2 moles

A: 2 moles
B: 4 moles

A in excess
B is limiting

A: 1 mole excess

Limiting & Excess Reagents

A SYSTEMATIC APPROACH

"Find the maximum number of moles of ammonia that can be formed with ***42 grams*** *of nitrogen and* ***6 grams*** *of hydrogen."*

We can identify this question as a *limiting reagent question* as data is provided for more than one reactant!

First step is to convert all data into moles.

Limiting & Excess Reagents

A SYSTEMATIC APPROACH

Step 1: Convert all data to moles.

mol of nitrogen = *mass* ÷ M_r
= 42 ÷ 28
= **1.5 mol**

mol of hydrogen = *mass* ÷ M_r
= 6 ÷ 2
= **3.0 mol**

Next, we will compare the moles of each reactant to determine which is insufficient (i.e. limiting).

The smaller number is not necessarily the limiting reagent!

Limiting & Excess Reagents

A SYSTEMATIC APPROACH

Step 2: Identify the limiting reagent.

$N_2 + 3H_2 \longrightarrow 2NH_3$

N_2	$3H_2$
Present: 1.5 mol	Present: 3.0 mol
Requires: 1.0 mol	Requires: 4.5 mol
Excess!	Limiting!

1.5 mol of nitrogen ***is present which*** *will require 4.5 mol of hydrogen* ***for complete reaction. Since only 3.0 mol of hydrogen is present, there is limiting hydrogen.***

3.0 mol of hydrogen ***is present, which will*** *require 1.0 mol of nitrogen* ***for complete reaction. Since 1.5 mol of nitrogen is present, there is excess nitrogen.***

Limiting & Excess Reagents

A SYSTEMATIC APPROACH

Step 3: Apply mole ratio to the limiting reagent.

$N_2 + 3H_2 \longrightarrow 2NH_3$

Since hydrogen is the limiting reagent, we will apply the mole ratio to hydrogen.

mol of ammonia = *2/3 mol of hydrogen*
= *2/3 (3)*
= ***2.00 mol (3 s.f.)***