Stoichiometric Relationships Part one (answers)

IB CHEMISTRY SL/HL

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese	Sulfur		Carbon	Helium	Manganese
54.938045	32.065		12.0107	4.002602	54.938045

Syllabus objectives:

Understandings:

- Atoms of different elements combine in fixed ratios to form compounds, which have different properties from their component elements.

- Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.

- Mixtures are either homogeneous or heterogeneous.

- The mole is a fixed number of particles and refers to the amount, *n*, of substance.

- Masses of atoms are compared on a scale relative to ^{12}C and are expressed as relative atomic mass (A_r) and relative formula/molecular mass (M_r). Molar mass (M) has the unit g mol^{-1}.

- The empirical formula and molecular formula of a compound give the simplest ratio and the actual number of atoms present in a molecule respectively.

- Reactants can be either limiting or excess.

- The experimental yield can be different from the theoretical yield.

Applications and skills:

- Deduction of chemical equations when reactants and products are specified.

- Application of the state symbols (s), (l), (g) and (aq) in equations.

- Explanation of observable changes in physical properties and temperature during changes of state.

- Calculation of the molar masses of atoms, ions, molecules and formula units.

- Solution of problems involving the relationships between the number of particles, the amount of substance in moles and the mass in grams.

- Interconversion of the percentage composition by mass and the empirical formula.

- Determination of the molecular formula of a compound from its empirical formula and molar mass.

- Obtaining and using experimental data for deriving empirical formulas from reactions involving mass changes.

- Solution of problems relating to reacting quantities, limiting and excess reactants, theoretical, experimental and percentage yields.

Elements, compounds and mixtures

- All substances are made up of one or more elements.
- An element is a substance that cannot be broken down into a simpler substance by chemical means.
- All known elements are included on the periodic table which is shown below.

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** 1.01						Atomic number **Element** Relative atomic mass											2 **He** 4.00
2	3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.63	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.90
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.96	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
6	55 **Cs** 132.91	56 **Ba** 137.33	57 † **La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.20	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	89 ‡ **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (269)	107 **Bh** (270)	108 **Hs** (269)	109 **Mt** (278)	110 **Ds** (281)	111 **Rg** (281)	112 **Cn** (285)	113 **Uut** (286)	114 **Uuq** (289)	115 **Uup** (288)	116 **Uuh** (293)	117 **Uus** (294)	118 **Uuo** (294)
			†	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97	
			‡	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)	

Compounds

- A compound is formed from two or more different elements chemically joined in a fixed ratio.
- Compounds have different properties from the elements that they are made from.



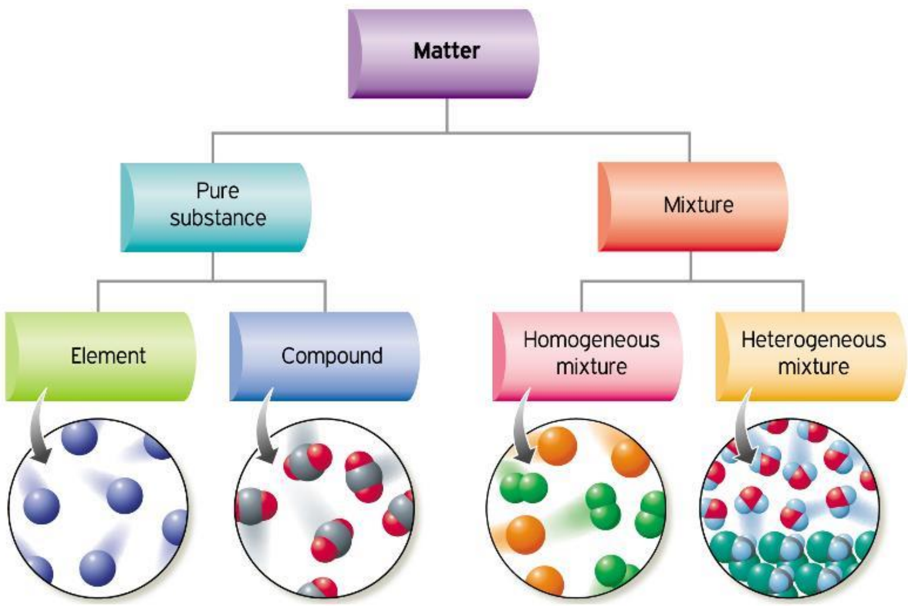
$$\text{sodium} + \text{chlorine} \rightarrow \text{sodium chloride}$$

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

- Note that the properties of the compound above (NaCl) are very different from the elements that it is made from. Sodium is a very reactive metal and chlorine is a poisonous gas. The product formed, NaCl, is safe for human consumption in small amounts.

Mixtures

- Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.
- Mixtures can be either homogeneous or heterogeneous.
- A homogeneous mixture has the same uniform appearance and composition throughout (salt solution).
- A heterogeneous mixture consists of visibly different substances or phases (sand and water).
- Matter can be divided into pure substances or mixtures, as can be seen in the flow chart below.



Exercises:

1. Distinguish between an element and compound

An element is composed of one type of atom, whereas a compound is composed of two or more different types of atom chemically combined.

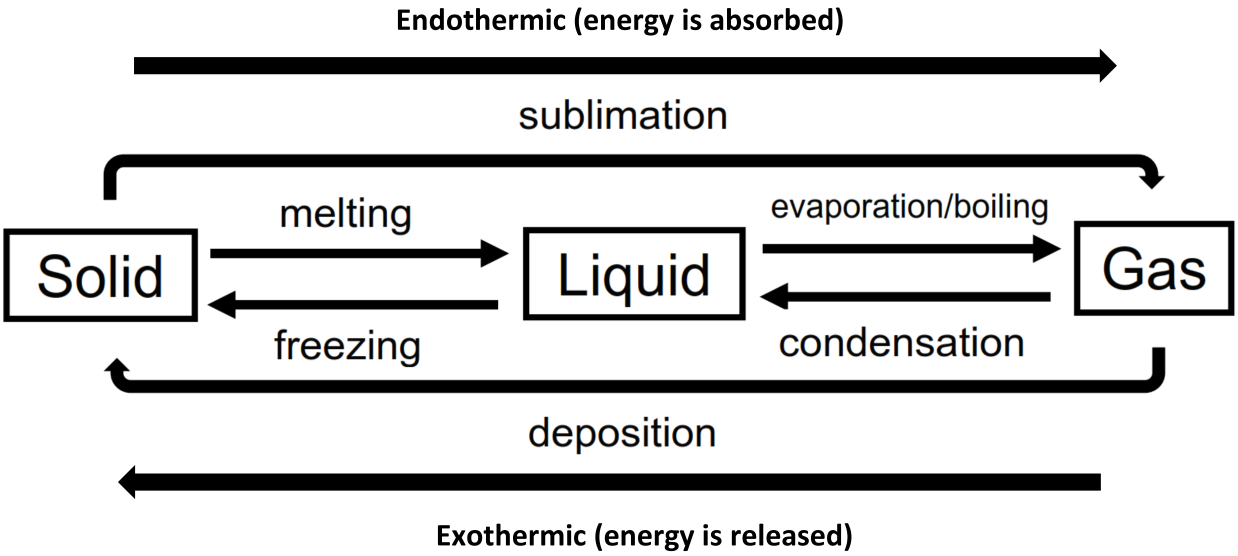
2. Distinguish between a homogenous and heterogeneous mixture.

Homogeneous mixtures have the same composition throughout whereas heterogeneous mixtures do not have the same composition throughout.

States of matter

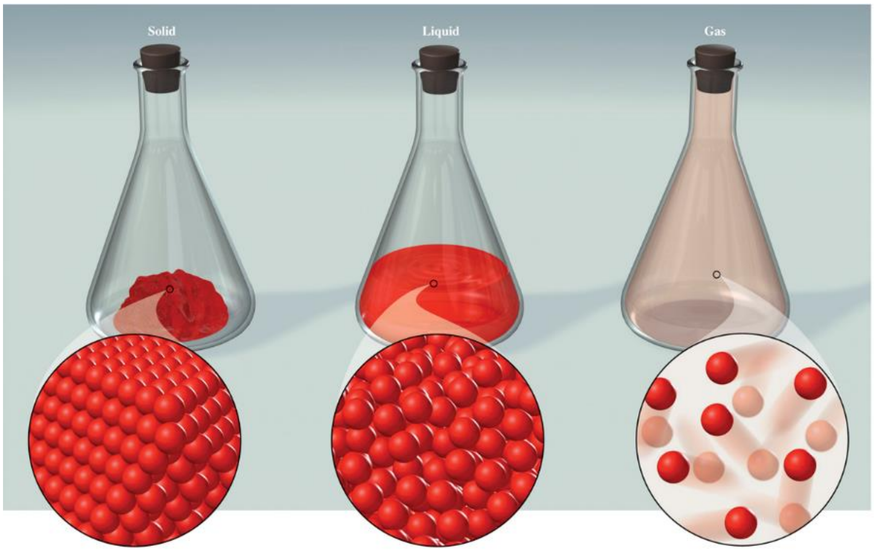
The changes of state are shown below.

- Melting is the change of state from a solid to a liquid.
- Freezing is the change of state from a liquid to a solid.
- Evaporating is the change of state form a liquid to a gas.
- Condensing is the change of state from a gas to a liquid.
- Sublimation is the change of state from a solid to a gas.
- Deposition is the change of state from a gas to a solid.



Particle models of solids, liquids and gases

- The particle models of a solid, liquid and gas are shown below.



Complete the table to show the properties of the following states of matter.

Property	solid	liquid	gas
shape	Fixed shape	No fixed shape	Have the same shape as the container
volume	Fixed volume	Fixed volume	No fixed volume
compressibility	Cannot be compressed	Cannot be compressed	Can be compressed
fluidity	Cannot flow	Can flow	Can flow

Physical and chemical changes

- In a physical change, no new substances are produced.
- The melting of ice is a physical change and can be represented by the following equation:

$$H_2O_{(s)} \rightarrow H_2O_{(l)}$$

- Evaporation of bromine:

$$Br_{2(l)} \rightarrow Br_{2(g)}$$

- Sublimation of iodine:

$$I_{2(s)} \rightarrow I_{2(g)}$$

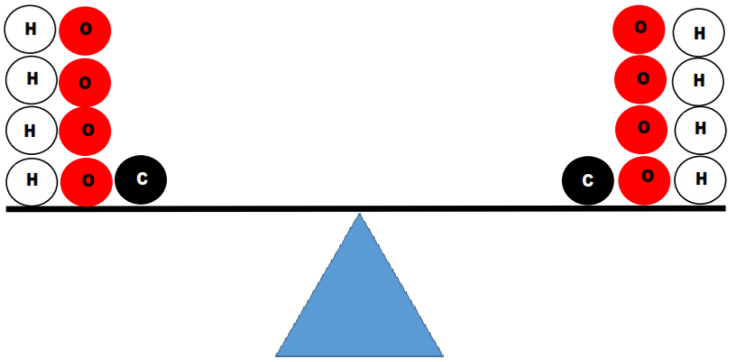
- A chemical change results in the formation of new chemical substances.
- In a chemical reaction, the atoms in the reactants are rearranged to form new products.
- Example:

$$\mathbf{CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}}$$

- The combustion of methane (shown in the equation above) is a chemical change as new chemical substances are formed (CO_2 and H_2O).

Balancing chemical equations

- The law of the conservation of mass states that mass (and therefore atoms) are conserver in a chemical reaction.
- Therefore, there must be the same number of each type of atom in the reactants and products, as shown in the diagram below.



$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

- To balance a chemical equation, we can only change the numbers in front of the reactants or products. These are called coefficients.

Example 1:

- There is one Na atom in the reactants and one in the products. However, there are two Cl atoms in the reactants but only one in the products.

$$Na_{(s)} + Cl_{2(g)} \rightarrow NaCl_{(s)}$$

Na 1	Na 1
Cl 2	Cl 1

Write the balanced equation:

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Example 2:

$$CaCO_{3(s)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Write the balanced equation:

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

State symbols

- State symbols show the physical state (solid, liquid, gas or aqueous) of the reactants and products in a chemical equation.

  (s) – solid

  (l) – liquid

  (g) – gas

  (aq) – aqueous (in solution)

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

Exercise:

Balance the following chemical equations. When each equation is balanced, calculate the sum of coefficients in the equations.

Answers:

1) $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

Sum of coefficients: 6

2) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$

Sum of coefficients: 13

3) $2CH_3OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 4H_2O_{(l)}$

Sum of coefficients: 11

4) $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

Sum of coefficients: 5

5) $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

Sum of coefficients: 6

6) $2NaCl_{(aq)} + CaO_{(aq)} \rightarrow CaCl_{2(aq)} + Na_2O_{(aq)}$

Sum of coefficients: 5

7) $8Al_{(s)} + 3Fe_3O_{4(s)} \rightarrow 4Al_2O_{3(s)} + 9Fe_{(s)}$

Sum of coefficients: 24

8) $Mg_3N_{2(s)} + 4H_2SO_{4(aq)} \rightarrow 3MgSO_{4(aq)} + (NH_4)_2SO_{4(aq)}$

Sum of coefficients: 9

9) $Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 2Fe_{(s)} + 3CO_{(g)}$

Sum of coefficients: 9

10) $2Al(OH)_{3(s)} + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 6H_2O_{(l)}$

Sum of coefficients: 12

The mole (n) amount of chemical substance

- The mole (mol - symbol n) is a unit of measurement used to measure the amount of a chemical substance.
- A mole of a substance has the same number of particles as 12.00 g of ^{12}C
- The number of particles (atoms, ions or molecules) in a mole is equal to Avogadro's constant (L) which is:

$$\mathbf{6.02 \times 10^{23}}$$

1. How many apples would you have if you had a mole of apples?

6.02×10^{23} apples

2. How many eggs would you have if you had two moles of eggs?

$2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$ eggs

3. How much money would you have if you had ten moles of dollars?

$10 \times 6.02 \times 10^{23} = 6.02 \times 10^{24}$ dollars

How big is a mole?

602, 000,000,000,000,000,000,000 (six hundred and two sextillion)

Why do we use the mole in chemistry?

- Atoms are very small – for example a sheet of aluminium foil is approximately 100,000 atoms thick.
- Because they are so small, it is almost impossible to count atoms, so we use the mole concept to 'count' atoms.
- For example, if you have one mole of copper atoms, then you have 6.02×10^{23} copper atoms.

Molar mass (M)

- The molar mass (M) is the mass of one mole of a substance in grams.
- The unit for molar mass is $gmol^{-1}$

Exercise: Use the periodic table to find the molar mass of the following elements:

Element symbol	Molar mass ($gmol^{-1}$)	Element symbol	Molar mass ($gmol^{-1}$)	Element symbol	Molar mass ($gmol^{-1}$)
C	12.01	S	32.07	I	126.90
Ne	20.18	Se	78.96	Pb	207.20
Mg	24.31	Rb	85.47	U	238.04

How to determine the molar mass of a compound

Example: Determine the molar mass of H_2O

- H_2O is composed of 2 H atoms and 1 O atom. Find the relative atomic mass (A_r) of the elements from the periodic table and add them together to get the molar mass.
- $(2 \times 1.01) + (1 \times 16.00) = 18.02$
- The molar mass of H_2O is 18.02 $gmol^{-1}$
- This means that one mole of H_2O has a mass of 18.02 g

Exercise: determine the molar mass of the following:

Substance	Molar mass ($gmol^{-1}$)	Substance	Molar mass ($gmol^{-1}$)	Substance	Molar mass ($gmol^{-1}$)
H_2	2.02	CO_2	44.01	$CaCl_2$	110.98
O_2	32.00	HCl	36.46	Al_2O_3	101.96
Cl_2	70.90	CH_4	16.05	NH_4NO_3	80.04
I_2	253.80	NH_3	17.04	$Al_2(SO_4)_3$	342.15

Calculations involving moles (n), mass (m) and molar mass (M)

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

Symbol	Meaning
n	amount in moles (mol)
m	mass (g)
M	molar mass ($gmol^{-1}$)

This equation can be rearranged to find mass (m) and molar mass (M):

$$m = n \times M \qquad M = \frac{m}{n}$$

Excercises:

1) Calculate the mass in grams of the following: (use the equation n = nM)

a) 3.00 mol NaOH

$3.00 \times 40.00 = 1.20 \times 10^2$ g

b) 0.100 mol C_3H_8

$0.100 \times 44.11 = 4.41$ g

c) 0.400 $CuSO_4$

$0.400 \times 159.61 = 63.8$ g

d) 100.0 mol SO_3

$100.0 \times 80.07 = 8.01 \times 10^3$ g

e) 0.270 mol HNO_3

$0.270 \times 63.01 = 17.0$ g

f) 0.600 mol $CaCl_2$

$0.600 \times 110.98 = 66.6$ g

g) 3.56 mol Al_2O_3

$3.56 \times 101.96 = 363$ g

h) 2.40 mol NH_4NO_3

$2.40 \times 80.04 = 192$ g

i) 0.850 mol $Al_2(SO_4)_3$

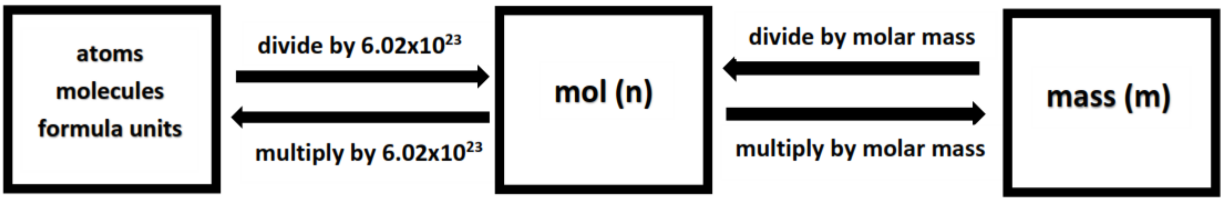
$0.850 \times 342.15 = 291$ g

j) 0.0593 mol Fe_2O_3

$0.0593 \times 159.69 = 9.47$ g

The relationship between number of particles, mol (n) and mass (m)

- One mole of any substance contains 6.02×10^{23} particles (atoms, molecules, formula units).
- The molar mass (M) of a substance is the mass (g) of one mole of a substance.



Example:

1) Calculate the number of H_2O molecules in 18.02 g of pure water.

- Firstly, convert to moles:

$$n = \frac{m}{M} \qquad n = \frac{18.02}{18.02} = 1 \text{ mol } H_2O$$

- Secondly, convert to number of molecules:

  One mole of any substance contains 6.02×10^{23} molecules

  1 mol of H_2O contains 6.02×10^{23} H_2O molecules

2) Calculate the mass of **one molecule** of H_2O:

One mole of H_2O (6.02×10^{23} H_2O molecules) has a mass of 18.02 g

One molecule has a mass of $\frac{18.02}{6.02 \times 10^{23}} = 2.99 \times 10^{-23}$ g

3) Determine the number of H atoms in one mol of H_2O.

One molecule of H_2O is composed of 2 H atoms and 1 O atom.

One mole of H_2O has 6.02×10^{23} H_2O molecules

$2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$ H atoms

Exercises:

Answers:

1) One mol = 6.02×10^{23} molecules. Multiply amount in mol by 6.02×10^{23}

a) 0.500 mol CH_4 — 3.01×10^{23} molecules CH_4

b) 0.750 mol SO_2 — 4.52×10^{23} molecules SO_2

c) 1.08 mol C_2H_5OH — 6.50×10^{23} molecules C_2H_5OH

d) 2.50 mol C_3H_8 — 1.51×10^{24} molecules C_3H_8

e) 1.45×10^{-3} mol NH_3 — 8.73×10^{20} molecules NH_3

2) Multiply the amount in mol by the number of atoms in the molecule

a) 0.500 mol CH_4 5 atoms — $3.01 \times 10^{23} \times 5 = 1.51 \times 10^{24}$

b) 0.750 mol SO_2 3 atoms — $4.52 \times 10^{23} \times 3 = 1.36 \times 10^{24}$

c) 1.08 mol C_2H_5OH 9 atoms — $6.50 \times 10^{23} \times 9 = 5.85 \times 10^{24}$

d) 2.50 mol C_3H_8 11 atoms — $1.51 \times 10^{24} \times 11 = 1.66 \times 10^{25}$

e) 1.45×10^{-3} mol NH_3 4 atoms — $8.73 \times 10^{20} \times 4 = 3.49 \times 10^{21}$

3) Count the number of hydrogen atoms in a molecule and multiply by Avogadro's constant, then multiply by the amount in mol.

a) 0.750 mol CH_4 — $6.02 \times 10^{23} \times 4 \times 0.750 = 1.81 \times 10^{24}$ H atoms

b) 1.24 mol C_2H_5OH — $6.02 \times 10^{23} \times 6 \times 1.24 = 4.48 \times 10^{24}$ H atoms

c) 0.913 mol C_3H_8 — $6.02 \times 10^{23} \times 8 \times 0.913 = 4.40 \times 10^{24}$ H atoms

d) 2.45 mol C_5H_{10} — $6.02 \times 10^{23} \times 10 \times 2.45 = 1.47 \times 10^{25}$ H atoms

e) 6.90×10^{-4} mol NH_3 — $6.02 \times 10^{23} \times 3 \times 6.90 \times 10^{-4} = 1.25 \times 10^{21}$ H atoms

4) Count the number of ions, then multiply by 6.02×10^{23}, then multiply by the amount in mol.

a) 1.00 mol of NaCl (Na^+ Cl^-) $6.02 \times 10^{23} \times 2 \times 1.00 = 1.20 \times 10^{24}$ ions

b) 0.500 mol of Na_2O ($2 \times Na^+$ O^{2-}) $6.02 \times 10^{23} \times 3 \times 0.500 = 9.03 \times 10^{23}$ ions

c) 1.45 mol of $MgCl_2$ (Mg^{2+} $2 \times Cl^-$) $6.02 \times 10^{23} \times 3 \times 1.45 = 2.62 \times 10^{24}$ ions

5)

a) First calculate amount in mol, then multiply by Avogadro's constant

M_r C_2H_5OH = 46.07 $gmol^{-1}$

n = m÷M = $2.30 \times 10^{-3} \div 46.07 = 4.99 \times 10^{-5}$ mol C_2H_5OH

$4.99 \times 10^{-5} \times 6.02 \times 10^{23} = 3.00 \times 10^{19}$ molecules C_2H_5OH

b) One mol of ethane has a molar mass of 30.07 $gmol^{-1}$ = 6.02×10^{23} molecules

Mass of one molecule = $30.07 \div 6.02 \times 10^{23} = 5.00 \times 10^{-23}$g

c) Divide number of molecules by Avogadro's constant to get amount in mol

$1.8 \times 10^{22} \div 6.02 \times 10^{23} = 0.0299$ mol O_2

d) Divide number of molecules by Avogadro's constant to get amount in mol

$3.01 \times 10^{23} \div 6.02 \times 10^{23} = 0.500$ mol H_2O

Multiply mass x molar mass to get grams

m = nM = $0.500 \times 18.02 = 9.01$g H_2O

e) Multiply amount in mol by Avogadro's constant to get molecules

$0.835 \times 6.02 \times 10^{23} = 5.03 \times 10^{23}$ molecules of I_2

One molecule of I_2 = 2 atoms of iodine

$5.03 \times 10^{23} \times 2 = 1.01 \times 10^{24}$ iodine atoms

Empirical formula and molecular formula

- Empirical formula is defined as the lowest whole number ratio of atoms in a compound.
- Molecular formula is the actual number of atoms in a compound.

Example:

- Butane has the molecular formula C_4H_{10}
- The empirical formula is C_2H_5 – how was this determined?

Exercise:

State the empirical formula of the following compounds:

1) H_2O_2 HO

2) C_2H_6 CH_3

3) C_2H_8 CH_4

4) $C_6H_{12}O_6$ CH_2O

5) $C_{20}H_{14}O_4$ $C_{10}H_7O_2$

Concept check:

Explain, giving an example, the difference between empirical and molecular formula.

Empirical formula is the lowest whole number ratio of atoms in a compound.

Molecular formula is the actual number of atoms in a compound.

Calculating empirical formula from percentage composition by mass

Example:

The relative molecular mass of aluminium chloride is 267 and its composition by mass is 20.3% aluminium (Al) and 79.7% chlorine (Cl).
Determine the empirical and molecular formula of aluminium chloride.

(i) Check that the % add up to 100 %

(ii) Divide the % of each element by its relative atomic mass.

(iii) Divide each number in part (ii) by the smallest ratio - this will give you the empirical formula of the compound.

(iv) To find the molecular formula from the empirical formula – determine the mass of the empirical formula and divide the molecular formula by the mass of the empirical formula.

Exercises:

1) Define the terms empirical formula and molecular formula.

Empirical formula is the lowest whole number ratio of atoms in a compound.

Molecular formula is the actual number of atoms in a compound.

2) Compound **B** has the following percentage composition by mass: C 26.7%, O 71.1% and H 2.2%. Calculate the empirical formula of compound **B**.

C	H	O
26.7	2.2	71.1
12.01	1.01	16.00
2.22	2.2	4.44
2.2	2.2	2.2
1	1	2

Empirical formula: CHO_2

3) Compound **C** has the following percentage composition by mass: 48.6% C, 10.8% H, 21.6% O and 18.9% N. Calculate the empirical formula of compound **C.**

C	H	O	N
48.6	10.8	21.6	18.9
12.01	1.01	16.00	14.01
4.04	10.7	1.35	1.35
1.35	1.35	1.35	1.35
3	8	1	1

Empirical formula: C_3H_8ON

4) Work out the molecular formula of each of the following given the empirical formula and the relative molecular mass:

a CH_2, $M_r = 70$

b OH, $M_r = 34$

c C_2H_5O, $M_r = 90$

a CH_2, $M_r = 70$ $(12.01) + (2 \times 1.01) = 14.03$

$70 \div 14.03 = 5$

$CH_2 \times 5 = C_5H_{10}$

b OH, M_r = 34 (16.00) + (1.01) = 17.01

$34 \div 17.01 = 2$

$OH \times 2 = H_2O_2$

c C_2H_5O, M_r = 90 (2 × 12.01) + (5 × 1.01) + (16.00) = 45.07

$90 \div 45.07 = 2$

$C_2H_5O \times 2 = C_4H_{10}O_2$

5) An organic compound A contains 62.0% by mass of carbon, 24.1% by mass of nitrogen, the remainder being hydrogen.

a) Determine the percentage by mass of hydrogen and the empirical formula of **A**.

C	N	H
62.0	24.1	13.9
12.01	14.01	1.01
5.16	1.72	13.8
1.72	1.72	1.72

Empirical formula: C_3NH_8

b) The relative molecular mass of **A** is 116. Determine the molecular formula of **A**.

(3 × 12.01) + (14.01) + (8 × 1.01) = 58.12

$116 \div 58.12 = 2$

$2 \times C_3NH_8 = C_6N_2H_{16}$

Molecular formula: $C_6N_2H_{16}$

Percentage composition by mass

- Percentage composition by mass is the percentage by mass of elements in a compound.

Example: Find the percentage by mass of carbon in ethanol (C_2H_5OH).

(24.02 / 44.08) × 100 = 54.5 %

Exercises:
Calculate the percentage by mass of carbon in the following:

a) CO_2
(12.01 ÷ 44.01) × 100 = 27.3 %

b) C_2H_6
(24.02 ÷ 30.08) × 100 = 79.9 %

c) $C_6H_5NO_2$
(72.06 ÷ 123.11) × 100 = 58.5 %

d) $C_6H_{12}O_6$
(72.06 ÷ 180.16) × 100 = 40.0 %

e) $C_6H_5COCH_3$
(96.08 ÷ 120.16) × 100 = 80.0 %

Percentage purity

- Percentage purity is the percentage of a pure compound in an impure sample.

$$\% \text{ purity} = \frac{\text{mass of pure compound in sample}}{\text{total mass of impure sample}} \times 100$$

Exercise:
A 150.0 g sample of copper ore contains 87.3 g of pure copper. Calculate the percentage purity.

$$\% \text{ purity} = \frac{87.3}{150.0} \times 100 = 58.2 \ \%$$

Stoichiometry

Molar ratios

- The coefficients in a balanced chemical equation tell us the molar ratios of reactants and products.

$$2A + 3B \rightarrow C + 2D$$

- 2 mol of A react with 3 mol of B to form 1 mol of C and 2 mol of D

Exercises:

State the molar ratios in the following chemical equations:

1) $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

   a) $CH_4 : H_2O$ 1 : 2

   b) $CH_4 : CO_2$ 1 : 1

   c) $O_2 : CO_2$ 2 : 1

2) $C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$

   a) $C_3H_8 : CO_2$ 1 : 3

   b) $C_3H_8 : H_2O$ 1 : 4

   c) $O_2 : H_2O$ 5 : 4

3) $2CH_3OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 4H_2O_{(l)}$

   a) $CH_3OH : H_2O$ 2 : 4

   b) $CH_3OH : CO_2$ 2 : 2

   c) $CH_3OH : O_2$ 2 : 3

4) $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$

   a) $Mg : HCl$ 1 : 2

   b) $Mg: H_2$ 1 : 1

   c) $HCl : MgCl_2$ 2 : 1

5) $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

   a) $CaCO_3 : CO_2$ 1 : 1

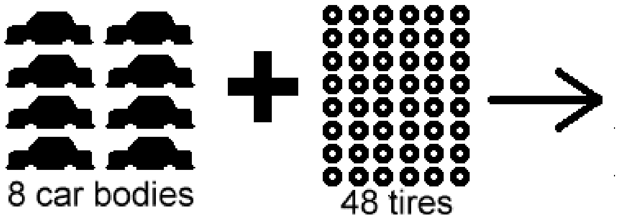
   b) $CaCO_3 : CaCl_2$ 1 : 1

   c) $HCl : CO_2$ 2 : 1

   d) $HCl : CaCl_2$ 2 : 1

Limiting reactant (reagent)

- The limiting reactant (reagent) is the reactant that is completely used up during a chemical reaction.
- The reactant that is in excess is the reactant that is not completely used up during the chemical reaction - there is some of this reactant left over at the end of the reaction.



- How many cars can be made with 8 car bodies and 48 tires? 8 cars
- Which is the limiting reactant? Car bodies
- Which is excess? Tires

Exercises:

1) 2 mol of CH_4 are reacted with excess oxygen (O_2) according to the following equation. Determine the maximum amount (in mol) of CO_2 and H_2O that can be produced.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

2 mol CO_2

4 mol H_2O

2) 3 mol of C_3H_8 is reacted with excess oxygen (O_2) according to the following equation. Determine the maximum amount (in mol) of CO_2 and H_2O that can be produced.

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$$

9 mol CO_2
12 mol H_2O

3) 1.5 mol of CH_3OH is reacted with excess oxygen (O_2) according to the following equation. Determine the maximum amount (in mol) of CO_2 and H_2O that can be produced.

$$2CH_3OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 4H_2O_{(l)}$$

1.5 mol CO_2
3 mol H_2O

4) 50.0 g of Mg is added to excess HCl. Determine the mass, in g, of $MgCl_2$ produced.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

n Mg = 50.0 / 24.31 = 2.06 mol

molar ratio Mg : $MgCl_2$ = 1:1

m $MgCl_2$ = 2.06 × 95.21 = 196 g

5) 75.0 g of $CaCO_3$ is added to excess HCl. Determine the mass, in g, of $CaCl_2$ and H_2O produced.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

n $CaCO_3$ = (75.0 / 100.09) = 0.749 mol

Molar ratio $CaCO_3$: $CaCl_2$: H_2O = 1:1:1

m $CaCl_2$ = 0.749 × 110.98 = 83.1 g

m H_2O = 0.749 × 18.02 = 13.5 g

6) A 50.6 g sample of $Mg(OH)_2$ is reacted with 45.0 g of HCl. Identify which reactant is in excess and which is the limiting reactant. Determine the mass (in g) of $MgCl_2$ that can be produced.

$$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$$

n HCl = (45.0 / 36.46) = 1.23 mol

n $Mg(OH)_2$ = 50.6 / 58.32 = 0.867 mol

HCl: 1.23 / 2 = 0.615 (limiting reactant)

$Mg(OH)_2$: 0.867 / 1 = 0.867 (excess)

Molar ratio HCl : $MgCl_2$ = 2:1

m $MgCl_2$ = 0.615 × 95.21 = 58.6 g

7) 30.0 g of ammonium nitrate (NH_4NO_3) and 50.0 g of sodium phosphate (Na_3PO_4) are reacted together. Identify which reactant is in excess and which is the limiting reactant. Determine the maximum mass (in g) of $NaNO_3$ that can be produced.

$$3NH_4NO_3 + Na_3PO_4 \rightarrow (NH_4)_3PO_4 + 3NaNO_3$$

M NH_4NO_3 = 80.04 gmol^{-1}
M Na_3PO_4 = 163.94 gmol^{-1}

n NH_4NO_3 = $\frac{30.0}{80.04}$ = 0.375 mol

0.375 ÷ 3 = 0.125

n Na_3PO_4 = $\frac{50.0}{163.94}$ = 0.305 mol

0.305 ÷ 1 = 0.305

NH_4NO_3 is limiting reactant, Na_3PO_4 is excess reagent

Molar ratio of NH_4NO_3 to $(NH_4)_3PO_4$ is 3:1
0.375 mol of NH_4NO_3 will produce (0.375 ÷ 3) = 0.125 mol $(NH_4)_3PO_4$
m = nM
m $(NH_4)_3PO_4$ = 149.08 × 0.125 = 18.6 g

Percentage yield

- The percentage yield is the actual yield divided by the theoretical yield

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

- The actual yield is the actual amount of product made.
- The theoretical yield is the amount of product made based on the stoichiometry of the reaction.

Example:
Aluminium reacts with oxygen according to the following equation. Determine the percentage yield if 20.0 g of Al reacts with excess oxygen to produce 32.7 g of Al_2O_3.

$$4Al_{(s)} + 3O_{2(g)} \rightarrow 2Al_2O_{3(s)}$$

M Al = 26.98 $gmol^{-1}$

n = $\frac{20.0}{26.98}$ = 0.741 mol

M Al_2O_3 = 101.96 $gmol^{-1}$
0.741 ÷ 2 = 0.371 mol K_2CO_3
0.371 × 101.96 = 37.8 g

$\%\ \text{yield} = \frac{32.7}{37.8} \times 100 =$ 86.5%

Exercises:
1) A 15.0 g sample of pure K_2O produces 7.62 g of K_2CO_3. Determine the percentage yield of the reaction.

$$4\ K_2O_{(s)} + 2CO_{2(g)} \rightarrow 2K_2CO_{3(s)} + 3O_{2(g)}$$

M K_2O = 71.09 $gmol^{-1}$

n = $\frac{15.0}{71.09}$ = 0.211 mol

M K_2CO_3 = 138.19 $gmol^{-1}$
0.211 ÷ 2 = 0.106 mol K_2CO_3
0.106 × 138.19 = 14.6 g

$\%\ \text{yield} = \frac{7.62}{14.6} =$ 52.2%

2) A 20.0 g sample of pure Fe_3O_4 produces 5.98 g of Fe. Determine the percentage yield of the reaction.

$$Fe_3O_{4(s)} + 4H_{2(g)} \rightarrow 3Fe_{(s)} + 4H_2O_{(l)}$$

M Fe_3O_4 = 231.53 $gmol^{-1}$

$n = \frac{20.0}{231.54}$ = 0.0864 mol

M Fe = 55.85 $gmol^{-1}$
0.0864 × 3 = 0.259 mol Fe
0.259 × 55.85 = 14.5 g

$\%\ \text{yield} = \frac{5.98}{14.5}$ = 41.2%