

Stoichiometric Relationships

Part one

IB CHEMISTRY SL/HL

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

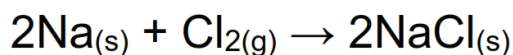
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
1	1 H 1.01																	2 He 4.00	
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18	
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95	
4	19 K 39.10	20 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 (209)	85 Po (210)	86 At (222)	87 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.



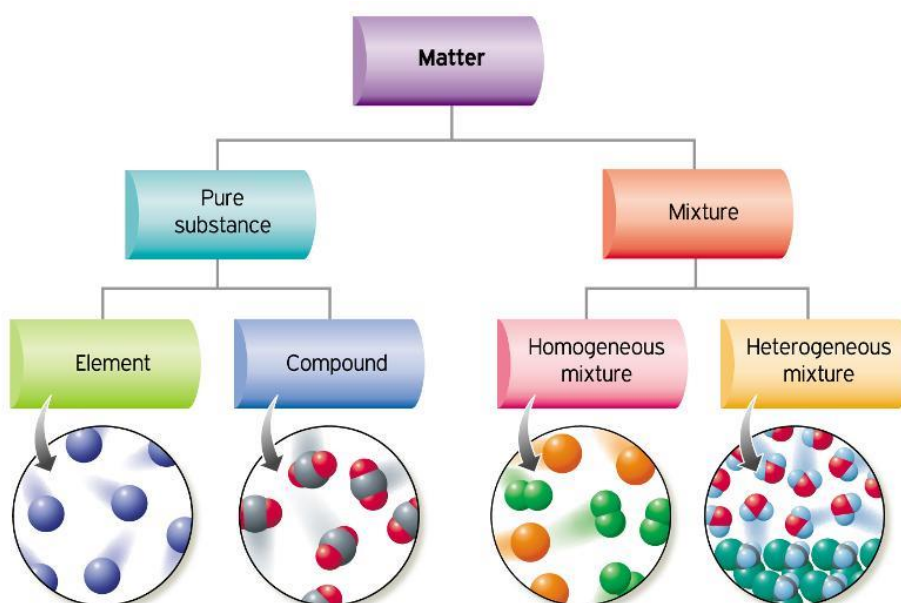
sodium + chlorine → sodium chloride



- 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:

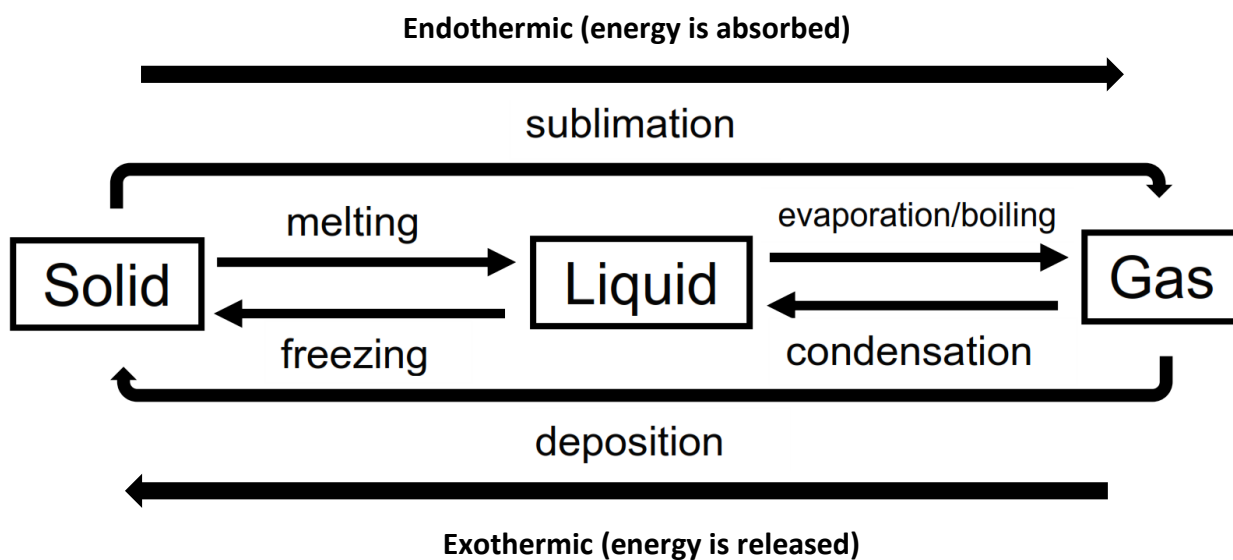
1. Distinguish between an element and compound

2. Distinguish between a homogenous and heterogeneous mixture.

States of matter

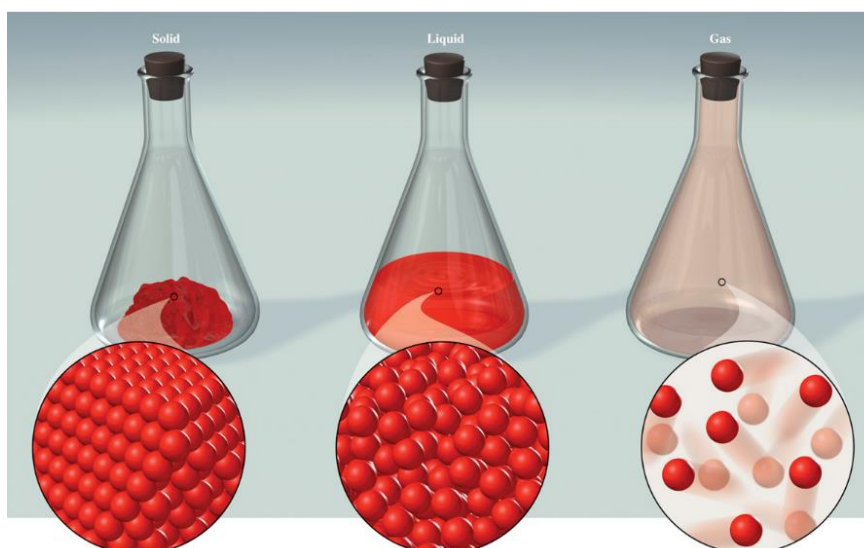
The changes of state are shown below.

- 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 from 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.

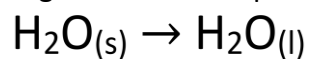


Complete the table to show the properties of the following states of matter.

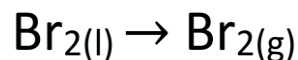
Property	solid	liquid	gas
shape			
volume			
compressibility			
fluidity			

Physical and chemical changes

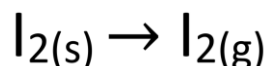
- In a physical change, no new substances are produced.
- The melting of ice is a physical change and can be represented by the following equation:



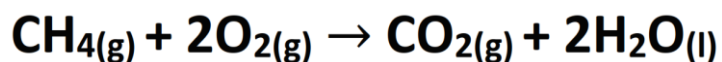
- Evaporation of bromine:



- Sublimation of iodine:



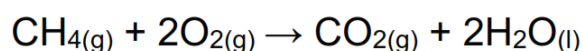
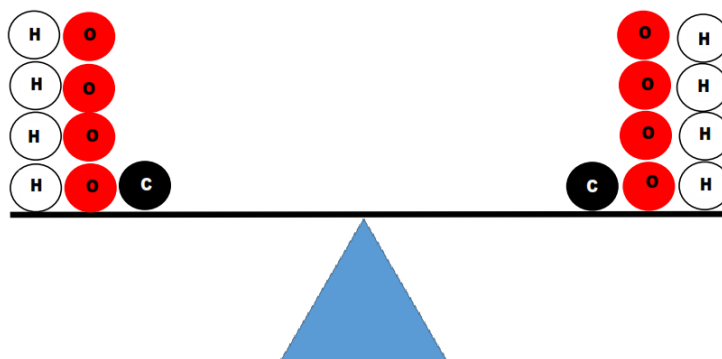
- A chemical change results in the formation of new chemical substances.
- In a chemical reaction, the atoms in the reactants are rearranged to form new products.
- Example:



- 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

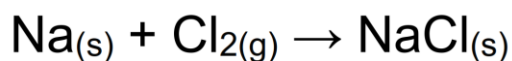
- The law of the conservation of mass states that mass (and therefore atoms) are conserved in a chemical reaction.
- Therefore, there must be the same number of each type of atom in the reactants and products, as shown in the diagram below.



- 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.

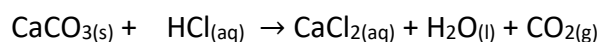


Na 1 Na 1

Cl 2 Cl 1

Write the balanced equation:

Example 2:



Write the balanced equation:

State symbols

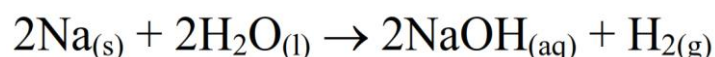
- State symbols show the physical state (solid, liquid, gas or aqueous) of the reactants and products in a chemical equation.

(s) – solid

(l) – liquid

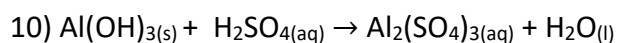
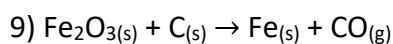
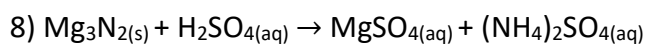
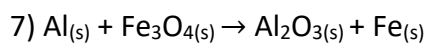
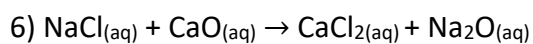
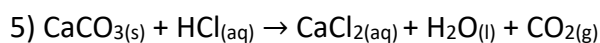
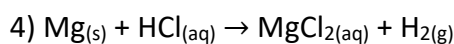
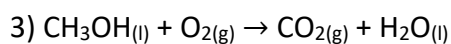
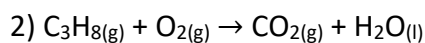
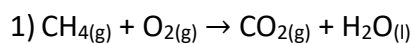
(g) – gas

(aq) – aqueous (in solution)



Exercise:

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



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:

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

1. How many apples would you have if you had a mole of apples?
2. How many eggs would you have if you had two moles of eggs?
3. How much money would you have if you had ten moles of dollars?

How big is a mole?

602, 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 g mol^{-1}

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

Element symbol	Molar mass (g mol^{-1})	Element symbol	Molar mass (g mol^{-1})	Element symbol	Molar mass (g mol^{-1})
C		S		I	
Ne		Se		Pb	
Mg		Rb		U	

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 g mol^{-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 (g mol^{-1})	Substance	Molar mass (g mol^{-1})	Substance	Molar mass (g mol^{-1})
H_2		CO_2		CaCl_2	
O_2		HCl		Al_2O_3	
Cl_2		CH_4		NH_4NO_3	
I_2		NH_3		$\text{Al}_2(\text{SO}_4)_3$	

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 (g mol ⁻¹)

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

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

Exercises:

1) Calculate the mass in grams of the following:

a) 3.00 mol NaOH

f) 0.600 mol CaCl₂

b) 0.100 mol C₃H₈

g) 3.56 mol Al₂O₃

c) 0.400 CuSO₄

h) 2.40 mol NH₄NO₃

d) 100.0 mol SO₃

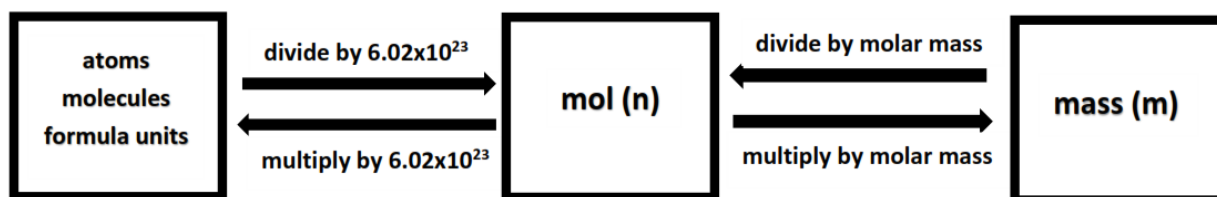
i) 0.850 mol Al₂(SO₄)₃

e) 0.270 mol HNO₃

j) 0.0593 mol Fe₂O₃

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 H₂O molecules in 18.02 g of pure water.

- Firstly, convert to moles:

$$n = \frac{m}{M} \quad n = \frac{18.02}{18.02} = 1 \text{ mol H}_2\text{O}$$

- Secondly, convert to number of molecules:

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

1 mol of H₂O contains 6.02×10^{23} H₂O molecules

2) Calculate the mass of **one molecule** of H₂O:

One mole of H₂O (6.02×10^{23} H₂O 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 H₂O.

One molecule of H₂O is composed of 2 H atoms and 1 O atom.

One mole of H₂O has 6.02×10^{23} H₂O molecules

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

Exercises:

1) Calculate the number of molecules in the following:

a) 0.500 mol CH₄

b) 0.750 mol SO₂

c) 1.08 mol C₂H₅OH

d) 2.50 mol C₃H₈

e) 1.45×10^{-3} mol NH₃

2) Calculate the total number of atoms in the following:

a) 0.500 mol CH₄

b) 0.750 mol SO₂

c) 1.08 mol C₂H₅OH

d) 2.50 mol C₃H₈

e) 1.45×10^{-3} mol NH₃

3) Calculate the number of hydrogen atoms in:

a) 0.750 mol CH₄

b) 1.24 mol C₂H₅OH

c) 0.913 mol C₃H₈

d) 2.45 mol C₅H₁₀

e) 6.90×10^{-4} mol NH₃

4) Calculate the number of ions in:

a) 1.00 mol of NaCl

b) 0.500 mol of Na₂O

c) 1.45 mol of MgCl₂

5) Calculate the following:

a) The number of ethanol molecules in a drop of ethanol (2.30×10^{-3} g)

b) The mass of one molecule of ethane (C₂H₆)

c) The amount (in mol) of O₂ that contains 1.80×10^{22} molecules

d) The mass of 3.01×10^{23} molecules of H₂O

e) The number of iodine atoms in 0.835 mol of I₂

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

2) C_2H_6

3) C_2H_8

4) $C_6H_{12}O_6$

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

Concept check:

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

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.

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**.

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**.

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 $\text{C}_2\text{H}_5\text{O}$, $M_r = 90$

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**.

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

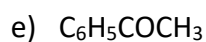
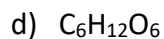
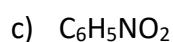
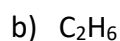
Percentage composition by mass

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

Example: Find the percentage by mass of carbon in ethanol (C₂H₅OH).

Exercises:

Calculate the percentage by mass of carbon in the following:



Percentage purity

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

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

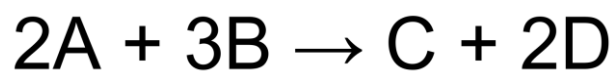
Exercise:

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

Stoichiometry

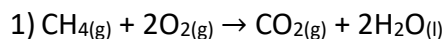
Molar ratios

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



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

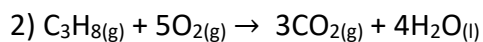
Exercises: State the molar ratios in the following chemical equations:



a) $\text{CH}_4 : \text{H}_2\text{O}$

b) $\text{CH}_4 : \text{CO}_2$

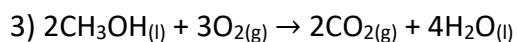
c) $\text{O}_2 : \text{CO}_2$



a) $\text{C}_3\text{H}_8 : \text{CO}_2$

b) $\text{C}_3\text{H}_8 : \text{H}_2\text{O}$

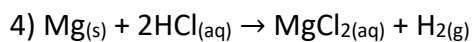
c) $\text{O}_2 : \text{H}_2\text{O}$



a) $\text{CH}_3\text{OH} : \text{H}_2\text{O}$

b) $\text{CH}_3\text{OH} : \text{CO}_2$

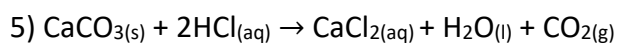
c) $\text{CH}_3\text{OH} : \text{O}_2$



a) $\text{Mg} : \text{HCl}$

b) $\text{Mg} : \text{H}_2$

c) $\text{HCl} : \text{MgCl}_2$



a) $\text{CaCO}_3 : \text{CO}_2$

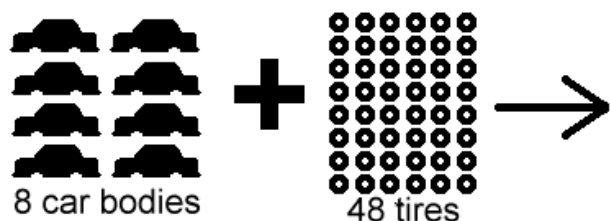
b) $\text{CaCO}_3 : \text{CaCl}_2$

c) $\text{HCl} : \text{CO}_2$

d) $\text{HCl} : \text{CaCl}_2$

Limiting reactant (reagent)

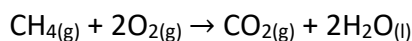
- The limiting reactant (reagent) is the reactant that is completely used up during a chemical reaction.
- The reactant that is in excess is the reactant that is not completely used up during the chemical reaction - there is some of this reactant left over at the end of the reaction.



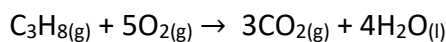
- How many cars can be made with 8 car bodies and 48 tires?
- Which is the limiting reactant?
- Which is excess?

Exercises:

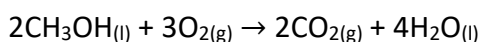
1) 2 mol of CH_4 are reacted with excess oxygen (O_2) according to the following equation. Determine the maximum amount (in mol) of CO_2 and H_2O that can be produced.



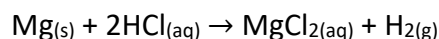
2) 3 mol of C_3H_8 is reacted with excess oxygen (O_2) according to the following equation. Determine the maximum amount (in mol) of CO_2 and H_2O that can be produced.



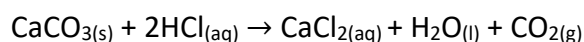
3) 1.5 mol of CH_3OH is reacted with excess oxygen (O_2) according to the following equation. Determine the maximum amount (in mol) of CO_2 and H_2O that can be produced.



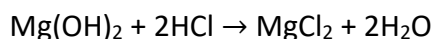
4) 50.0 g of Mg is added to excess HCl. Determine the mass, in g, of MgCl₂ produced.



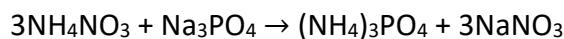
5) 75.0 g of CaCO₃ is added to excess HCl. Determine the mass, in g, of CaCl₂ and H₂O produced.



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 MgCl₂ that can be produced.



7) 30.0 g of ammonium nitrate (NH₄NO₃) and 50.0 g of sodium phosphate (Na₃PO₄) are reacted together. Identify which reactant is in excess and which is the limiting reactant. Determine the maximum mass (in g) of NaNO₃ that can be produced.



Percentage yield

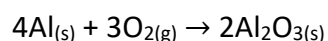
- The percentage yield is the actual yield divided by the theoretical yield

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

- 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.

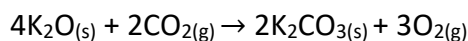
Example:

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 .



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.



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

