Stoichiometric Relationships Part two (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:

- The mole (mol) 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.

Applications and skills:

- 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.

The mole and amount of substance

- The mole, symbol mol, is the SI unit of amount of substance (n).
- It is a measure of the number of specified elementary entities (an elementary entity can refer to an atom, a molecule, an ion, an electron, or any other particle).
- One mole contains exactly $6.02214076 \times 10^{23}$ elementary entities (usually rounded to 6.02×10^{23}).
- This is numerically equal to the Avogadro constant (L or N_A) which is 6.02×10^{23} mol^{-1}

Elementary entity	Number of elementary entities in one mole
Atoms	6.02×10^{23}
Molecules	6.02×10^{23}
Ions	6.02×10^{23}
Formula units	6.02×10^{23}

Relative atomic mass and relative molecular mass

- Relative atomic mass, A_r, is the weighted average mass of the naturally occurring isotopes of an element relative to 1/12 the mass of an atom of carbon-12.
- The relative atomic mass scale is based on the isotope carbon-12 which has a mass of exactly 12 amu.
- Relative molecular mass, M_r, is the weighted average mass of a molecule relative to 1/12 the mass of an atom of ^{12}C.
- The M_r is the sum of the A_r of the atoms in a molecule.
- Both relative atomic mass and relative molecular mass do not have units.
- Relative formula mass is mostly used for compounds that do not form molecules, such as ionic compounds.

Exercise: Calculate the relative molecular mass/formula mass of the following.

1. C_2H_5OH $M_r = 46.08$

2. CH_3COCH_3 $M_r = 58.09$

3. $C_6H_{12}O_6$ $M_r = 180.18$

4. KCl $M_r = 74.55$

5. $MgBr_2$ $M_r = 184.11$

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}
- The molar mass of a substance is numerically equal to its relative atomic mass.
- To convert A_r to M, multiply by the molar mass constant, M_u, which is approximately equal to 1 g mol^{-1}

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. Multiply by the molar mass constant to get the molar mass.

(2 × 1.01) + (1 × 16.00) = 18.02

18.02 × 1 g mol^{-1} = 18.02 g mol^{-1}

The molar mass of H_2O is 18.02 g mol^{-1}

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	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 amount (*n*), mass (*m*) and molar mass (*M*)

- To convert from mass (in g) to amount (in mol), divide the mass of the substance by its molar mass.

$$\textbf{amount (mol)} = \frac{\textbf{mass (g)}}{\textbf{molar mass (g mol}^{-1}\textbf{)}}$$

$$n(\text{mol}) = \frac{m\ (\text{g})}{M\ (\text{g mol}^{-1})} \qquad n = \frac{m}{M}$$

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

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

Excercises:

1. Calculate the amount in mol of the following:

a. 30.00 g Mg

$30.00 \div 24.31 = 1.234$ mol

b. 75.00 g O_2

$75.00 \div 32.00 = 2.344$ mol

c. 26.93 g $CuSO_4$

$26.93 \div 159.61 = 0.1687$ mol

d. 15.00 g NaOH

$15.00 \div 40.00 = 0.3750$ mol

e. 1.78 g C_3H_8

$1.78 \div 44.11 = 0.0404$ mol

f. 45.82 g $CaCl_2$

$45.82 \div 110.98 = 0.4129$ mol

g. 98.36 g Al_2O_3

$98.36 \div 101.96 = 0.9647$ mol

h. 173.81 g NH_4NO_3

$173.81 \div 80.04 = 2.172$ mol

i. 118.62 g $Al_2(SO_4)_3$

$118.62 \div 342.15 = 0.3467$ mol

j. 261.04 g Fe_2O_3

$261.04 \div 159.69 = 1.635$ mol

2. Calculate the mass in grams of the following:

a. 3.00 mol Mg

$3.00 \times 24.31 = 72.93$ g

b. 0.100 mol O_2

$0.100 \times 32.00 = 3.20$ g

c. 0.400 mol $CuSO_4$

$0.400 \times 159.61 = 63.8$ g

d. 9.84 mol NaOH

$9.84 \times 40.00 = 394$ g

e. 0.270 mol C_3H_8

$0.270 \times 44.11 = 11.9$ 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 $Al_2(SO_4)_3$

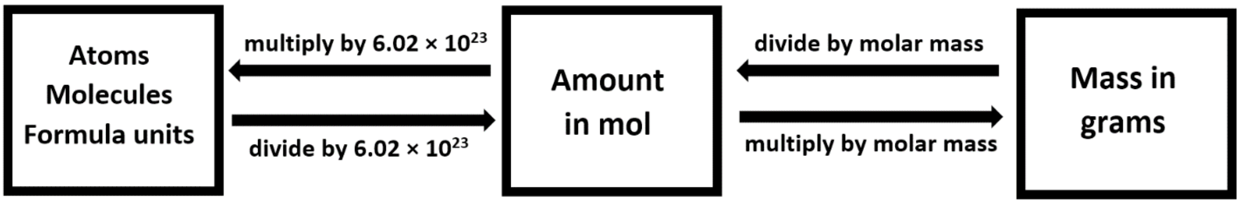
$0.850 \times 342.15 = 291$ g

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.

   First, convert to amount (in mol):

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

   Next, 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:

1. Calculate the number of molecules in the following:

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. Calculate the total number of atoms in the following:

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

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

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

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

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

3. Calculate the number of molecules in the following:

a. 25.00 g of propanone, C_3H_6O — $(25.00 \div 58.09) \times 6.02 \times 10^{23} = 2.59 \times 10^{23}$

b. 50.12 g of ethane, C_2H_6 — $(50.12 \div 30.08) \times 6.02 \times 10^{23} = 1.00 \times 10^{24}$

c. 13.74 g of glucose, $C_6H_{12}O_6$ — $(13.74 \div 180.18) \times 6.02 \times 10^{23} = 4.59 \times 10^{22}$

d. 71.83 g of water, H_2O — $(71.83 \div 18.02) \times 6.02 \times 10^{23} = 2.40 \times 10^{24}$

e. 134.20 g of hexane, C_6H_{14} — $(134.20 \div 86.20) \times 6.02 \times 10^{23} = 9.37 \times 10^{23}$

4. Calculate the number of hydrogen atoms in:

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

5. Calculate the number of ions in:

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

6. Calculate the following:

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

M_r $C_2H_5OH = 46.07$ g mol^{-1}

$n = m \div M = 2.30 \times 10^{-3} \div 46.07 = 4.99 \times 10^{-5}$ 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. The mass of one molecule of ethane (C_2H_6).

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

c. The amount (in mol) of O_2 that contains 1.80×10^{22} molecules.

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

d. The mass of 3.01×10^{23} molecules of H_2O.

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

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

e. The number of iodine atoms in 0.835 mol of I_2

$0.835 \times 6.02 \times 10^{23} = 5.03 \times 10^{23}$ 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 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?
  Divide the 4 and 10 by 2 to give 2 and 5.

Exercise:

State the empirical formula of the following compounds:

1. H_2O_2 HO
2. C_2H_6 CH_3
3. C_4H_8 CH_2
4. $C_6H_{12}O_6$ CH_2O
5. $C_{20}H_{14}O_4$ $C_{10}H_7O_2$

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.

1. Check that the % add up to 100 %

   20.3 % + 79.7 % = 100 %

2. Divide the % of each element by its relative atomic mass.

Al	Cl
$\frac{20.3}{26.98}$	$\frac{79.7}{35.45}$

3. Divide each number in part (2) by the smallest ratio - this will give you the empirical formula of the compound.

Al	Cl
$\frac{0.752}{0.752}$	$\frac{2.25}{0.752}$
1	3

   Empirical formula $AlCl_3$

4. 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.

$$\frac{267}{133.33} = 2.00$$

   Molecular formula Al_2Cl_6

Exercises:

1. 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**.

2. 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.**

4. 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. Determine the molecular formula of each of the following given the empirical formula and the relative molecular mass, M_r

   a. CH_2, $M_r = 70$

   CH_2, $M_r = 70$ (12.01) + (2 × 1.01) = 14.03

   70 ÷ 14.03 = 5

   $CH_2 \times 5 = C_5H_{10}$

b. OH, M_r = 34

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

34 ÷ 17.01 = 2

OH × 2 = H_2O_2

c. C_2H_5O, M_r = 90

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

90 ÷ 45.07 = 2

C_2H_5O × 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 ÷ 58.12 = 2

2 × 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 / 46.08) × 100 = 52.1 %

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

1. CO_2
   (12.01   44.01) × 100 = 27.3 %

2. C_2H_6
   (24.02 ÷ 30.08) × 100 = 79.9 %

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

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

5. $C_6H_5COCH_3$
   (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.

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

Calculating empirical formula from combustion analysis

Menthol is an organic compound composed of C, H and O atoms. The complete combustion of 0.1005 g of menthol produces 0.2829 g of CO_2 and 0.1159 g of H_2O. Calculate the empirical formula of menthol.

1. Calculate the mass of carbon in CO_2 and convert to mol.

Calculate the mass of C in 0.2829 g of CO_2

$$\frac{12.01}{44.01} \times 0.2829 = 0.07720 \text{ g of C}$$

Convert to amount in mol (*n*)

$$n = \frac{0.07720}{12.01} = 6.428 \times 10^{-3} \text{ mol C}$$

2. Calculate the mass of H in H_2O and convert to mol.

Calculate the mass of H in 0.1159 g of H_2O

$$\frac{2.02}{18.02} \times 0.1159 = 0.01299 \text{ g of H}$$

Convert to amount in mol (*n*)

$$n = \frac{0.01299}{1.01} = 0.01286 \text{ mol H}$$

3. Calculate the mass of O by subtracting the mass of carbon and mass of hydrogen from original mass of menthol. Convert to amount in mol.

Calculate the mass of O in 0.1005 g of menthol

0.1005 – 0.07720 – 0.01299 = 0.01031 g O

Convert to amount in mol (*n*)

$$n = \frac{0.01031}{16.00} = 6.444 \times 10^{-4} \text{ mol O}$$

4. Divide each amount by the smallest to get the lowest whole number ratio.

6.428×10^{-3} mol C	**0.01286 mol H**	**6.444×10^{-4} mol O**
6.444×10^{-4}	**6.444×10^{-4}**	**6.444×10^{-4}**
10	**20**	**1**

Empirical formula: $C_{10}H_{20}O$