Structure 1.4 Answers

IB CHEMISTRY SL

Structure 1.4.1

Understandings:

- The mole (mol) is the SI unit of amount of substance. One mole contains exactly the number of elementary entities given by the Avogadro constant.

Learning outcomes:

- Convert the amount of substance, *n*, to the number of specified elementary entities.

Additional notes:

- An elementary entity may be an atom, a molecule, an ion, an electron, any other particle or a specified group of particles.
- The Avogadro constant N_A is given in the data booklet. It has the units mol^{-1}.

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 6.02×10^{23} elementary entities.
- 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}

Structure 1.4.2

Understandings:

- Masses of atoms are compared on a scale relative to ^{12}C and are expressed as relative atomic mass A_r and relative formula mass M_r.

Learning outcomes:

- Determine relative formula masses M_r from relative atomic masses A_r.

Additional notes:

- Relative atomic mass and relative formula mass have no units.
- The values of relative atomic masses given to two decimal places in the data booklet should be used in calculations.

Relative atomic mass and relative formula 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 formula mass, M_r, is the weighted average mass of a substance 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 substance.
- Both relative atomic mass and relative formula mass do not have units.

Exercise: Calculate the relative formula masses 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. $MgBr_2$ $M_r = 184.11$

Structure 1.4.3

Understandings:

- Molar mass M has the units g mol^{-1}.

Learning outcomes:

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

Additional notes:

- The relationship $n = m/M$ is given in the data booklet.

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.

$(2 \times 1.01) + (1 \times 16.00) = 18.02$

$18.02 \times 1 \text{ g mol}^{-1} = 18.02 \text{ g mol}^{-1}$

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.

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

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

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

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

Exercises:

1. Calculate the amount (in mol) of the following.

a. 30.00 g Mg

30.00 ÷ 24.31 = 1.234 mol

b. 75.00 g O_2

75.00 ÷ 32.00 = 2.344 mol

c. 26.93 g $CuSO_4$

26.93 ÷ 159.61 = 0.1687 mol

d. 15.00 g NaOH

15.00 ÷ 40.00 = 0.3750 mol

e. 1.78 g C_3H_8

1.78 ÷ 44.11 = 0.0404 mol

f. 45.82 g $CaCl_2$

45.82 ÷ 110.98 = 0.4129 mol

g. 98.36 g Al_2O_3

98.36 ÷ 101.96 = 0.9647 mol

h. 173.81 g NH_4NO_3

173.81 ÷ 80.04 = 2.172 mol

i. 118.62 g $Al_2(SO_4)_3$

118.62 ÷ 342.15 = 0.3467 mol

j. 261.04 g Fe_2O_3

261.04 ÷ 159.69 = 1.635 mol

2. Calculate the mass in grams of the following.

a. 3.00 mol Mg

3.00 × 24.31 = 72.93 g

b. 0.100 mol O_2

0.100 × 32.00 = 3.20 g

c. 0.400 mol $CuSO_4$

0.400 × 159.61 = 63.8 g

d. 9.84 mol NaOH

9.84 × 40.00 = 394 g

e. 0.270 mol C_3H_8

0.270 × 44.11 = 11.9 g

f. 0.600 mol $CaCl_2$

0.600 × 110.98 = 66.6 g

g. 3.56 mol Al_2O_3

3.56 × 101.96 = 363 g

h. 2.40 mol NH_4NO_3

2.40 × 80.04 = 192 g

i. 0.850 mol $Al_2(SO_4)_3$

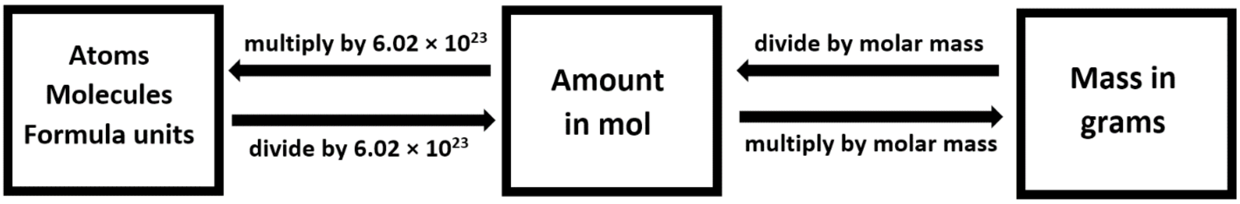
0.850 × 342.15 = 291 g

j. 0.0593 mol Fe_2O_3

0.0593 × 159.69 = 9.47 g

The relationship between number of particles, amount in 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 (in g) of one mole of a substance.



Examples:

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 of H_2O 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 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

4. 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

5. 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

Structure 1.4.4
Understandings:

- The empirical formula of a compound gives the simplest ratio of atoms of each element present in that compound.
- The molecular formula gives the actual number of atoms of each element present in a molecule.

Learning outcomes:

- Interconvert the percentage composition by mass and the empirical formula.
- Determine the molecular formula of a compound from its empirical formula and molar mass.

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?
  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.

$$\begin{array}{cc} \mathrm{Al} & \mathrm{Cl} \\ \frac{20.3}{26.98} & \frac{79.7}{35.45} \end{array}$$

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

$$\begin{array}{cc} \mathrm{Al} & \mathrm{Cl} \\ \frac{0.752}{0.752} & \frac{2.25}{0.752} \\ 1 & 3 \end{array}$$

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.

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

2. 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

3. 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, $M_r = 70$ $(12.01) + (2 \times 1.01) = 14.03$

$70 \div 14.03 = 5$

$CH_2 \times 5 = C_5H_{10}$

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 = $C_4H_{10}O_2$

4. 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 %

Calculating empirical formula from combustion analysis

Example: 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 the original mass of menthol. Convert to amount in mol.

Calculate the mass of O in 0.1005 g of menthol

$$0.1005 - 0.07720 - 0.01299 = 0.01031 \text{ g O}$$

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.

$$\frac{6.428 \times 10^{-3} \text{ mol C}}{6.444 \times 10^{-4}} \qquad \frac{0.01286 \text{ mol H}}{6.444 \times 10^{-4}} \qquad \frac{6.444 \times 10^{-4} \text{ mol O}}{6.444 \times 10^{-4}}$$

10 **20** **1**

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

Structure 1.4.5

Understandings:

- The molar concentration is determined by the amount of solute and the volume of solution.

Learning outcomes:

- Solve problems involving the molar concentration, amount of solute and volume of solution.

Additional notes:

- The use of square brackets to represent molar concentration is required.
- Units of concentration should include g dm^{-3} and mol dm^{-3} and conversion between these.
- The relationship $n = CV$ is given in the data booklet.

Calculating the concentration of a solution

- The concentration of a solution can be expressed in mol dm^{-3} or g dm^{-3}.
- The equation for calculating concentration in mol dm^{-3} is shown below.
- In this equation, **volume must be in dm^3** (to convert from cm^3 to dm^3, divide by 1000).

$$c\ (\text{mol dm}^{-3}) = \frac{\text{amount of solute (mol)}}{\text{volume of solution (dm}^3\text{)}}$$

$$c = \frac{n}{V}$$

c = concentration in mol dm^{-3}
n = amount in mol
V = volume in dm^3

$$n = cV \qquad V = \frac{n}{c}$$

- The equation for calculating the concentration in g dm^{-3} is shown below.

$$c\ (\text{g dm}^{-3}) = \frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3\text{)}}$$

Example: 50.0 g of NaCl are dissolved in 100 cm^3 of water which is then made up to 500.0 cm^3 in a volumetric flask. Calculate the concentration of the solution in mol dm^{-3} and g dm^{-3}.

In g dm^{-3}
c = mass of solute ÷ volume of solution
$c = 50.0 \div (500.0 \div 1000)$
$c = 100.0$ g dm^{-3}

In mol dm^{-3}
Convert from mass to amount in mol
$n = m \div M$
$n = 50.0 \div 58.44 = 0.856$ mol NaCl
$c = n \div V$
$c = 0.856 \div (500 \div 1000)$
$c = 1.71$ mol dm^{-3}

Exercises:

1. Calculate the concentration (in mol dm^{-3} and g dm^{-3}) of these solutions:

a. 10.6 g of sodium carbonate (Na_2CO_3) in 1.00 dm^3 of solution.

$n = m \div M$
$n = 10.6 \div 105.99 = 0.100$ mol Na_2CO_3
$c = n \div V$
$c = 0.100 \div (1.00)$
$c = 0.100$ mol dm^{-3}

$c = 10.6 \div (1.00)$
$c = 10.6$ g dm^{-3}

b. 117 g of sodium chloride (NaCl) in 5.00 dm^3 of solution.

$n = m \div M$
$n = 117 \div 58.44 = 2.00$ mol NaCl
c = n ÷ V
$c = 2.00 \div 5.00$
$c = 0.400$ mol dm^{-3}

$c = 117 \div 5.00$
$c = 23.4$ g dm^{-3}

c. 0.830 g of potassium iodide (KI) in 25.0 cm^3 of solution.

$n = m \div M$
$n = 0.830 \div 166.00 = 5.00 \times 10^{-3}$ mol KI
$c = n \div V$
$c = 5.00 \times 10^{-3} \div (25.0 \div 1000)$
$c = 0.200$ mol dm^{-3}

$c = 0.830 \div (25.0 \div 1000)$
$c = 33.2$ g dm^{-3}

2. Calculate the amount (in mol) of solute in each of the following solutions:

a. 0.250 dm^3 of 0.400 mol dm^{-3} ammonium chloride solution.

$n = cV$

$n = 0.400 \times 0.250$

$n = 0.100$ mol

b. 200.0 cm^3 of 0.800 mol dm^{-3} sodium carbonate solution.

$n = cV$

$n = 0.800 \times (200.0 \div 1000)$

$n = 0.160$ mol

c. 300.0 cm^3 of 4.00 mol dm^{-3} sodium hydroxide solution.

$n = cV$

$n = 4.00 \times (300.0 \div 1000)$

$n = 1.20$ mol

3. Calculate the mass of solute in the following solutions:

a. 2.00 dm^3 of 0.200 mol dm^{-3} potassium hydroxide (KOH) solution.

$n = cV$

$n = 0.200 \times 2.00$

$n = 0.400$ mol

$m = nM$

$m = 0.400 \times 56.11$

$m = 22.6$ g KOH

b. 200.0 cm^3 of 0.100 mol dm^{-3} sodium carbonate (Na_2CO_3) solution.

$n = cV$

$n = 0.100 \times (200.0 \div 1000)$

$n = 0.0200$ mol

$m = nM$

$m = 0.0200 \times 105.99$

$m = 2.12$ g Na_2CO_3

c. 25.0 cm^3 of 0.0500 mol dm^{-3} copper(II) sulphate ($CuSO_4 \bullet 5H_2O$) solution.

$n = cV$

$n = 0.0500 \times (25.0 \div 1000)$

$n = 1.25 \times 10^{-3}$ mol

$m = nM$

$m = 1.25 \times 10^{-3} \times 249.72$

$m = 0.312$ g $CuSO_4 \bullet 5H_2O$

Structure 1.4.6

Understandings:

- Avogadro's law states that equal volumes of all gases measured under the same conditions of temperature and pressure contain equal numbers of molecules.

Learning outcomes:

- Solve problems involving the mole ratio of reactants and/or products and the volume of gases.

Avogadro's law

- Equal volumes of gases at the same temperature and pressure contain the same number of particles.

- At STP (273 K and 100 kPa):

Amount (mol)	1 mol H_2	1 mol N_2	1 mol O_2
Volume (dm^3)	22.7	22.7	22.7
Number of particles	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Example: 40 cm^3 of CO reacts with 40 cm^3 of O_2. What volume of CO_2 is produced? What volume of the excess reactant remains?

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

Molar ratio of CO to O_2 is 2:1

40 cm^3 of CO reacts with 20 cm^3 of O_2 (molar ratio is 2:1)

Volume of O_2 is 40 cm^3, therefore O_2 is excess reactant.

Excess reactant remaining = 40 – 20 = 20 cm^3 of O_2

Exercises:

1. What volume of sulfur trioxide, in cm^3, can be prepared using 40 cm^3 sulfur dioxide and 20 cm^3 oxygen gas by the following reaction? Assume all volumes are measured at the same temperature and pressure.

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

Ratio of SO_2 to O_2 is 2:1

40 cm^3 of SO_2 would react with 20 cm^3 of O_2 to produce 40 cm^3 of SO_3

2. 5 dm^3 of carbon monoxide, $CO_{(g)}$, and 2 dm^3 of oxygen, $O_{2(g)}$, at the same temperature and pressure are mixed together. What is the maximum volume of carbon dioxide, $CO_{2(g)}$, in dm^3, that can be formed? What volume of the excess reactant remains?

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

Ratio of CO to O_2 is 2:1

5 dm^3 of CO would need 5 ÷ 2 = 2.5 dm^3 of O_2 to react completely.

Therefore, O_2 is the limiting reactant and CO is the excess reactant.

Ratio of O_2 to CO_2 is 1:2, therefore, 2 dm^3 of O_2 would produce 4 dm^3 of CO_2

2 dm^3 of O_2 reacts with 4 dm^3 of CO, 5-4 = 1 dm^3 of CO remains.

3. 100 cm^3 of ethene, C_2H_4, is burned in 400 cm^3 of oxygen, producing carbon dioxide and some liquid water. Some oxygen remains unreacted (excess).

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(g)}$$

Calculate the volume of carbon dioxide produced and the volume of oxygen remaining.

Molar ratio of reactants and products is 1:3:2:2

100 cm^3 of C_2H_4 reacts with 300 cm^3 of O_2 to produce 200 cm^3 of CO_2 and 200 cm^3 of H_2O

Volume of O_2 remaining: 400 – 300 = 100 cm^3