Mastering Chemistry

- Book 1C
- Topic 3 Metals

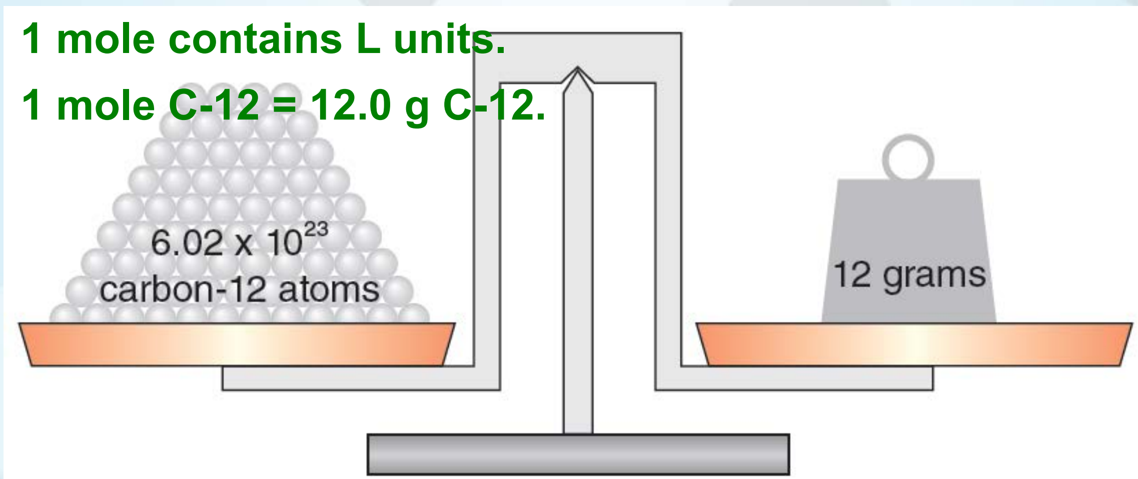
Content

Continued on next page

Content

12.1 What is a mole (p.71)

- Counting eggs: 1 dozen = 12 pieces
- Counting atoms: 1 **mole (摩爾)** = 6.02×10^{23} atoms
- 6.02×10^{23} mol^{-1} = **Avogadro constant (亞佛加德羅常數)** = L
  = the number of particles as there are atoms in 12.0 g of C-12



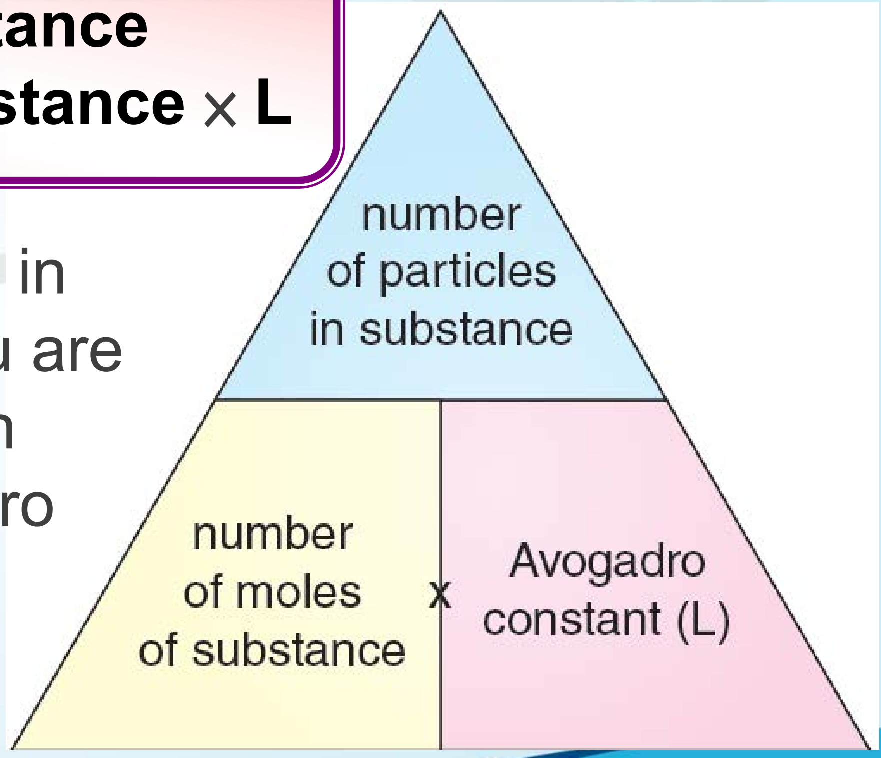
- 1 mole of chlorine <u>molecules</u> (Cl_2) has
  2 moles of chlorine <u>atoms</u> (Cl)

12.1 What is a mole (p.71)

- 0.10 mole of water molecules = 0.10 $\times$ 6.02 $\times$ 10^{23} molecules

number or particles in a substance
= number of moles of the substance $\times$ L

- If 'number of moles of substance' in the triangle below is covered, you are left with the number of particles in substance divided by the Avogadro constant:



$$\text{number of moles of substance} = \frac{\text{number of particles in substance}}{\text{Avogadro constant (L)}}$$

12.1 What is a mole (p.71)

Q (Example 12.1)

A sample contains 5.60 moles of sodium carbonate (Na_2CO_3). How many a) formula units; and b) ions are present?

A

a) Number of formula units = number of moles $\times$ L
$= 5.60 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} = 3.37 \times 10^{24}$

b) Number of ions = 3 $\times$ number of formula units
$= 3 \times 3.37 \times 10^{24} = 1.01 \times 10^{25}$

$\therefore$ the sample contains 3.37×10^{24} formula units and 1.01×10^{25} ions.

12.1 What is a mole (p.71)

Practice 12.1

1 Calculate the number of molecules in 0.160 mole of carbon dioxide.

Number of molecules
= number of moles x L
= 0.160 mol x 6.02 x 10^{23} mol^{-1}
= 9.63 x 10^{22}

2 Calculate the number of moles of nitric acid (HNO_3) that contains 9.03 x 10^{23} molecules.

Number of moles of HNO_3

$$= \frac{\text{number of molecules}}{L}$$

$$= \frac{9.03 \times 10^{23}}{6.02 \times 10^{23}\ mol^{-1}}$$

= 1.50 mol

12.1 What is a mole (p.71)

3 Calculate the total number of atoms in 3.50 moles of ammonia (NH_3).
(Avogadro constant = 6.02×10^{23} mol^{-1})

Number of molecules
= number of moles x L
= 3.50 mol x 6.02×10^{23} mol^{-1}
= 2.11×10^{24}

One NH_3 molecule contains 4 atoms.
Number of atoms
= $4 \times 2.11 \times 10^{24}$
= 8.44×10^{24}

12.2 Molar mass (p.74)

- The masses of atoms are measured relative to a scale on which carbon-12 has a mass of exactly 12.00.
- relative atomic mass of ^{4}He = 4.0 = ⅓ that of ^{12}C

i.e. 1 mole of ^{4}He atoms contains 4.0 g of ^{4}He;
1 mole of ^{12}C atoms contains 12.0 g of ^{12}C.

12.2 Molar mass (p.74)

The mass of one mole of a substance is its molar mass (摩爾質量). The unit used to measure molar mass is g mol^{-1}.

12.2 Molar mass (p.74)

- molar mass of a covalent compound (g mol^{-1}) has a value = its <u>relative molecular mass</u> (no unit)
  e.g. relative molecular mass of CO_2 = 44.0,
  i.e. molar mass of CO_2 = 44.0 g mol^{-1}

- molar mass of an ionic compound (g mol^{-1}) has a value = its <u>formula mass</u> (no unit)
  e.g. formula mass of KCl = 74.6,
  i.e. molar mass of KCl = 74.6 g mol^{-1}

12.2 Molar mass (p.74)

Practice 12.2

Calculate the molar masses of the following substances:

a) methane (CH_4);
(12.0 + 4 x 1.0) g mol^{-1} = 16.0 g mol^{-1}

b) potassium carbonate (K_2CO_3);
(2 x 39.1 + 12.0 + 3 x 16.0) g mol^{-1} = 138.2 g mol^{-1}

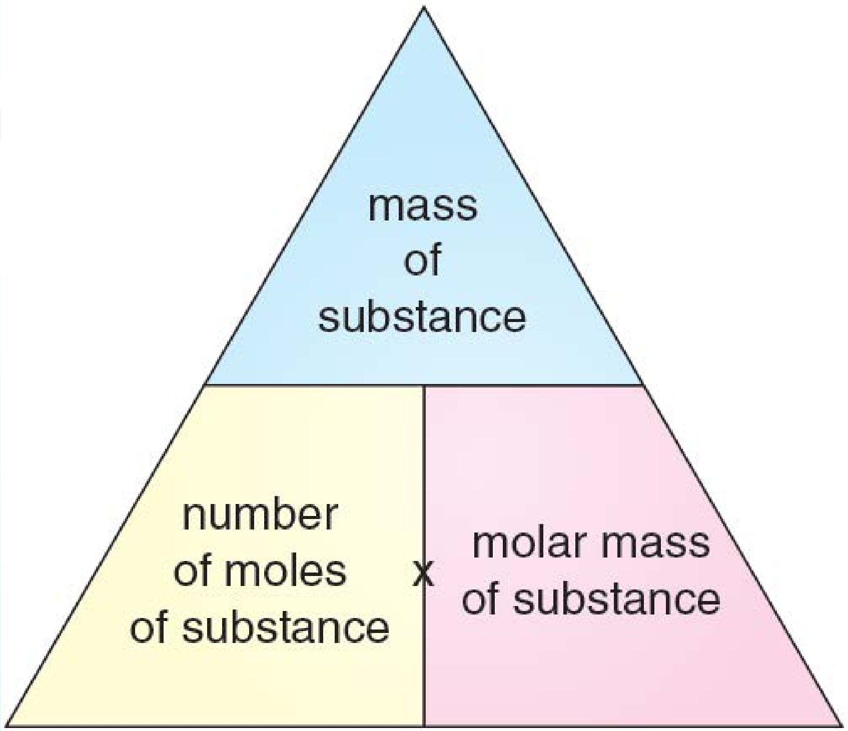
c) ethanol (C_2H_5OH);
(2 x 12.0 + 6 x 1.0 + 16.0) g mol^{-1} = 46.0 g mol^{-1}

d) magnesium phosphate ($Mg_3(PO_4)_2$).
(3 x 24.3 + 2 x 31.0 + 8 x 16.0) g mol^{-1} = 262.9 g mol^{-1}

(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0, Mg = 24.3, P = 31.0, K = 39.1)

12.3 Calculations involving moles and masses (p.76)

$$\text{number of moles of substance (mol)} = \frac{\text{mass of substance (g)}}{\text{molar mass of substance (g mol}^{-1}\text{)}}$$



12.3 Calculations involving moles and masses (p.76)

Q (Example 12.2)

How many moles of phosphoric acid molecules (H_3PO_4) are there in 84.3 g of the acid?
(Relative atomic masses: H = 1.0, O = 16.0, P = 31.0)

A

$$\text{Number of moles of } H_3PO_4 = \frac{\text{mass of } H_3PO_4}{\text{molar mass of } H_3PO_4}$$

$$= \frac{84.3 \text{ g}}{(3 \times 1.0 + 31.0 + 4 \times 16.0) \text{ g mol}^{-1}} = \frac{84.3 \text{ g}}{98.0\text{g mol}^{-1}} = 0.860 \text{ mol}$$

$\therefore$ there are 0.860 mole of phosphoric acid molecules.

12.3 Calculations involving moles and masses (p.76)

Q (Example 12.3)

What is the mass 0.0760 mole of aluminium sulphate?
(Relative atomic masses: O = 16.0, Al = 27.0, S = 32.1)

A

Mass of $Al_2(SO_4)_3$ = # of moles of $Al_2(SO_4)_3 \times$ molar mass of $Al_2(SO_4)_3$

$= 0.0760 \text{ mol} \times [2 \times 27.0 + 3 \times (32.1 + 4 \times 16.0)] \text{ g mol}^{-1}$

$= 0.0760 \text{ mol} \times 342.3 \text{ g mol}^{-1} = 26.0 \text{ g}$

$\therefore$ the mass of 0.0760 mole of aluminium sulphate is 26.0 g.

12.3 Calculations involving moles and masses (p.76)

Practice 12.3

1 A reaction uses 214 g of calcium nitrate.
How many moles of calcium nitrate have been used?
(Relative atomic masses: N = 14.0, O = 16.0, Ca = 40.1)

Molar mass of $Ca(NO_3)_2$
$= (40.1 + 2 \times 14.0 + 6 \times 16.0)\ g\ mol^{-1}$
$= 164.1\ g\ mol^{-1}$

Number of moles of $Ca(NO_3)_2$ used

$$= \frac{\text{mass of } Ca(NO_3)_2}{\text{molar mass of } Ca(NO_3)_2}$$

$$= \frac{214\ g}{164.1\ g\ mol^{-1}}$$

$= 1.30\ mol$

12.3 Calculations involving moles and masses (p.76)

2 Complete the table.
(Relative atomic masses: H = 1.0, C = 12.0, N = 14.0, O = 16.0, Na = 23.0, S = 32.1, K = 39.1, Cr = 52.0)

Compound	Molar mass (g mol^{-1})	Mass (g)	Number of moles (mol)
Ammonia (NH_3)	17.0	59.5	3.50
Glucose ($C_6H_{12}O_6$)	180.0	8.98	0.0499
Potassium dichromate ($K_2Cr_2O_7$)	294.2	648	2.20
Sodium sulphate (Na_2SO_4)	142.1	8.07×10^{-3}	5.68×10^{-5}

12.3 Calculations involving moles and masses (p.76)

Q (Example 12.4)

A beaker contains 82.8 g of water. How many
a) molecules; and b) atoms are present?
(Relative atomic masses: H = 1.0, O = 16.0, L = 6.02×10^{23} mol^{-1})

A

a) # of molecules = # of moles of $H_2O \times L$

$$= \frac{\text{mass of } H_2O}{\text{molar mass of } H_2O} \times L = \frac{82.8 \text{ g}}{(2 \times 1.0 + 16.0) \text{ g mol}^{-1}} \times L$$

$$= 4.60 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} = 2.77 \times 10^{24}$$

12.3 Calculations involving moles and masses (p.76)

b) Each H_2O molecule contains 3 atoms.

Number of atoms present $= 3 \times 2.77 \times 10^{24}$

$= 8.31 \times 10^{24}$

$\therefore$ there are 2.77×10^{24} molecules and 8.31×10^{24} atoms in 82.8 g of water.

12.3 Calculations involving moles and masses (p.76)

Practice 12.4

1 You are given a sample of 49.2 g of iron(III) sulphate.
How many a) formula units; b) ions are present?
(Relative atomic masses: O = 16.0, S = 32.1, Fe = 55.8; Avogadro constant = 6.02 x 10^{23} mol^{-1})

a) Molar mass of $Fe_2(SO_4)_3$
= (2 x 55.8 + 3 x 32.1 + 12 x 16.0) g mol^{-1}
= 399.9 g mol^{-1}

Number of moles of $Fe_2(SO_4)_3$

$$= \frac{\text{mass of } Fe(SO_4)_3}{\text{molar mass of } Fe_2(SO_4)_3}$$

$$= \frac{49.2 \text{ g}}{399.9 \text{ g mol}^{-1}}$$

= 0.123 mol

Number of formula units
= number of moles x L
= 0.123 mol x 6.02 x 10^{23} mol^{-1}
= 7.40 x 10^{22}

b) One formula unit of $Fe_2(SO_4)_3$ contains 5 ions.
Number of ions = 5 x 7.40 x 10^{22}
= 3.70 x 10^{23}

12.3 Calculations involving moles and masses (p.76)

2 Which of the following samples contains a greater number of atoms?
8.00 g of methane (CH_4) molecules or 6.02×10^{23} carbon dioxide molecules
(Relative atomic masses: H = 1.0, C = 12.0; Avogadro constant = 6.02×10^{23} mol^{-1})

Molar mass of CH_4 = $(12.0 + 4 \times 1.0)$ g mol^{-1}
= 16.0 g mol^{-1}

Number of moles of CH_4

$$= \frac{\text{mass of } CH_4}{\text{molar mass of } CH_4}$$

$$= \frac{8.00 \text{ g}}{16.0 \text{ g mol}^{-1}}$$

= 0.500 mol

Number of atoms in the sample of CH_4
= $0.500 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} \times 5$
= $2.50 \times 6.02 \times 10^{23}$

Number of atoms in the sample of CO_2
= $3 \times 6.02 \times 10^{23}$

∴ the sample of CO_2 contains a greater number of atoms.

12.4 Percentage by mass of an element in a compound (p.80)

Percentage by mass of an element in a compound

$$= \frac{\text{number of atoms of the element in formula} \times \text{relative atomic mass of the element}}{\text{formula mass or relative molecular mass of the compound}} \times 100\%$$

$$= \frac{(n_x)(M_x)}{M_{compound}} \times 100\%$$

12.4 Percentage by mass of an element in a compound (p.80)

Q (Example 12.5)

Calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$, is used in fertilisers for providing phosphorus to plants. What is the percentage by mass of phosphorus in the compound?
(Relative atomic masses: H = 1.0, O = 16.0, P = 31.0, Ca = 40.1)

A

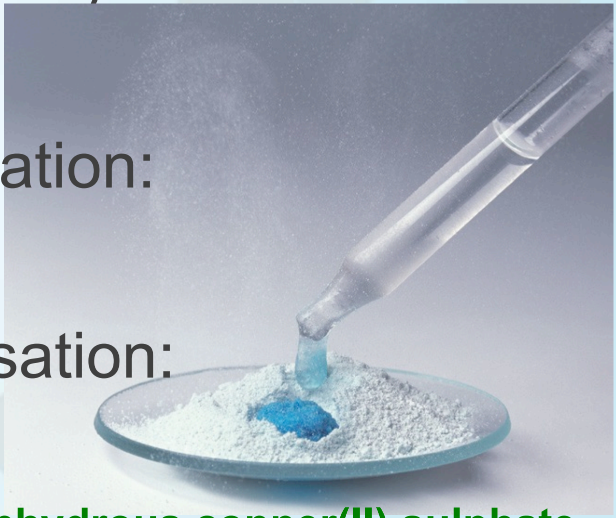
$$\text{\% by mass of P in } Ca(H_2PO_4)_2 = \frac{(n_P)(M_P)}{M_{Ca(H_2PO_4)_2}} \times 100\%$$

$$= \frac{2 \times 31.0}{[40.1 + 2 \times (2 \times 1.0 + 31.0 + 4 \times 16.0)]} \times 100\%$$

$$= \frac{2 \times 31.0}{234.1} \times 100\% = 26.5\%$$

12.4 Percentage by mass of an element in a compound (p.80)

- Some compounds crystallise with water as an integral part of the crystal lattice.
- The water is necessary for the formation of crystals and is called **water of crystallisation (結晶水)**.
- Compounds that contain water of crystallisation: **hydrated (水合)**, *e.g.* $CuSO_4 \cdot 5H_2O(s)$
- Those that have lost their water of crystallisation: **anhydrous (無水)**, *e.g.* $CuSO_4(s)$



When water is added, white anhydrous copper(II) sulphate becomes blue hydrated copper(II) sulphate

12.4 Percentage by mass of an element in a compound (p.80)

Q (Example 12.6)

Calculate the percentage by mass of water in hydrated barium chloride ($BaCl_2 \cdot 2H_2O$).
(Relative atomic masses: H = 1.0, O = 16.0, Cl = 35.5, Ba = 137.3)

A

% by mass of H_2O in $BaCl_2 \cdot 2H_2O$

$$= \frac{(n_{H_2O})(M_{H_2O})}{M_{BaCl_2 \cdot 2H_2O}} \times 100\%$$

$$= \frac{2 \times (2 \times 1.0 + 16.0)}{[137.3 + 2 \times 35.5 + 2 \times (2 \times 1.0 + 16.0)]} \times 100\%$$

$$= \frac{2 \times 18.0}{244.3} \times 100\% = 14.7\%$$

12.4 Percentage by mass of an element in a compound (p.80)

Practice 12.5

1 Urea (CH_4N_2O) is a common nitrogen-containing compound found in fertilisers. What is the percentage by mass of nitrogen in urea?
(Relative atomic masses: H = 1.0, C = 12.0, N = 14.0, O = 16.0)

2 What is the percentage by mass of water in hydrated zinc nitrate, $Zn(NO_3)_2 \cdot 6H_2O$?
(Relative atomic masses: H = 1.0, N = 14.0, O = 16.0, Zn = 65.4)

3 A metal carbonate (XCO_3) has 51.4% by mass of X and 9.72% by mass of carbon.
a) Calculate the relative atomic mass of X.
b) Hence identify X. (Refer to the Periodic Table)
(Relative atomic masses: C = 12.0, O = 16.0)

12.4 Percentage by mass of an element in a compound (p.80)

1 Relative molecular mass of urea
$= 12.0 + 4 \times 1.0 + 2 \times 14.0 + 16.0$
$= 60.0$

Percentage by mass of nitrogen in urea

$$= \frac{2 \times 14.0}{60.0} \times 100\%$$

$= 46.7\%$

∴ the percentage by mass of nitrogen in urea is 46.7%.

2 Formula mass of hydrated zinc nitrate
$= 65.4 + 2 \times 14.0 + 6 \times 16.0 + 6 \times (2 \times 1.0 + 16.0)$
$= 297.4$

Percentage by mass of water in hydrated zinc nitrate

$$= \frac{6 \times (2 \times 1.0 + 16.0)}{297.4} \times 100\%$$

$= 36.3\%$

∴ the percentage by mass of water in hydrated zinc nitrate is 36.3%.

12.4 Percentage by mass of an element in a compound (p.80)

3 a) Let m be the relative atomic mass of X.

Formula mass of XCO_3
$= m + 12.0 + 3 \times 16.0$
$= m + 60.0$

Percentage by mass of X in XCO_3

$$= 51.4\% = \frac{m}{m + 60.0} \times 100\%$$

$$51.4 \times (m + 60.0) = 100m$$
$$51.4m + 3\,084 = 100m$$
$$m = 63.5$$

∴ the relative atomic mass of X is 63.5.

b) Copper

12.5 Empirical formula (p.82)

- **Empirical formula (實驗式)**—gives the simplest whole number ratio of atoms or ions in the compound

Working out the empirical formula from % composition

- Assume that you have 100.0 g of the compound. Obtain the mass of each element in such a sample.
- Calculate the number of moles of atoms of each element.
- Divide each element's molar amount by the smallest molar amount.
- If the results of the calculations do not approximate to whole numbers, multiply them all by 2 to obtain whole numbers. (In some cases, you might have to multiply by 3 or 4 to obtain whole numbers.)

12.5 Empirical formula (p.82)

Q (Example 12.8)

An oxide of chromium contains 31.6% of oxygen by mass.

What is the empirical formula of the oxide?

(Relative atomic masses: O = 16.0, Cr = 52.0)

12.5 Empirical formula (p.82)

A 100.0 g of the oxide contain 68.4 g of Cr and 31.6 g of O.

	Chromium	Oxygen
Mass:	68.4 g	31.6 g
Mole:	$\frac{68.4 \text{ g}}{52.0 \text{ g mol}^{-1}} = 1.32 \text{ mol}$	$\frac{31.6 \text{ g}}{16.0 \text{ g mol}^{-1}} = 1.98 \text{ mol}$
Mole ratio:	$\frac{1.32}{1.32} = 1.00$	$\frac{1.98}{1.32} = 1.50$
Whole # mole ratio:	1.00 x 2 = 2	1.50 x 2 = 3

$\therefore Cr_2O_3$

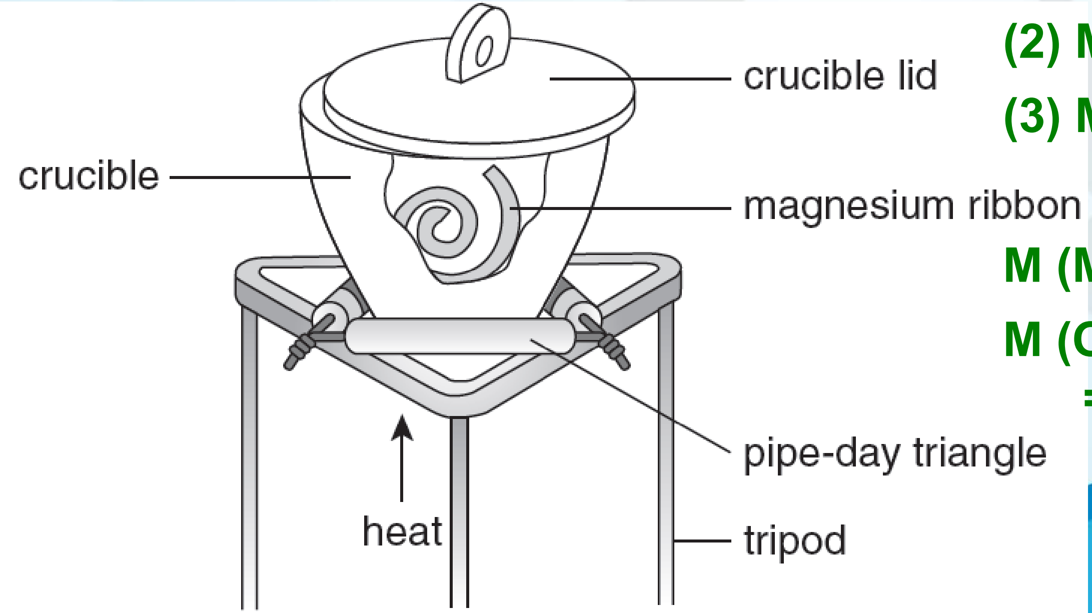
12.5 Empirical formula (p.82)

Working out the empirical formula from experimental data

Magnesium oxide

magnesium + oxygen ⟶ magnesium oxide

- Convert known mass of Mg to Mg_xO_y.
- Measure the mass of Mg_xO_y to get the mass of O combined.
- From mass of Mg : mass of O, get mole of Mg : mole of O to obtain the empirical formula Mg_xO_y.



(1) M (crucible) = 21.884 g

(2) M (crucible + Mg) = 22.202 g

(3) M (crucible + Mg_xO_y) = 22.394 g

M (Mg) = (2) – (1) = 0.318 g

M (O combined) = (3) – (2)
= 0.192 g

12.5 Empirical formula (p.82)

Magnesium	Oxygen
0.318 g	0.192 g
$\frac{0.318\ \text{g}}{24.3\ \text{g mol}^{-1}} = 0.0131\ \text{mol}$	$\frac{0.192\ \text{g}}{16.0\ \text{g mol}^{-1}} = 0.0120\ \text{mol}$
$\frac{0.0131}{0.0120} = 1.09$	$\frac{0.0120}{0.0120} = 1.00$

$\therefore$ the empirical formula of magnesium oxide is MgO.

Determining the empirical formula of magnesium oxide

12.5 Empirical formula (p.82)

Working out the empirical formula from experimental data

Iron oxide

- Reduce known mass of Fe_xO_y to Fe and CO_2.

$$\text{oxide of iron} + \text{carbon monoxide} \longrightarrow \text{iron} + \text{carbon dioxide}$$

- Measure the mass of Fe produced to get the orginal mass of O.
- From mass of Fe : mass of O, get mole of Fe : mole of O to obtain the empirical formula Fe_xO_y.

(1) M (iron oxide) = 2.315 g

(2) M (Fe produced) = 1.675 g

(3) M (O in Fe_xO_y) = (1) – (2) = 0.640 g

12.5 Empirical formula (p.82)

Iron	Oxygen
1.675 g	0.640 g
$\dfrac{1.675\ \text{g}}{55.8\ \text{g mol}^{-1}} = 0.0300\ \text{mol}$	$\dfrac{0.640\ \text{g}}{16.0\ \text{g mol}^{-1}} = 0.0400\ \text{mol}$
$\dfrac{0.0300}{0.0300} = 1.00$	$\dfrac{0.0400}{0.0300} = 1.33$
1.00 x 3 = 3	1.33 x 3 = 4

∴ the empirical formula of the iron oxide is Fe_3O_4.

12.5 Empirical formula (p.82)

Working out the empirical formula from experimental data

A hydrated compound

Q (Example 12.9)

12.016 g of hydrated magnesium sulphate, $MgSO_4{\bullet}nH_2O$, gives 5.915 g o anhydrous magnesium sulphate on heating. Deduce the value of *n*.
(Relative atomic masses: H = 1.0, O = 16.0, Mg = 24.3, S = 32.1)

12.5 Empirical formula (p.82)

A

Mass of water driven off = (12.106 – 5.915) g = 6.191 g

	$MgSO_4$	H_2O
Mass:	5.915 g	6.191 g
Mole:	$\frac{5.915\text{ g}}{120.4\text{ g mol}^{-1}} = 0.0491\text{ mol}$	$\frac{6.191\text{ g}}{18.0\text{ g mol}^{-1}} = 0.344\text{ mol}$
Mole ratio:	$\frac{0.0491}{0.0491} = 1.00$	$\frac{0.344}{0.0491} = 7.01$

∴ the value of *n* is 7.

12.5 Empirical formula (p.82)

Practice 12.6

1 A fluoride of sulphur contains 29.7% by mass of sulphur. What is the empirical formula of the fluoride?
(Relative atomic masses: F = 19.0, S = 32.1)

1 100.0 g of the fluoride contain 29.7 g of sulphur and 70.3 g of fluoride.

	Sulphur	Fluorine
1 Mass of element	29.7 g	70.3 g
2 Number of moles of atoms	$\frac{29.7\text{ g}}{32.1\text{ g mol}^{-1}} = 0.925\text{ mol}$	$\frac{70.3\text{ g}}{19.0\text{ g mol}^{-1}} = 3.70\text{ mol}$
3 Mole ratio of atoms	$\frac{0.925}{0.925} = 1.00$	$\frac{3.70}{0.925} = 4.00$

$\therefore$ the empirical formula of the fluoride is SF_4.

12.5 Empirical formula (p.82)

2 A 1.000 g sample of red phosphorus was burnt in air and reacted with oxygen to give 2.291 g of an oxide. What is the empirical formula of the oxide? (Relative atomic masses: O = 16.0, P = 31.0)

2 Mass of oxygen in the oxide = (2.291 – 1.000) g
= 1.291 g

	Phosphorus	Oxygen
1 Mass of element	1.000 g	1.291 g
2 Number of moles of atoms	$\frac{1.000\ \text{g}}{31.0\ \text{g mol}^{-1}} = 0.0323\ \text{mol}$	$\frac{1.291\ \text{g}}{16.0\ \text{g mol}^{-1}} = 0.0807\ \text{mol}$
3 Mole ratio of atoms	$\frac{0.0323}{0.0323} = 1.00$	$\frac{0.0807}{0.0323} = 2.50$
4 Whole number mole ratio of atoms	1.00 x 2 = 2	2.50 x 2 = 5

∴ the empirical formula of the oxide is P_2O_5.

12.5 Empirical formula (p.82)

3 A sample of hydrated sodium carbonate ($Na_2CO_3 \cdot nH_2O$) has a mass of 4.317 g before heating. After heating, the mass of the anhydrous sodium carbonate is found to be 3.224 g. Deduce the value of *n*.
(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0, Na = 23.0)

3 Mass of water driven off = (4.317 – 3.224) g
= 1.093 g

	Na_2CO_3	H_2O
1 Mass of compound	3.224 g	1.093 g
2 Number of moles of compound	$\frac{3.224\ g}{106.0\ g\ mol^{-1}} = 0.0304\ mol$	$\frac{1.093\ g}{18.0\ g\ mol^{-1}} = 0.0607\ mol$
3 Mole ratio of compound	$\frac{0.0304}{0.0304} = 1.00$	$\frac{0.0607}{0.0304} = 2.00$

∴ the value of *n* is 2.

12.6 Molecular formula (p.88)

- The empirical formula of a compound may or may not be the same as its molecular formula.

Table 12.2 Examples of compounds with different empirical and molecular formulae

Compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$

- To determine the molecular formula of a compound, you need
  - the empirical formula; and
  - the relative molecular mass.

12.6 Molecular formula (p.88)

Q (Example 12.10)

Compound X contains 38.7% carbon, 9.7% hydrogen and 51.6% oxygen. Its relative molecular mass is 62.0.

Find its

a) empirical formula; and

b) Molecular formula.

(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

12.6 Molecular formula (p.88)

A

a) 100.0 g of compound X contain 38.7 g of carbon, 9.7 g of hydrogen and 51.6 g of oxygen.

	Carbon	Hydrogen	Oxygen
Mass:	38.7 g	9.7 g	51.6 g
Mole:	$\frac{38.7\text{ g}}{12.0\text{ g mol}^{-1}} = 3.23\text{ mol}$	$\frac{9.7\text{ g}}{1.0\text{ g mol}^{-1}} = 9.7\text{ mol}$	$\frac{51.6\text{ g}}{16.0\text{ g mol}^{-1}} = 3.23\text{ mol}$
Mole ratio:	$\frac{3.23}{3.23} = 1.00$	$\frac{9.7}{3.23} = 3.00$	$\frac{3.23}{3.23} = 1.00$

∴ the empirical formula of compound X is CH_3O.

12.6 Molecular formula (p.88)

b) Let $(CH_3O)_n$ be the molecular formula of compound X.

Relative molecular mass of X = $n(12.0 + 3 \times 1.0 + 16.0)$
$= 31n$
$\therefore 31n = 62.0$
$n = 2$

$\therefore$ the molecular formula of compound X is $(CH_3O)_2$ or $C_2H_6O_2$.

12.6 Molecular formula (p.88)

Practice 12.7

1 Write the empirical formula for each of the following compounds.
   a) hydrazine (N_2H_4) NH_2
   b) butane (C_4H_{10}) C_2H_5

2 Vitamic C has an empirical formula of $C_3H_4O_3$ and a relative molecular mass of 176.0. Deduce the molecular formula of vitamin C.
(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

Let $(C_3H_4O_3)_n$ be the molecular formula of vitamin C.
Relative molecular mass of vitamin C = $n(3 \times 12.0 + 4 \times 1.0 + 3 \times 16.0) = 88n$
$\therefore 88n = 176.0$
$n = 2$
$\therefore$ the molecular formula of vitamin C is $C_6H_8O_6$.

12.7 Amounts and equations (p.90)

- Chemists need to know exactly how much of each reactant they need to use in a reaction to make a certain amount of product.
- $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ tells you that

1 hydrogen molecule	reacts with	1 chlorine molecule	to make	2 hydrogen chloride molecules
1 mole of hydrogen molecules	reacts with	1 mole of chlorine molecules	to make	2 moles of hydrogen chloride molecules

12.7 Amounts and equations (p.90)

- Based on relative molecular masses:

2.0 g of hydrogen	reacts with	71.0 g of chlorine	to make	73.0 g of hydrogen chloride

- Based on reaction in the same ratio:

1.0 g of hydrogen	reacts with	35.5 g of chlorine	to make	36.5 g of hydrogen chloride

12.8 Calculating masses from equations (p.91)

- Chemists want to know in a reaction how much is produced or required.

- Method 1—using number of moles of substances involved

Given mass of X → how many moles of X → **Use the coefficients → how many moles of Y** → **From moles of Y → how many grams of Y**

- Method 2—using ratios

Calculate molar masses of X and Y → **Use the coefficients → mass ratio X : Y** → **Use mass ratio X : Y and given mass of X → how many grams of Y**

Thermally decomposing sodium hydrogencarbonate

Determining the percentage by mass of certain components

12.8 Calculating masses from equations (p.91)

Q (Example 12.11)

Iron(III) oxides reacts with carbon monoxide according to:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

638 g of iron(III) oxide are allowed to react with excess carbon monoxide. Calculate the mass of iron made.

(Relative atomic masses: O = 16.0, Fe = 55.8)

12.8 Calculating masses from equations (p.91)

A

Given mass of X → how many moles of X	**Use the coefficients → how many moles of Y**	**From moles of Y → how many grams of Y**

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

Method 1

1) # of moles of Fe_2O_3 = $\dfrac{\text{mass of } Fe_2O_3}{\text{molar mass of } Fe_2O_3} = \dfrac{638\text{ g}}{(2 \times 55.8 + 3 \times 16.0)\text{ g mol}^{-1}}$

$= \dfrac{638\text{ g}}{159.6\text{ g mol}^{-1}} = 4.00\text{ mol}$

2) # of moles of Fe made = 2×4.00 mol = 8.00 mol

3) Mass of Fe made = 8.00 mol $\times$ 55.8 g mol^{-1} = 446 g

12.8 Calculating masses from equations (p.91)

A

Calculate molar masses of X and Y → Use the coefficients → mass ratio X : Y → Use mass ratio X : Y and given mass of X → how many grams of Y

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

Method 2

1) Molar mass of Fe_2O_3 = $(2 \times 55.8 + 3 \times 16.0)$ g mol^{-1} = 159.6 g mol^{-1}

Molar mass of Fe = 55.8 g mol^{-1}

2) 159.6 g of Fe_2O_3 reacts to make 2×55.8 g of Fe.

3) Mass of Fe made $= 638 \text{ g} \times \dfrac{2 \times 55.8 \text{ g}}{159.6 \text{ g}} = 446 \text{ g}$

∴ 446 g of iron are made.

12.8 Calculating masses from equations (p.91)

Practice 12.8

1 Sulphur dioxide reacts with oxygen to make sulphur trioxide.

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

Calculate the mass of sulphur trioxide that can be made by the reaction of 45.3 g of sulphur dioxide and excess oxygen.
(Relative atomic masses: O = 16.0, S = 32.0)

12.8 Calculating masses from equations (p.91)

1 $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$

<u>*Method 1*</u>

1 Molar mass of SO_2 = (32.1 + 2 x 16.0) g mol^{-1} = 64.1 g mol^{-1}

Number of moles of $SO_2 = \dfrac{\text{mass of } SO_2}{\text{molar mass of } SO_2}$

$= \dfrac{45.3 \text{ g}}{64.1 \text{ g mol}^{-1}}$

= 0.707 mol

2 According to the equation, 2 moles of SO_2 react with O_2 to make 2 moles of SO_3.

i.e. number of moles of SO_3 made = 0.707 mol

3 Molar mass of SO_3 = (32.1 + 3 x 16.0) g mol^{-1} = 80.1 g mol^{-1}

Mass of SO_3 made = number of moles of SO_3 x molar mass of SO_3

= 0.707 mol x 80.1 g mol^{-1}

= 56.6 g

12.8 Calculating masses from equations (p.91)

Method 2

1 Molar mass of SO_2 = (32.1 + 2 x 16.0) g mol^{-1} = 64.1 g mol^{-1}

Molar mass of SO_3 = (32.1 + 3 x 16.0) g mol^{-1} = 80.1 g mol^{-1}

2 According to the equation, 2 moles of SO_2 react with O_2 to make 2 moles of SO_3.

i.e. 2 x 64.1 g of SO_2 react with O_2 to make 2 x 80.1 g of SO_3.

3 Mass of SO_3 made $= 45.3 \text{ g} \times \dfrac{2 \times 80.1 \text{ g}}{2 \times 64.1 \text{ g}}$

$= 56.6$ g

∴ 56.6 g of SO_3 are made.

12.8 Calculating masses from equations (p.91)

2 A sample contains 78.0% of lead(II) oxide by mass.
Assuming that the other components in this sample do not contain lead, what mass of the sample is required to extract 15.0 tonnes of lead?
(Relative atomic masses: O = 16.0, Pb = 207.2; 1 tonne = 106 g)

12.8 Calculating masses from equations (p.91)

2 $PbO(s) \longrightarrow Pb(s)$

15.0 tonnes

Method 1

1 Number of moles of Pb $= \dfrac{\text{mass of Pb}}{\text{molar mass of Pb}}$

$= \dfrac{15.0 \times 10^6 \text{ g}}{207.2 \text{ g mol}^{-1}}$

$= 72\ 400$ mol

2 According to the equation, 1 mole of PbO gives 1 mole of Pb in extraction.

i.e. number of moles of PbO = 72 400 mol

3 Molar mass of PbO $= (207.2 + 16.0) \text{ g mol}^{-1} = 223.2 \text{ g mol}^{-1}$

Mass of PbO = number of moles of PbO x molar mass of PbO

$= 72\ 400 \text{ mol} \times 223.2 \text{ g mol}^{-1}$

$= 1.62 \times 10^7$ g

= 16.2 tonnes

Mass of sample required $= \dfrac{16.2}{78.0\%}$ tonnes

= 20.8 tonnes

12.8 Calculating masses from equations (p.91)

Method 2

1 Molar mass of PbO = (207.2 + 16.0) g mol^{-1} = 223.2 g mol^{-1}

2 According to the equation, 1 mole of PbO gives 1 mole of Pb in extraction.

i.e. 223.2 g of PbO give 207.2 g of Pb in extraction.

3 Mass of PbO = 15.0 tonnes x $\frac{223.2 \text{ g}}{207.2 \text{ g}}$

= 16.2 tonnes

Mass of sample required = $\frac{16.2}{78.0\%}$ tonnes

= 20.8 tonnes

∴ 20.8 tonnes of sample are required.

12.8 Calculating masses from equations (p.91)

3 Strontium is extracted from strontium oxide (SrO) by heating a mixture of strontium oxide and aluminium.

$$3SrO(s) + 2Al(s) \rightarrow 3Sr(s) + Al_2O_3(s)$$

Calculate the mass of aluminium required to react completely with 55.3 g of strontium oxide.

(Relative atomic masses: O = 16.0, Al = 27.0, Sr = 87.6)

12.8 Calculating masses from equations (p.91)

3 $3SrO(s) + 2Al(s) \longrightarrow 3Sr(s) + Al_2O_3(s)$

Method 1

1 Molar mass of SrO = (87.6 + 16.0) g mol^{-1} = 103.6 g mol^{-1}

$$\text{Number of moles of SrO} = \frac{\text{mass of SrO}}{\text{molar mass of SrO}} = \frac{55.3\ \text{g}}{103.6\ \text{g mol}^{-1}} = 0.534\ \text{mol}$$

2 According to the equation, 3 moles of SrO react with 2 moles of Al.

$$\text{i.e. number of moles of Al} = \frac{2}{3} \times 0.534\ \text{mol} = 0.356\ \text{mol}$$

3 Mass of Al = number of moles of Al x molar mass of Al

$$= 0.356\ \text{mol} \times 27.0\ \text{g mol}^{-1} = 9.61\ \text{g}$$

12.8 Calculating masses from equations (p.91)

Method 2

1 Molar mass of SrO = (87.6 + 16.0) g mol^{-1} = 103.6 g mol^{-1}

Molar mass of Al = 27.0 g mol^{-1}

2 According to the equation, 3 moles of SrO react with 2 moles of Al.

i.e. 3 x 103.6 g of SrO react with 2 x 27.0 g of Al.

3 Mass of Al $= 55.3 \text{ g} \times \dfrac{2 \times 27.0 \text{ g}}{3 \times 103.6 \text{ g}}$

$= 9.61 \text{ g}$

∴ 9.61 g of aluminium are required.

12.9 Limiting reactant (p.98)

- A chemical reaction stops when one of the reactants is used up:
  - the one used up is the **limiting reactant (限量反應物)**;
  - the one not used up is *in excess*.

Q (Example 12.15)

Consider the reaction: $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

16.0 g of magnesium are allowed to react with 8.00 g of nitrogen.
a) Identify the limiting reactant.
b) Calculate the mass of magnesium nitride that can be obtained.
(Relative atomic masses: N = 14.0, Mg = 24.3)

12.9 Limiting reactant (p.98)

A

a) $\# \text{ of moles of Mg} = \dfrac{\text{mass of Mg}}{\text{molar mass of Mg}} = \dfrac{16.0 \text{ g}}{24.3 \text{ g mol}^{-1}} = 0.658 \text{ mol}$

$\# \text{ of moles of } N_2 = \dfrac{\text{mass of } N_2}{\text{molar mass of } N_2} = \dfrac{8.00 \text{ g}}{2 \times 14.0 \text{ g mol}^{-1}} = 0.286 \text{ mol}$

According to the equation, 3 moles of Mg react with 1 mole of N_2.
In this reaction, 0.658 mole of Mg reacts with 0.219 mole of N_2.
Thus, N_2 is in excess. Mg is the limiting reactant.

12.9 Limiting reactant (p.98)

A

b) # of moles of Mg_3N_2 made $= \frac{1}{3} \times$ number of moles of Mg $= 0.219$ mol

Molar mass of Mg_3N_2 = $(3 \times 24.3 + 2 \times 14.0)$ g mol^{-1} = 100.9 g mol^{-1}

Mass of Mg_3N_2 made = number of moles of Mg_3N_2 $\times$ molar mass of Mg_3N_2

= 0.219 mol $\times$ 100.9 g mol^{-1} = 22.1 g

12.9 Limiting reactant (p.98)

Practice 12.9

1 At high temperatures, iron combines with sulphur to make iron(II) sulphide.

$$Fe(s) + S(s) \rightarrow FeS(s)$$

In an experiment, 7.48 g of iron are allowed to react with 8.67 g of sulphur.

a) Identify the limiting reactant.

b) Calculate the mass of iron(II) sulphide made.

(Relative atomic masses: S = 32.1, Fe = 55.8)

1 a) Number of moles of Fe $= \dfrac{\text{mass of Fe}}{\text{molar mass of Fe}}$

$= \dfrac{7.48 \text{ g}}{55.8 \text{ g mol}^{-1}}$

$= 0.134 \text{ mol}$

Number of moles of S $= \dfrac{\text{mass of S}}{\text{molar mass of S}}$

$= \dfrac{8.67 \text{ g}}{32.1 \text{ g mol}^{-1}}$

$= 0.270 \text{ mol}$

According to the equation, 1 mole of Fe reacts with 1 mole of S.

In this reaction, 0.134 mole of Fe reacts with 0.134 mole of S.

Thus, S is in excess. Fe is the limiting reactant.

b) The amount of FeS made is determined by the amount of Fe.

Number of moles of FeS made = number of moles of Fe
$= 0.134 \text{ mol}$

Molar mass of FeS $= (55.8 + 32.1) \text{ g mol}^{-1}$
$= 87.9 \text{ g mol}^{-1}$

Mass of FeS made = number of moles of FeS x molar mass of FeS
$= 0.134 \text{ mol} \times 87.9 \text{ g mol}^{-1}$
$= 11.8 \text{ g}$

12.9 Limiting reactant (p.98)

2 Refer to the following chemical equation:

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

N moles of Al are allowed to react with *N* moles of O_2 under suitable conditions until the reaction stops. How many moles of Al_2O_3 are formed?

According to the equation, 4 moles of Al react with 3 moles of O_2 to give 2 moles of Al_2O_3.

In this reaction, *N* moles of Al react with $\frac{3}{4}N$ moles of O_2 to give $\frac{1}{2}N$ moles of Al_2O_3.

Key terms (p.100)

mole	摩爾	hydrated	水合的
Avogadro constant	亞佛加德羅常數	anhydrous	無水的
molar mass	摩爾質量	empirical formula	實驗式
water of crystallisation	結晶水	limiting reactant	限量反應物

Summary (p.101)

1 The value of 6.02×10^{23} mol^{-1} is called Avogadro constant (symbol L).

2 Number of particles in a substance
= number of moles of the substance x L

3 One mole (symbol: mol) is the amount of a substance that contains the same number of particles as there are atoms in 12.0 g of carbon-12.

4 The mass of one mole of a substance is its molar mass. The units of molar mass are g mol^{-1}.

Summary (p.101)

5 $$\text{Number of moles of substance (mol)} = \frac{\text{mass of substance (g)}}{\text{molar mass of substance (g mol}^{-1}\text{)}}$$

6 Percentage by mass of an element in a compound

$$= \frac{\text{number of atoms of the element in formula} \times \text{relative atomic mass of the element}}{\text{formula mass or relative molecular mass of the compound}} \times 100\%$$

7 The empirical formula of a compound gives the simplest whole number ratio of atoms or ions in the compound.

8 Choose one of the methods to calculate the masses of substances involved in a reaction:
a) using numbers of moles of substances involved; or
b) using ratios.

Unit Exercise (p.102)

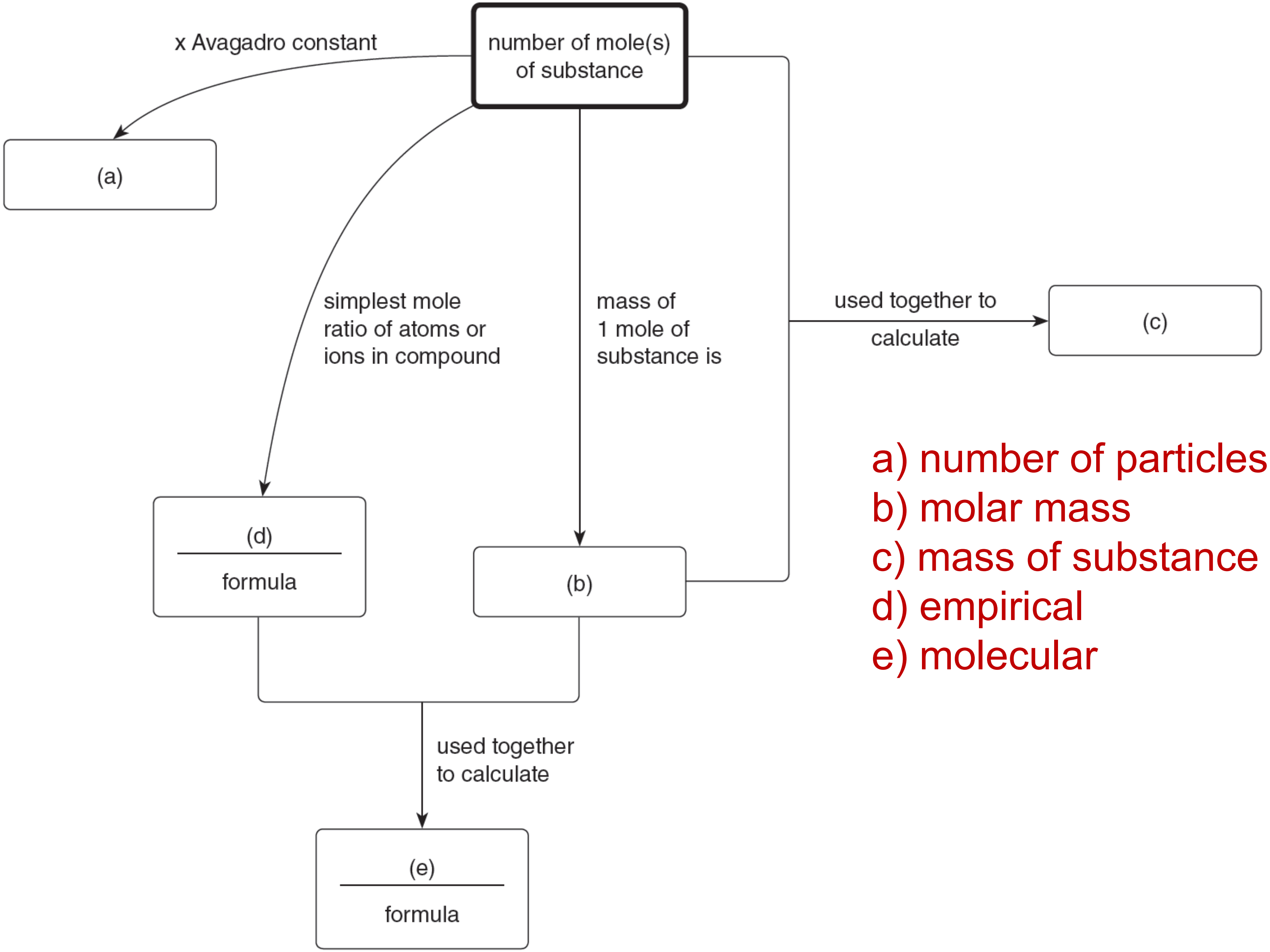
Note: Questions are rated according to ascending level of difficulty (from 1 to 5):

question targeted at level 3 and above;
question targeted at level 4 and above;
question targeted at level 5.
' * ' indicates 1 mark is given for effective communication.

Unit Exercise (p.102)

PART I KNOWLEDGE AND UNDERSTANDING

1 Complete the following concept map.



a) number of particles
b) molar mass
c) mass of substance
d) empirical
e) molecular

Unit Exercise (p.102)

PART II MULTIPLE CHOICE QUESTIONS

2 What is the number of ATOMS in 1 mole of methane (CH_4)? (Avogadro constant = 6.02×10^{23} mol^{-1})

A 3.01×10^{24}
B 1.50×10^{24}
C 6.02×10^{23}
D 1.20×10^{23}

Answer: A

Explanation: One methane molecule contains 5 atoms.

Unit Exercise (p.102)

3 What is the total number of ions in 25.7 g of barium hydroxide, $Ba(OH)_2$?
(Relative atomic masses: H = 1.0, O = 16.0, Ba = 137.3;
(Avogadro constant = 6.02×10^{23} mol^{-1})

A 9.03×10^{22}
B 1.20×10^{23}
C 2.71×10^{23}
D 5.87×10^{23}

Answer: C

(HKDSE, Paper 1A, 2016, 4)

Unit Exercise (p.102)

Explanation:

$$\text{Molar mass of } Ba(OH)_2 = (137.3 + 2 \times 16.0 + 2 \times 1.0)\ g\ mol^{-1}$$
$$= 171.3\ g\ mol^{-1}$$

$$\text{Number of moles of } Ba(OH)_2 = \frac{\text{mass of } Ba(OH)_2}{\text{molar mass of } Ba(OH)_2}$$
$$= \frac{25.7\ g}{171.3\ g\ mol^{-1}}$$
$$= 0.150\ mol$$

$$\text{Number of formula units} = 0.150\ mol \times 6.02 \times 10^{23}\ mol^{-1}$$
$$= 9.03 \times 10^{22}$$

One formula unit of $Ba(OH)_2$ contains 3 ions.

$$\text{Number of ions} = 3 \times 9.03 \times 10^{22}$$
$$= 2.71 \times 10^{23}$$

Unit Exercise (p.102)

4 Nickel makes up 18.0% of the total mass of a coin.
The coin has a mass of 9.440 g.
How many nickel atoms are in the coin?
(Relative atomic mass: Ni = 58.7; Avogadro constant = $6.02 \times 10^{23}\ mol^{-1}$)

A 1.74×10^{22}
B 1.96×10^{22}
C 1.02×10^{23}
D 1.85×10^{24}

Answer: A

Unit Exercise (p.102)

Explanation:

Mass of nickel = 9.440 g x 18.0%
= 1.699 g

$$\text{Number of moles of nickel} = \frac{\text{mass of nickel}}{\text{molar mass of nickel}}$$

$$= \frac{1.699 \text{ g}}{58.7 \text{ g mol}^{-1}}$$

$$= 0.0289 \text{ mol}$$

Number of nickel atoms = 0.0289 mol x 6.02 x 10^{23} mol^{-1}
= 1.74 x 10^{22}

Unit Exercise (p.102)

5 Which mass of substance below contains the greatest number of atoms?
(Relative atomic masses: H = 1.0, C = 12.0, N = 14.0, S = 32.1, Cl = 35.5)

A 3.00 g of ammonia, NH_3
B 3.00 g of chloromethane, $CHCl_3$
C 4.00 g of hydrogen sulphide, H_2S
D 4.00 g of hydrogen chloride, HCl

Answer: A

(OCR Advanced Subsidiary GCE, Chem. A, H032/01, Sample Question Paper, 2014, 9)

Unit Exercise (p.102)

Explanation:

Substance	Molar mass (g mol^{-1})	Number of moles of molecules (mol)	Number of moles of atoms (mol)
3.00 g of NH_3	14.0 + 3 x 1.0 = 17.0	$\frac{3.00}{17.0} = 0.176$	4 x 0.176 = 0.704
3.00 g of $CHCl_3$	12.0 + 1.0 + 3 x 35.5 = 119.5	$\frac{3.00}{119.5} = 0.0251$	5 x 0.0251 = 0.126
4.00 g of H_2S	2 x 1.0 + 32.1 = 34.1	$\frac{4.00}{34.1} = 0.117$	3 x 0.117 = 0.351
4.00 g of HCl	1.0 + 35.5 = 36.5	$\frac{4.00}{36.5} = 0.110$	2 x 0.110 = 0.220

∴ 3.00 g of ammonia contain the greatest number of atoms.

Unit Exercise (p.102)

6 If 8.0 g of sulphur dioxide gas contains *n* molecules, how many molecules does 2.0 g of oxygen gas contain? (Relative atomic masses: O = 16.0, S = 32.1)

A 2.0 *n*
B 4.0 *n*
C 0.25 *n*
D 0.50 *n*

Answer: D

(HKDSE, Paper 1A, 2018, 4)

Unit Exercise (p.102)

7 Which of the following potassium compounds contains 52.4% by mass of potassium?
(Relative atomic masses: C = 12.0, O = 16.0, S = 32.1, Cl = 35.5, K = 39.1)

A Potassium carbonate
B Poatssium chloride
C Potassium sulphate
D Potassium oxide

Answer: B

Explanation:

Formula mass of KCl
$= 39.1 + 35.5 = 74.6$

Percentage by mass of potassium in KCl

$$= \frac{39.1}{74.6} \times 100\%$$

$= 52.4\%$

Unit Exercise (p.102)

8 A compound is found to contain 5.89 g of nitrogen and 10.1 g of oxygen. What is the empirical formula of the compound?
(Relative atomic masses: N = 14.0, O = 16.0)

A NO
B NO_2
C N_2O_3
D N_2O_5

Answer: C

Unit Exercise (p.102)

Explanation:

	Nitrogen	Oxygen
1 Mass of element	5.89 g	10.1 g
2 Number of moles of atoms	$\frac{5.89\text{ g}}{14.0\text{ g mol}^{-1}} = 0.421\text{ mol}$	$\frac{10.1\text{ g}}{16.0\text{ g mol}^{-1}} = 0.631\text{ mol}$
3 Mole ratio of atoms	$\frac{0.421}{0.421} = 1.00$	$\frac{0.631}{0.421} = 1.50$
4 Whole number mole ratio of atoms	1.00 x 2 = 2	1.50 x 2 = 3

$\therefore$ the empirical formula of the compound is N_2O_3.

Unit Exercise (p.102)

9 Camphor is used as an ingredient in cooking in India. It contains the same percentage by mass of hydrogen and oxygen.
What is the molecular formula of camphor?
(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

A $C_{10}H_8O_8$
B $C_{10}H_{10}O_{10}$
C $C_{10}H_{16}O$
D $C_{10}H_{12}O_2$

Answer: C

Explanation:

Percentage by mass of hydrogen in $C_{10}H_{16}O$

$$= \frac{16.0 \times 1.0}{\text{relative molecular mass of } C_{10}H_{16}O} \times 100\%$$

Percentage by mass of oxygen in $C_{10}H_{16}O$

$$= \frac{16.0}{\text{relative molecular mass of } C_{10}H_{16}O} \times 100\%$$

Thus, $C_{10}H_{16}O$ contains the same percentage by mass of hydrogen and oxygen.

Unit Exercise (p.102)

10 Hydrated salt X•*n*H2O contains 51.16% of water by mass. Given that the molar mass of X is 120.3 g, what is *n*? (Relative atomic masses: H = 1.0, O = 16.0)

A 2
B 5
C 7
D 10

Answer: C

(HKDSE, Paper 1A, 2014, 5)

Unit Exercise (p.102)

11 When heated, mercury(II) oxide decomposes to form mercury and oxgyen.

$2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

What mass of mercury(II) oxide is required to produce 0.230 mole of oxygen?

(Relative atomic masses: O = 16.0, Hg = 200.6)

Answer: D

A 49.8 g
B 62.4 g
C 84.2 g
D 99.6 g

Explanation:

According to the equation, 2 moles of HgO produce 1 mole of O_2.
i.e. 0.460 mole of HgO is required to produce 0.230 mole of O_2.

Molar mass of HgO = (200.6 + 16.0) g mol^{-1}
= 216.6 g mol^{-1}

Mass of HgO = number of moles of HgO x molar mass of HgO
= 0.460 mol x 216.6 g mol^{-1}
= 99.6 g

Unit Exercise (p.102)

12 $KClO_3$(s) decomposes to form KCl(s) and O_2(g) when heated.

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

What mass of $KClO_3$ has decomposed when 19.2 g of O_2 are produced?
(Relative atomic masses: O = 16.0, Cl = 35.5, K = 39.1)

A 49.0 g
B 62.2 g
C 75.3 g
D 113 g

Answer: A

Unit Exercise (p.102)

Explanation:

Molar mass of O_2 = (2 x 16.0) g mol^{-1} = 32.0 g mol^{-1}

Number of moles of $O_2 = \dfrac{19.2\text{ g}}{32.0\text{ g mol}^{-1}}$

$= 0.600$ mol

According to the equation, 2 moles of $KClO_3$ decompose to produce 3 moles of O_2.

i.e. 0.400 mole of $KClO_3$ decomposes to produce 0.600 mole of O_2.

Molar mass of $KClO_3$ = (39.1 + 35.5 + 3 x 16.0) g mol^{-1}

= 122.6 g mol^{-1}

Mass of $KClO_3$ = 0.400 mol x 122.6 g mol^{-1}

= 49.0 g

Unit Exercise (p.102)

13 Aluminium reacts with copper(II) oxide according to the equation below.

$$2Al(s) + 3CuO(s) \rightarrow Al_2O_3(s) + 3Cu(s)$$

What is the mass of aluminium needed to react with 35.8 g of copper(II) oxide?

(Relative atomic masses: O = 16.0, Al = 27.0, Cu = 63.5)

A 8.10 g
B 11.5 g
C 12.1 g
D 18.2 g

Answer: A

Unit Exercise (p.102)

Explanation:

1 Molar mass of CuO = (63.5 + 16.0) g mol^{-1} = 79.5 g mol^{-1}

$$\text{Number of moles of CuO} = \frac{\text{mass of CuO}}{\text{molar mass of CuO}} = \frac{35.8\ \text{g}}{79.5\ \text{g mol}^{-1}} = 0.450\ \text{mol}$$

2 According to the equation, 3 moles of CuO react with 2 moles of Al.

$$\text{i.e. number of moles of Al} = \frac{2}{3} \times 0.450\ \text{mol} = 0.300\ \text{mol}$$

3 Mass of Al = number of moles of Al x molar mass of Al

= 0.300 mol x 27.0 g mol^{-1}

= 8.10 g

Unit Exercise (p.102)

14 Ammonia reacts with oxygen according to the equation below.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

What is the mass of oxygen required to react with 20.4 g of ammonia?
(Relative atomic masses: H = 1.0, N = 14.0, O = 16.0)

Answer: B

A 38.4 g
B 48.0 g
C 66.5 g
D 94.8 g

Unit Exercise (p.102)

Explanation:

1 Molar mass of NH_3 = (14.0 + 3 x 1.0) g mol^{-1} = 17.0 g mol^{-1}

$$\text{Number of moles of } NH_3 = \frac{\text{mass of } NH_3}{\text{molar mass of } NH_3}$$

$$= \frac{20.4 \text{ g}}{17.0 \text{ g mol}^{-1}}$$

$$= 1.20 \text{ mol}$$

2 According to the equation, 4 moles of NH_3 react with 5 moles of O_2.

$$\text{i.e. number of moles of } O_2 = \frac{5}{4} \times 1.20 \text{ mol}$$

$$= 1.50 \text{ mol}$$

3 Molar mass of O_2 = 2 x 16.0 g mol^{-1} = 32.0 g mol^{-1}

Mass of O_2 = number of moles of O_2 x molar mass of O_2

= 1.50 mol x 32.0 g mol^{-1}

= 48.0 g

Unit Exercise (p.102)

15 Cerussite is an ore of lead. A sample of cerussite contains 68.0% of lead(II) carbonate by mass. Assuming that the other components in this ore do not contain lead, what is the mass of ore required to extract 29.1 g of lead?
(Relative atomic masses: C = 12.0, O = 16.0, Pb = 207.2)

Answer: D

A 19.8 g
B 23.4 g
C 42.8 g
D 55.0 g

Unit Exercise (p.102)

Explanation:

$PbCO_3(s) \longrightarrow Pb(s)$

1 Number of moles of Pb $= \dfrac{\text{mass of Pb}}{\text{molar mass of Pb}}$

$= \dfrac{29.1\ g}{207.2\ g\ mol^{-1}}$

$= 0.140\ mol$

2 1 mole of $PbCO_3$ gives 1 mole of Pb in extraction.

i.e. number of moles of $PbCO_3$ = 0.140 mol

3 Molar mass of $PbCO_3$ = $(207.2 + 12.0 + 3 \times 16.0)\ g\ mol^{-1}$

$= 267.2\ g\ mol^{-1}$

Mass of $PbCO_3$

= number of moles of $PbCO_3$ x molar mass of $PbCO_3$

$= 0.140\ mol \times 267.2\ g\ mol^{-1}$

$= 37.4\ g$

Mass of ore $= \dfrac{37.4}{68.0\%}\ g$

$= 55.0\ g$

Unit Exercise (p.102)

16 Refer to the following chemical equation:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

N moles of Fe_2O_3 are allowed to react with $2N$ moles of CO under suitable conditions until the reaction stops. How many moles of Fe are formed?

Answer: D

A N

B $2N$

C $\frac{2}{3}N$

D $\frac{4}{3}N$

(HKDSE, Paper 1A, 2014, 4)

Unit Exercise (p.102)

17 NO(g) reacts with O_2(g) to form NO_2(g).

$2NO(g) + O_2(g)\ 2NO_2(g)$

In an experiment, 6.0 moles of NO(g) are allowed to react with 2.5 moles of O_2(g). How many moles of NO_2(g) are formed?

Answer: B

A 4.5 moles
B 5.0 moles
C 5.5 moles
D 6.0 moles

(HKDSE, Paper 1A, 2014, 4)

Explanation:

According to the equation, 2 moles of NO react with 1 mole of O_2 to give 2 moles of NO_2.
In this reaction, 5.0 moles of NO react with 2.5 moles of O_2 to give 5.0 moles of NO_2.

Unit Exercise (p.102)

PART III STRUCTURED QUESTIONS

18 The mineral ilmenite is usually mined and processed for titanium. A sample of the main compound in ilmenite contains 5.41 g of iron, 4.64 g of titanium and 4.65 g of oxygen. Deduce the empirical formula of the compound in ilmenite.

(Relative atomic masses: O = 16.0, Ti = 47.9, Fe = 55.8)

Unit Exercise (p.102)

	Iron	Titanium	Oxygen	
1 Mass of element	5.41 g	4.64 g	4.65 g	
2 Number of moles of atoms	$\frac{5.41\text{ g}}{55.8\text{ g mol}^{-1}} = 0.0970\text{ mol}$	$\frac{4.64\text{ g}}{47.9\text{ g mol}^{-1}} = 0.0969\text{ mol}$	$\frac{4.65\text{ g}}{16.0\text{ g mol}^{-1}} = 0.291\text{ mol}$	(1)
3 Mole ratio of atoms	$\frac{0.0970}{0.0969} = 1.00$	$\frac{0.0969}{0.0969} = 1.00$	$\frac{0.291}{0.0969} = 3.00$	(1)

∴ the empirical formula of the compound is $FeTiO_3$.

Unit Exercise (p.102)

19 A compound of phosphorus and chlorine has a relative molecular mass smaller than 250. It contains 22.6% of phosphorus by mass.

a) Deduce the molecular formula of the compound.

b) Draw the electron diagram for the compound, showing electrons in the *outermost shells* only.

(HKDSE, Paper 1B, 2016, 1(c))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.102)

20 a) A compound contains tin (Sn) and oxygen only. In an experiment for determining the empirical formula of this compound, 4.27 g of the compound were heated with carbon. Carbon monoxide and 3.36 g of tin were formed.
Calculate the empirical formula of this compound.
(Relative atomic masses: O = 16.0, Sn = 118.7)

Mass of oxygen in the compound = (4.27 – 3.36) g = 0.91 g

	Tin	Oxygen	
1 Mass of element	3.36 g	0.91 g	
2 Number of moles of atoms	$\frac{3.36\text{ g}}{118.7\text{ g mol}^{-1}}$ = 0.0283 mol	$\frac{0.91\text{ g}}{16.0\text{ g mol}^{-1}}$ = 0.0569 mol	(1)
3 Mole ratio of atoms	$\frac{0.0283}{0.0283}$ = 1.00	$\frac{0.0569}{0.0283}$ = 2.01	(1)

∴ the empirical formula of this compound is SnO_2.

Unit Exercise (p.102)

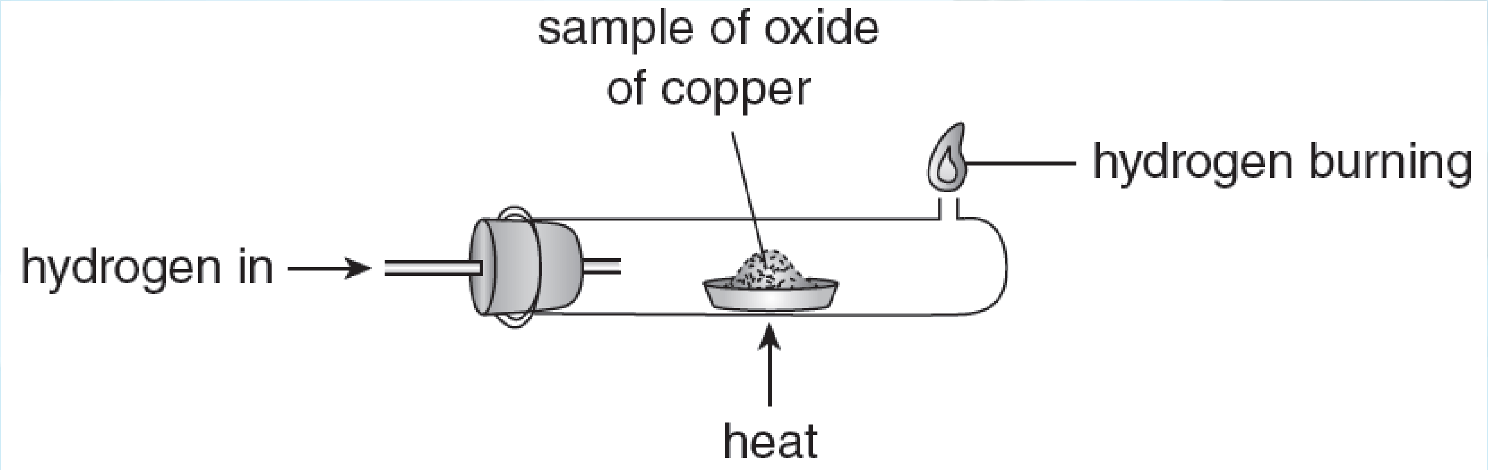
b) Draw the hazard warning label that should be displayed on a cylinder containing carbon monoxide gas.

or

(1)

Unit Exercise (p.102)

21 A sample of 3.93 g of an oxide of copper was reduced by heating in a stream of hydrogen gas, using the apparatus shown below. After heating, 3.49 g of copper were produced.



Calculate the empirical formula of the oxide of copper.
(Relative atomic masses: O = 16.0, Cu = 63.5)

Unit Exercise (p.102)

Mass of oxygen in oxide = (3.93 – 3.49) g = 0.44 g

	Copper	Oxygen	
1 Mass of element	3.49 g	0.44 g	
2 Number of moles of atoms	$\frac{3.49\text{ g}}{63.5\text{ g mol}^{-1}} = 0.0550\text{ mol}$	$\frac{0.44\text{ g}}{16.0\text{ g mol}^{-1}} = 0.0275\text{ mol}$	(1)
3 Mole ratio of atoms	$\frac{0.0550}{0.0275} = 2.00$	$\frac{0.0275}{0.0275} = 1.00$	(1)

∴ the empirical formula of the oxide is Cu_2O.

Unit Exercise (p.102)

22 The percentage by mass of water in a sample of hydrated calcium phosphate is 18.8%.
A 9.00 g sample of hydrated calcium phosphate is heated to remove all the water of crystallisation. What is the mass of anhydrous calcium phosphate obtained?

Percentage by mass of calcium phosphate in sample
= (100 – 18.8)% = 81.2%

Mass of anhydrous calcium phosphate in sample
= 9.00 g × 81.2% (1)
= 7.31 g (1)

Unit Exercise (p.102)

23 The chemical formula of hydrated barium chloride is $BaCl_2•xH_2O$. Hydrated barium chloride contains 14.7% by mass of water.
Calculate the value of x.
(Relative atomic masses: H = 1.0, O = 16.0, Cl = 35.5, Ba = 137.3)

Formula mass of $BaCl_2•xH_2O = 137.3 + 2 \times 35.5 + x(2 \times 1.0 + 16.0)$
$= 208.3 + 18x$

Percentage by mass of water in $BaCl_2•xH_2O = 14.7\% = \dfrac{18x}{208.3 + 18x} \times 100\%$ (1)

$14.7(208.3 + 18x) = 100(18x)$
$x = 1.99$ (1)
$\therefore$ the value of x is 2.

Unit Exercise (p.102)

24 A sample of hydrated nickel sulphate ($NiSO_4 \cdot xH_2O$) with a mass of 2.287 g was heated to remove all water of crystallisation. The solid remaining had a mass of 1.344 g.

a) Calculate the value of the integer *x*. Show your working.

(Relative atomic masses: H = 1.0, O = 16.0, S = 32.1, Ni = 58.7)

Mass of water removed = (2.287 – 1.344) g = 0.943 g

	$NiSO_4$	H_2O	
1 Mass of compound	1.344 g	0.943 g	
2 Number of moles of compound	$\frac{1.344 \text{ g}}{154.8 \text{ g mol}^{-1}} = 8.68 \times 10^{-3}$ mol	$\frac{0.943 \text{ g}}{18.0 \text{ g mol}^{-1}} = 0.0524$ mol	(1)
3 Mole ratio of compound	$\frac{8.68 \times 10^{-3}}{8.68 \times 10^{-3}} = 1.00$	$\frac{0.0524}{8.68 \times 10^{-3}} = 6.04$	(1)

∴ the value of *x* is 6.

Unit Exercise (p.102)

b) Suggest how a student doing this experiment could check that all the water had been removed.

(AQA Advanced Subsidiary GCE, Unit 1, Jun. 2015, 4(a)–(b))

Repeat heating the solid. (1)
Check that the mass is unchanged. (1)

Unit Exercise (p.102)

25 When heated, iron(III) nitrate is converted into iron(III) oxide, nitrogen dioxide and oxygen.

$$4Fe(NO_3)_3(s)\ 2Fe_2O_3(s) + 12NO_2(g) + 3O_2(g)$$

A 2.90 g sample of iron(III) nitrate was completely converted to the products shown.

(Relative atomic masses: N = 14.0, O = 16.0, Fe = 55.8; Avogadro constant = $6.02 \times 10^{23}\ mol^{-1}$)

a) Calculate

i) the mass of iron(III) oxide,

ii) the number of oxygen molecules produced in this reaction.

b) Suggest why the iron(III) oxide obtained is pure.

Unit Exercise (p.102)

a) Molar mass of $Fe(NO_3)_3$ = (55.8 + 3 x 14.0 + 9 x 16.0) g mol^{-1} = 241.8 g mol^{-1}

Number of moles of $Fe(NO_3)_3$ = $\dfrac{\text{mass of } Fe(NO_3)_3}{\text{molar mass of } Fe(NO_3)_3}$ = $\dfrac{2.90 \text{ g}}{241.8 \text{ g mol}^{-1}}$

= 0.0120 mol (1)

According to the equation, 4 moles of $Fe(NO_3)_3$ are converted to 2 moles of Fe_2O_3 and 3 moles of O_2 when heated.

i.e. number of moles of Fe_2O_3 = $\dfrac{0.0120}{2}$ mol = 0.00600 mol (1)

number of moles of O_2 = $\dfrac{3}{4}$ x 0.0120 mol = 0.00900 mol (1)

Unit Exercise (p.102)

i) Molar mass of Fe_2O_3 = (2 x 55.8 + 3 x 16.0) g mol^{-1} = 159.6 g mol^{-1}
Mass of Fe_2O_3 = number of moles of Fe_2O_3 x molar mass of Fe_2O_3
= 0.00600 mol x 159.6 g mol^{-1}
= 0.958 g (1)

ii) Number of O_2 molecules = 0.00900 mol x 6.02 x 10^{23} mol
= 5.42 x 10^{21} (1)

b) Other products are gases. (1)

Unit Exercise (p.102)

26 The combustion of glucose ($C_6H_{12}O_6$) can be represented by the equation:

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$$

What mass of oxygen is required for the combustion of 64.8 g of glucose?

(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

Method 1

1) # of moles of $C_6H_{12}O_6$

$$= \frac{\text{mass of } C_6H_{12}O_6}{\text{molar mass of } C_6H_{12}O_6} = \frac{64.8\text{ g}}{(6\times12.0 + 12\times1.0 + 6\times16.0)\text{ g mol}^{-1}}$$

$$= \frac{64.8\text{ g}}{180.0\text{ g mol}^{-1}} = 0.360\text{ mol} \quad (1)$$

2) # of moles of O_2 = 6 $\times$ 0.360 mol = 2.16 mol (1)

3) Mass of O_2 = 2.16 mol $\times$ 32.0 g mol^{-1} = 69.1 g (1)

Unit Exercise (p.102)

Method 2

1) Molar mass of $C_6H_{12}O_6$ = $(6\times12.0 + 12\times1.0 + 6\times16.0)$ g mol^{-1}= 180.0 g mol^{-1}

Molar mass of O_2 = (2×16.0) g mol^{-1}= 32.0 g mol^{-1}

2) 180.0 g of $C_6H_{12}O_6$ require 6×32.0 g of O_2. (1)

3) Mass of O_2 = $64.8\text{ g} \times \dfrac{6 \times 32.0\text{ g}}{180.0\text{ g}}$ (1) = 69.1 g (1)

$\therefore$ 69.1 g of oxygen are required.

Unit Exercise (p.102)

27 Some airbags in cars contain sodium azide (NaN3). Sodium azide is made by reacting dinitrogen monoxide gas with sodium amide ($NaNH_2$) according to the equation below.

$$2NaNH_2 + N_2O \rightarrow NaN_3 + NaOH + NH_3$$

Calculate the mass of sodium amide needed to make 1.17 tonnes of sodium azide.

(Relative atomic masses: H = 1.0, N = 14.0, O = 16.0, Na = 23.0;
1 tonne = 10^6 g)

Unit Exercise (p.102)

Method 1

1 Molar mass of NaN_3 = (23.0 + 3 x 14.0) g mol^{-1} = 65.0 g mol^{-1}

$$\text{Number of moles of } NaN_3 = \frac{\text{mass of } NaN_3}{\text{molar mass of } NaN_3}$$
$$= \frac{1.17 \times 10^6 \text{ g}}{65.0 \text{ g mol}^{-1}}$$
$$= 1.80 \times 10^4 \text{ mol}$$ (1)

2 According to the equation, 2 moles of $NaNH_2$ react with N_2O to make 1 mole of NaN_3.

i.e. number of moles of $NaNH_2$ = 2 x 1.80 x 10^4 mol
= 3.60 x 10^4 mol (1)

3 Molar mass of $NaNH_2$ = (23.0 + 14.0 + 2 x 1.0) g mol^{-1} = 39.0 g mol^{-1}

Mass of $NaNH_2$ = number of moles of $NaNH_2$ × molar mass of $NaNH_2$
= 3.60 x 10^4 mol x 39.0 g mol^{-1}
= 1.40 x 10^6 g
= 1.40 tonnes (1)

Unit Exercise (p.102)

Method 2

1 Molar mass of $NaNH_2$ = (23.0 + 14.0 + 2 x 1.0) g mol^{-1} = 39.0 g mol^{-1}

Molar mass of NaN_3 = (23.0 + 3 x 14.0) g mol^{-1} = 65.0 g mol^{-1}

2 According to the equation, 2 moles of $NaNH_2$ react with N_2O to make 1 mole of NaN_3.

i.e. 2 x 39.0 g of $NaNH_2$ react with N_2O to make 65.0 g of NaN_3. (1)

3 Mass of $NaNH_2$ = 1.17 tonnes x $\frac{2 \times 39.0\text{ g}}{65.0\text{ g}}$ (1)

= 1.40 tonnes (1)

∴ 1.40 tonnes of $NaNH_2$ are needed.

Unit Exercise (p.102)

28 a) In an experiment, 1.25 g of lithium nitride is formed when a piece of lithium is burnt in air.

i) Write a chemical equation for the reaction involved.

ii) Calculate the mass of lithium that reacted with nitrogen.

(Relative atomic masses: Li = 6.9, N = 14.0)

b) Name another compound which will also be formed when lithium is burnt in air.

(HKDSE, Paper 1B, 2018, 1(b)–(c))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.102)

29 A bracelet, originally made of pure silver, became tarnished over time with black silver sulphide forming on the surface. The bracelet was cleaned by converting the silver sulphide back to silver using aluminium in the reaction below. The mass of the bracelet decreased by 0.0128 g in the cleaning process.

$$3Ag_2S(s) + 2Al(s) \rightarrow 6Ag(s) + Al_2S_3(s)$$

(Relative atomic masses: Al = 2.70, S = 32.1, Ag = 107.9)

a) How many moles of sulphur were removed in the cleaning process?

$$\text{\# of moles of S} = \frac{\text{mass of S}}{\text{molar mass of S}} = \frac{0.0128\ \text{g}}{32.1\ \text{g mol}^{-1}}$$

$$= 3.99 \times 10^{-4}\ \text{mol} \qquad (1)$$

Unit Exercise (p.102)

b) What mass of aluminium was used in the reaction?

Number of moles of Ag_2S removed = number of moles of S removed
$= 3.99 \times 10^{-4}$ mol

According to the equation, 3 moles of Ag_2S react with 2 moles of Al.

i.e. number of moles of Al $= \dfrac{2}{3} \times 3.99 \times 10^{-4}$ mol
$= 2.66 \times 10^{-4}$ mol (1)

Mass of Al = number of moles of Al x molar mass of Al
$= 2.66 \times 10^{-4} \text{ mol} \times 27.0 \text{ g mol}^{-1}$
$= 7.18 \times 10^{-3}$ g (1)

$\therefore$ 7.18×10^{-3} g of aluminium was used.

Unit Exercise (p.102)

30 Zinc chloride can be prepared by the reaction between zinc and hydrogen chloride gas.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

An impure sample of zinc powder of 6.24 g reacted with hydrogen chloride gas until the reaction was complete. 11.6 g of zinc chloride were produced.

What is the percentage purity of the sample of zinc powder?
(Relative atomic masses: Cl = 35.5, Zn = 65.4)

Unit Exercise (p.102)

Molar mass of $ZnCl_2$ = (65.4 + 2 x 35.5) g mol^{-1}
= 136.4 g mol^{-1}

$$\text{Number of moles of } ZnCl_2 = \frac{\text{mass of } ZnCl_2}{\text{molar mass of } ZnCl_2}$$

$$= \frac{11.6 \text{ g}}{136.4 \text{ g mol}^{-1}}$$

$$= 0.0850 \text{ mol} \quad (1)$$

According to the equation, 1 mole of Zn reacts with HCl to produce 1 mole of $ZnCl_2$.

Unit Exercise (p.102)

i.e. number of moles of Zn = 0.0850 mol (1)

Mass of Zn = number of moles of Zn x molar mass of Zn

$= 0.0850 \text{ mol} \times 65.4 \text{ g mol}^{-1}$

$= 5.56 \text{ g}$ (1)

$$\text{Percentage purity of sample} = \frac{5.56 \text{ g}}{6.24 \text{ g}} \times 100\%$$

$= 89.1\%$ (1)

∴ the percentage purity of the sample of zinc powder is 89.1%.

Unit Exercise (p.102)

31 One use of sodium azide (NaN_3) is to make lead(II) azide, which can be used as a detonator for explosives. Lead(II) azide has the chemical formula of $Pb(N_3)_2$.

Lead(II) azide can be made by the following reaction:

$$Pb(NO_3)_2(aq) + 2NaN_3(aq) \quad Pb(N_3)_2(s) + 2NaNO_3(aq)$$

a) What method would you use to separate the lead(II) azide from the reaction mixture?

Filtration

b) 10.0 g of $Pb(NO_3)_2$ in solution is mixed with 5.20 g of NaN_3 in solution. What is the mass of $Pb(N_3)_2$ produced?
(Relative atomic masses: N = 14.0, O = 16.0, Na = 23.0, Pb = 207.2)

Unit Exercise (p.102)

Molar mass of $Pb(NO_3)_2$ = (207.2 + 2 x 14.0 + 6 x 16.0) g mol^{-1} = 331.2 g mol^{-1}

$$\text{Number of moles of } Pb(NO_3)_2 \text{ used} = \frac{\text{mass of } Pb(NO_3)_2}{\text{molar mass of } Pb(NO_3)_2} = \frac{10.0 \text{ g}}{331.2 \text{ g mol}^{-1}} = 0.0302 \text{ mol} \qquad (1)$$

Molar mass of NaN_3 = (23.0 + 3 x 16.0) g mol^{-1}
= 65.0 g mol^{-1}

$$\text{Number of moles of } NaN_3 = \frac{\text{mass of } NaN_3}{\text{molar mass of } NaN_3} = \frac{5.20 \text{ g}}{65.0 \text{ g mol}^{-1}} = 0.0800 \text{ mol} \qquad (1)$$

Unit Exercise (p.102)

According to the equation, 1 mole of $Pb(NO_3)_2$ reacts with 2 moles of NaN_3 to produce 1 mole of $Pb(N_3)_2$.

In this reaction, 0.0302 mole of $Pb(NO_3)_2$ reacts with 0.0604 mole of NaN_3.

Thus, NaN_3 is in excess. The amount of $Pb(NO_3)_2$ limits the amount of $Pb(N_3)_2$ produced. (1)

Number of moles of $Pb(N_3)_2$ = 0.0302 mol (1)

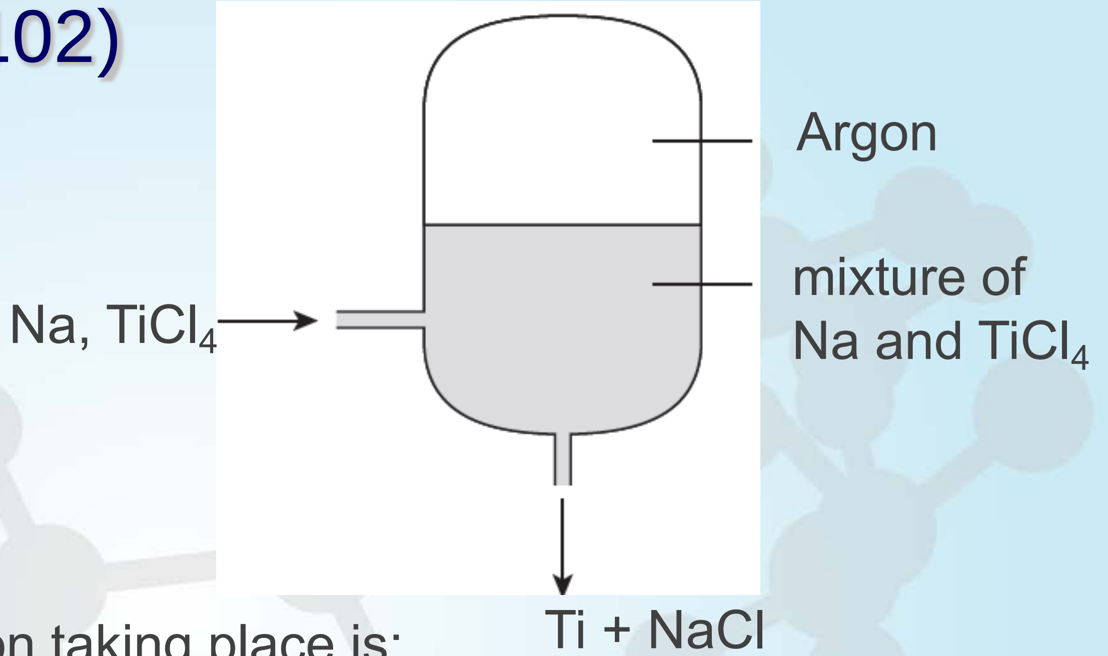
Molar mass of $Pb(N_3)_2$ = $(207.2 + 6 \times 14.0)$ g mol^{-1}
= 291.2 g mol^{-1}

Mass of $Pb(N_3)_2$ = number of moles of $Pb(N_3)_2$ x molar mass of $Pb(N_3)_2$
= 0.0302 mol x 291.2 g mol^{-1}
= 8.79 g (1)

∴ 8.79 g of $Pb(N_3)_2$ are produced.

Unit Exercise (p.102)

32 The schematic diagram shows a reactor used to produce titanium from titanium(IV) chloride.



a) The equation for the reaction taking place is:

$TiCl_4(g)$ + __4__ Na(l) → Ti(s) + __4__ NaCl(l) (1)

i) Balance the equation.

ii) What does this reaction suggest about the reactivity of titanium compared to that of sodium? Titanium is less reactive than sodium (1)

b) Suggest why the reaction is done in an atmosphere of dry argon instead of air containing water vapour. Argon is unreactive. (1)

Water vapour would react with sodium. /

Air contains oxygen that would react with the reactants. (1)

Unit Exercise (p.102)

c) In one reaction, 163 kg of titanium(IV) chloride were allowed to react with 81.0 kg of sodium.
Calculate the mass of titanium produced from this reaction.
(Relative atomic masses: Na = 23.0, Cl = 35.5, Ti = 47.9)

Molar mass of $TiCl_4$ = (47.9 + 4 x 35.5) g mol^{-1}
= 189.9 g mol^{-1}

$$\text{Number of moles of } TiCl_4 = \frac{\text{mass of } TiCl_4}{\text{molar mass of } TiCl_4} = \frac{163\ 000 \text{ g}}{189.9 \text{ g mol}^{-1}} = 858 \text{ mol} \quad (1)$$

$$\text{Number of moles of Na} = \frac{\text{mass of Na}}{\text{molar mass of Na}} = \frac{81\ 000 \text{ g}}{23.0 \text{ g mol}^{-1}} = 3\ 520 \text{ mol} \quad (1)$$

Unit Exercise (p.102)

According to the equation, 1 mole of $TiCl_4$ reacts with 4 moles of Na.

In this reaction, 858 moles of $TiCl_4$ react with 4 x 858 moles of Na.

Thus, Na is in excess. $TiCl_4$ is the limiting reactant. (1)

The amount of Ti produced is determined by the amount of $TiCl_4$.

Number of moles of Ti = 858 mol (1)

Mass of Ti = number of moles of Ti x molar mass of Ti
= 858 mol x 47.9 g mol^{-1}
= 41 100 g (1)
= 41.1 kg

∴ 41.1 kg of titanium were produced.

Unit Exercise (p.102)

33 An experiment is carried out to determine the empirical formula of magnesium oxide: magnesium + oxygen → magnesium oxide

The following results are obtained

mass of magnesium ribbon reacted = 0.420 g

mass of magnesium oxide formed = 0.700 g

Describe an experiment to produce these results. As part of your answer, show how these results can be used to obtain the empirical formula of the magnesium oxide.

(Relative atomic masses: O = 16.0, Mg = 24.3)

(Edexcel GCSE (Higher Tier), C2, Jun. 2016, 6(d))

Unit Exercise (p.102)

Experimental method

- Find the mass of the crucible and lid. (1)
- Find the mass of the crucible, lid and magnesium. (1)
- Heat the magnesium. (1)
- Lift the lid occasionally to allow oxygen in. Ensure that as little product escapes as possible. (1)
- Heat until there is no further change. (1)
- Allow to cool. (1)
- Find the mass of the crucible, lid and magnesium oxide. (1)
- Repeat heating until a constant mass is obtained. (1)

Unit Exercise (p.102)

Calculations

- Mass of magnesium
  = mass of crucible, lid and magnesium – mass of crucible and lid
- Mass of magnesium oxide
  = mass of crucible, lid and magnesium oxide – mass of crucible and lid
- Mass of oxygen = mass of magnesium oxide – mass of magnesium
  = (0.700 – 0.420) g = 0.280 g
- Mole ratio of magnesium atoms to oxygen atoms

$$= \frac{0.420\text{ g}}{24.3\text{ g mol}^{-1}} : \frac{0.280\text{ g}}{16.0\text{ g mol}^{-1}} \qquad (1)$$

  = 0.0173 : 0.0175
- Ratio of magnesium atoms : oxygen atoms is 1 : 1.
- The empirical formula of the oxide is MgO. (1)

Communication mark (1)