O Level Chemistry
Chap 10: Chemical Calculations

Basic Chemical Calculations

1) $2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$ means that
   - 2 mol/48g of magnesium reacts with 1 mol/32g of oxygen to produce 2 mol/80g of magnesium oxide.

2) $1dm^3 = 1000cm^3$

   Therefore,

   $50cm^3 = \frac{50}{1000}\ dm^3$ and $50dm^3 = 50 \times 1000$

   $= 0.05dm^3$ $= 50000cm^3$

2) Example: Calculate the mass of water produced in the complete combustion of 4g of methane, given the equation:

   $CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (l)$

   Amt of $CH_4 = \frac{4}{16}$
   $= 0.25$ mol

   According to the eqn, 1 mol of CH_4 forms 2 mol of H_2O.
   0.25 mol of CH_4 forms 0.5 mol of H_2O.

   Mass of water produced = 0.5mol x 18
   = 9g

3) For chemical calculations involving gases, we can use the volume of the gas to compare the ratio of the reactants and products (change moles to volume). This is because the volume of gas is proportional to its number of moles. [For gases to compare volume only!]

4) Example: Hydrogen reacts with water to form water according to the equation below. Given that the volume of hydrogen is $10cm^3$, calculate
   (a) the volume of oxygen gas required for the reaction
   (b) the amount of water (in moles) produced.

   $2H_2 + 0_2 \rightarrow 2H_2O$

   (a) According to the eqn, 2 mol of hydrogen reacts with 1 mol of oxygen.
   Volume of oxygen required $= \frac{10}{2}$
   $= 5cm^3$

   (b) Amt of hydrogen reacted $= \frac{10}{24000}$
   $\approx 4.167 \times 10^{-4}$ mol
   According to the eqn, 2 mol of hydrogen forms 2 mol of water.
   $\therefore$ Amt of water produced $= 4.167 \times 10^{-4}$ mol

http://sites.google.com/site/olevelexamnotes
Resources for the A1 candidate.

Concentration of solutions

5) The concentration of a solution is the amount of solute dissolved per unit volume in a solution.

6)
$$\textbf{Concentration (mol/dm}^3\textbf{)} = \frac{\textbf{Amt of solute (mol)}}{\textbf{Volume (dm}^3\textbf{)}}$$
$$\textbf{Concentration (g/dm}^3\textbf{)} = \frac{\textbf{Mass of solute (g)}}{\textbf{Volume (dm}^3\textbf{)}}$$
$$\textbf{Concentration (mol/dm}^3\textbf{)} = \frac{\textbf{Concentration (g/dm}^3\textbf{)}}{\mathbf{M_r}}$$

Note: Concentration in mol/dm^3 is also known as molar concentration (M).

7) Example: A $50cm^3$ solution is formed by 8g of NaOH. Calculate
   (a) the concentration in g/dm^3
   (b) the molar concentration (mol/dm^3), using your answer in (a)
   (c) the mass of NaOH in $75cm^3$ of the solution.

(a) Concentration (g/dm^3) = $\frac{8g}{50/1000}$

=$160g/dm^3$

(b) Concentration (mol/dm^3) = $\frac{160}{23+16+1}$

= $4mol/dm^3$

(c) Mass of NaOH = concentration x volume

= $160 \times \frac{75}{1000}$

= 12g

Limiting Reactant

8) Limiting reactant is the reactant that is completely used up in a reaction and which limits the amount of products formed.

9) Example: Zinc reacts with hydrochloric acid according to the equation below.

$Zn (s) + 2HCl (aq) \rightarrow ZnCl_2 (aq) + H_2 (g)$

Given that 0.05mol of zinc was added to 0.075mol of hydrochloric acid,

a) Identify the limiting reactant. Calculate the amount (in moles) of excess reactant which remained unreacted.
b) Calculate the mass of $ZnCl_2$ produced.

a) According to the eqn, 1 mol of Zn reacts with 2 mol of HCl.

∴ 0.05mol of Zn reacts with 0.10mol of HCl.

HCl is the limiting reactant.

Amt of Zn in excess = 0.05 – 0.0375

= 0.0125mol

b) According to the eqn, 2mol of HCl forms 1mol of $ZnCl_2$.

∴ 0.075mol of HCL forms 0.0375mol of $ZnCl_2$.

Mass of $ZnCl_2$ = 0.0375 x [65+2(35.5)]

=5.1g

10) Example: A mixture of 8.0g of hydrogen and 8.0g of oxygen is ignited according to the given equation. Calculate the mass of water formed.
$2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l)$

Amt of $H_2 = \frac{8.0}{2}$
$= 4$mol
Amt of $O_2 = \frac{8.0}{32}$
$= 0.25$mol
According to the eqn, 2mol of H_2 reacts with 1 mol of O_2.
4mol of H_2 reacts with 2mol of O_2.
$\therefore$ O_2 is the limiting reactant.

According to the eqn, 1 mol of O_2 forms 2 mol of H_2O.
Amt of H_2O formed = 0.25 x 2
= 0.5mol
Mass of H_2O formed = 0.5 x 18
=9g

Percentage yield and purity

11) $$\%\ \textbf{yield} = \frac{\textbf{Actual yield}}{\textbf{Theoretical yield}} \times \mathbf{100\%}$$
The actual yield is always lower than the theoretical yield (calculated yield according to an eqn) due to:
- Impure reactants
- Human error (transferring of solutions),
- Loss of reactants and products through evaporation
- Reversible reactions

12) $$\%\ \textbf{purity} = \frac{\textbf{Mass of pure substance}}{\textbf{Mass of sample}} \times \mathbf{100\%}$$

13) Example: When 128g of sulphur dioxide was reacted with excess oxygen, 140g of sulphur trioxide was produced.
$2SO_2 + O_2 \rightarrow 2SO_3$
Calculated the percentage yield of sulphur dioxide.

Amt of $SO_2 = \frac{128}{32+2(16)}$
$= 2$mol
According to the eqn, 2mol of SO_2 forms 2 mol of SO_3.
Theoretical mass of SO_3 = 2 x [32+3(16)]
= 160g
% yield = $\frac{\text{Actual yield}}{\text{Theoretical yield}}$ x 100%
= $\frac{140g}{160g}$ x 100%
= 87.5%

14) When excess hydrochloric acid was added to 6g of impure calcium carbonate, $1200cm^3$ of gas was produced.

$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Calculate the percentage purity of the calcium carbonate sample.

Amt of CO_2 produced = $\frac{1200}{24000}$

= 0.05mol

According to the eqn, 1 mol of $CaCO_3$ forms 1mol of CO_2.

0.05mol of $CaCO_3$ forms 0.05mol of CO_2.

% purity of $CaCO_3$ = $\frac{\text{Mass of pure substance}}{\text{Mass of sample}}$

= $\frac{0.05 \text{ x } 100}{6}$ x 100%

= 83.3% (3sf)

15) A 5.00g sample of copper was contaminated with copper (II) oxide, which was found to react with 0.020mol of hydrochloric acid.

$CuO\ (s) + 2HCl\ (ag) \rightarrow CuCl_2\ (aq) + H_2O\ (l)$

Calculate the percentage purity of copper metal in the sample.

According to the eqn, 1mol of CuO reacts with 2mol of HCl.

0.010mol of CuO reacts with 0.020mol of HCl.

Mass of CuO = 0.010 x (64+16)

= 0.80g

Mass of pure copper = 5.00g – 0.80g

= 4.2g

% purity of copper metal = $\frac{4.2}{5.0}$ x 100%

= 84%

Notes: