

O Level Chemistry

Chap 10: Chemical Calculations

Basic Chemical Calculations

- 1) $2\text{Mg (s)} + \text{O}_2 \text{ (g)} \rightarrow 2\text{MgO (s)}$ means that
- 2 mol/48g of magnesium reacts with 1 mol/32g of oxygen to produce 2 mol/80g of magnesium oxide.
- 2) $1\text{dm}^3 = 1000\text{cm}^3$
Therefore,
 $50\text{cm}^3 = \frac{50}{1000} \text{dm}^3$ and $50\text{dm}^3 = 50 \times 1000$
 $= 0.05\text{dm}^3$ $= 50000\text{cm}^3$
- 2) Example: Calculate the mass of water produced in the complete combustion of 4g of methane, given the equation:
 $\text{CH}_4 \text{ (g)} + 2\text{O}_2 \text{ (g)} \rightarrow \text{CO}_2 \text{ (g)} + 2\text{H}_2\text{O (l)}$
- Amt of $\text{CH}_4 = \frac{4}{16}$
 $= 0.25 \text{ mol}$
- 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 = $0.5\text{mol} \times 18$
 $= 9\text{g}$
- 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.
- $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
- (a) According to the eqn, 2 mol of hydrogen reacts with 1 mol of oxygen.
Volume of oxygen required = $\frac{10}{2}$
 $= 5\text{cm}^3$
- (b) Amt of hydrogen reacted = $\frac{10}{24000}$
 $\approx 4.167 \times 10^{-4} \text{ mol}$
- According to the eqn, 2 mol of hydrogen forms 2 mol of water.
 \therefore Amt of water produced = $4.167 \times 10^{-4} \text{ mol}$

Concentration of solutions

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

$$6) \quad \text{Concentration (mol/dm}^3) = \frac{\text{Amt of solute (mol)}}{\text{Volume (dm}^3)}$$

$$\text{Concentration (g/dm}^3) = \frac{\text{Mass of solute (g)}}{\text{Volume (dm}^3)}$$

$$\text{Concentration (mol/dm}^3) = \frac{\text{Concentration (g/dm}^3)}{M_r}$$

Note: Concentration in mol/dm³ is also known as molar concentration (M).

7) Example: A 50cm³ solution is formed by 8g of NaOH. Calculate

(a) the concentration in g/dm³

(b) the molar concentration (mol/dm³), using your answer in (a)

(c) the mass of NaOH in 75cm³ of the solution.

$$\begin{aligned} \text{(a) Concentration (g/dm}^3) &= \frac{8\text{g}}{50/1000} \\ &= 160\text{g/dm}^3 \end{aligned}$$

$$\begin{aligned} \text{(b) Concentration (mol/dm}^3) &= \frac{160}{23+16+1} \\ &= 4\text{mol/dm}^3 \end{aligned}$$

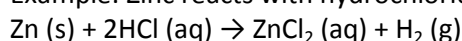
(c) Mass of NaOH = concentration x volume

$$\begin{aligned} &= 160 \times \frac{75}{1000} \\ &= 12\text{g} \end{aligned}$$

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.



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 ZnCl₂ produced.

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

$$\therefore 0.05\text{mol of Zn reacts with } 0.10\text{mol of HCl.}$$

HCl is the limiting reactant.

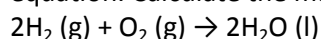
$$\begin{aligned} \text{Amt of Zn in excess} &= 0.05 - 0.0375 \\ &= 0.0125\text{mol} \end{aligned}$$

b) According to the eqn, 2mol of HCl forms 1mol of ZnCl₂.

$$\therefore 0.075\text{mol of HCl forms } 0.0375\text{mol of ZnCl}_2.$$

$$\begin{aligned} \text{Mass of ZnCl}_2 &= 0.0375 \times [65 + 2(35.5)] \\ &= 5.1\text{g} \end{aligned}$$

- 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.



$$\begin{aligned}\text{Amt of H}_2 &= \frac{8.0}{2} \\ &= 4\text{mol}\end{aligned}$$

$$\begin{aligned}\text{Amt of O}_2 &= \frac{8.0}{32} \\ &= 0.25\text{mol}\end{aligned}$$

According to the eqn, 2mol of H₂ reacts with 1 mol of O₂.

4mol of H₂ reacts with 2mol of O₂.

∴ O₂ is the limiting reactant.

According to the eqn, 1 mol of O₂ forms 2 mol of H₂O.

$$\begin{aligned}\text{Amt of H}_2\text{O formed} &= 0.25 \times 2 \\ &= 0.5\text{mol}\end{aligned}$$

$$\begin{aligned}\text{Mass of H}_2\text{O formed} &= 0.5 \times 18 \\ &= 9\text{g}\end{aligned}$$

Percentage yield and purity

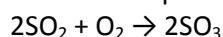
11)
$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

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

- 13) Example: When 128g of sulphur dioxide was reacted with excess oxygen, 140g of sulphur trioxide was produced.



Calculate the percentage yield of sulphur dioxide.

$$\begin{aligned}\text{Amt of SO}_2 &= \frac{128}{32+2(16)} \\ &= 2\text{mol}\end{aligned}$$

According to the eqn, 2mol of SO₂ forms 2 mol of SO₃.

$$\begin{aligned}\text{Theoretical mass of SO}_3 &= 2 \times [32+3(16)] \\ &= 160\text{g}\end{aligned}$$

$$\begin{aligned}\% \text{ yield} &= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% \\ &= \frac{140\text{g}}{160\text{g}} \times 100\% \\ &= 87.5\%\end{aligned}$$

- 14) When excess hydrochloric acid was added to 6g of impure calcium carbonate, 1200cm³ of gas was produced.



Calculate the percentage purity of the calcium carbonate sample.

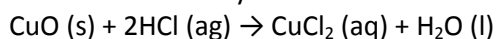
$$\begin{aligned}\text{Amt of CO}_2 \text{ produced} &= \frac{1200}{24000} \\ &= 0.05\text{mol}\end{aligned}$$

According to the eqn, 1 mol of CaCO₃ forms 1mol of CO₂.

0.05mol of CaCO₃ forms 0.05mol of CO₂.

$$\begin{aligned}\% \text{ purity of CaCO}_3 &= \frac{\text{Mass of pure substance}}{\text{Mass of sample}} \\ &= \frac{0.05 \times 100}{6} \times 100\% \\ &= 83.3\% \text{ (3sf)}\end{aligned}$$

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



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.

$$\begin{aligned}\text{Mass of CuO} &= 0.010 \times (64+16) \\ &= 0.80\text{g}\end{aligned}$$

$$\begin{aligned}\text{Mass of pure copper} &= 5.00\text{g} - 0.80\text{g} \\ &= 4.2\text{g}\end{aligned}$$

$$\begin{aligned}\% \text{ purity of copper metal} &= \frac{4.2}{5.0} \times 100\% \\ &= 84\%\end{aligned}$$

Notes: