

Stoichiometry

Isotopes

Same atomic number (number of protons or electrons)

Different atomic weights (atomic masses)

The Mol and Avogadro's Number

1 mol = 6.02×10^{23} (Avogadro's number)

1 mol = 1 molecular mass (or 1 mw)

1 atomic mass of an element = 1 mol of that element is the same as the atomic weight of that element (except for diatomic molecules)

Ex.

1 mol Na = 23 g = 1 at. mass of Na = 6.02×10^{23} Na atoms

1 mol Ag = 108 g = 1 at. mass of Ag = 6.02×10^{23} Ag atoms

1 mol of hydrogen (H₂) = 2 g (1 x 2) = 6.02×10^{23} hydrogen molecules

1 mol of oxygen (O₂) = 32 g (16 x 2) = 6.02×10^{23} oxygen molecules

1 mol of chlorine (Cl₂) = 71 g (35.5 x 2) = 6.02×10^{23} chlorine molecules

1 mol of any compound = 6.02×10^{23} molecules

Ex.

1 mol CO₂ = 44 g CO₂ = 6.02×10^{23} CO₂ molecules

1 mol SO₂ = 64 g SO₂ = 6.02×10^{23} SO₂ molecules

1 mol C₂H₆ = 30 g C₂H₆ = 6.02×10^{23} C₂H₆ molecules

2 mol C₂H₆ = 60 g C₂H₆ = 2 x 6.02×10^{23} C₂H₆ molecules

Calculations using Avogadro's Number

Calculate the number of CO₂ molecules in 2 g of CO₂.

$$\begin{array}{l}
 44 \text{ g CO}_2 = 6.02 \times 10^{23} \text{ CO}_2 \text{ molecules} \\
 2 \text{ g CO}_2 = x
 \end{array}
 \left. \vphantom{\begin{array}{l} 44 \text{ g CO}_2 = 6.02 \times 10^{23} \text{ CO}_2 \text{ molecules} \\ 2 \text{ g CO}_2 = x \end{array}} \right\}
 \frac{2 \text{ g CO}_2 \times 6.02 \times 10^{23} \text{ molecules}}{44 \text{ g CO}_2}
 = 2.7 \times 10^{22} \text{ CO}_2 \text{ molecules}$$

Calculate the mass of 3.01×10^{24} molecules of CH_4 .

$$\left. \begin{array}{l} 16 \text{ g CH}_4 = 6.02 \times 10^{23} \text{ CH}_4 \text{ molecules} \\ x = 3.01 \times 10^{24} \text{ CH}_4 \text{ molecules} \end{array} \right\} \frac{16 \text{ g CH}_4 \times 3.01 \times 10^{24} \text{ CH}_4 \text{ molecules}}{6.02 \times 10^{23} \text{ CH}_4 \text{ molecules}} \\ = \mathbf{80 \text{ CH}_4 \text{ molecules}}$$

Calculate the mass of 1.2×10^{20} Na atoms.

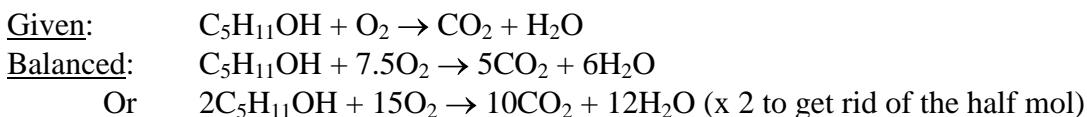
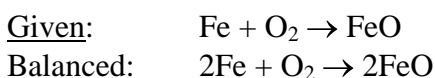
$$\left. \begin{array}{l} 23 \text{ g Na} = 6.02 \times 10^{23} \text{ Na atoms} \\ x = 1.2 \times 10^{20} \text{ Na atoms} \end{array} \right\} \frac{23 \text{ g Na} \times 1.2 \times 10^{20} \text{ Na atoms}}{6.02 \times 10^{23} \text{ Na atoms}} \\ = \mathbf{4.6 \times 10^{-3} \text{ g Na}}$$

Percents of Elements in Compounds

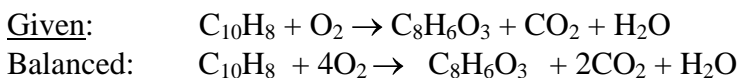
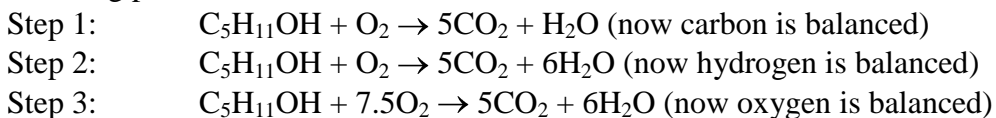
$$\begin{array}{l} \text{C}_2\text{H}_6 \quad \% \text{C} = \frac{2 \times 12}{(2 \times 12) + (6 \times 1)} \times 100 = \frac{24}{30} \times 100 = \mathbf{80 \% C} \\ \quad \quad \% \text{H} = 100 - 80 = \mathbf{20 \% H} \end{array}$$

$$\begin{array}{l} \text{CH}_3\text{OH} \quad \% \text{C} = \frac{12}{32} \times 100 = \mathbf{37.5 \% C} \\ \quad \quad \% \text{O} = \frac{16}{32} \times 100 = \mathbf{50 \% O} \\ \quad \quad \% \text{H} = 100 - (37.5 + 50) = 100 - 87.5 = \mathbf{12.5 \% H} \end{array}$$

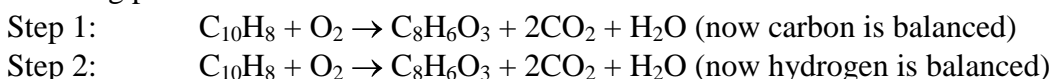
Balancing Equations

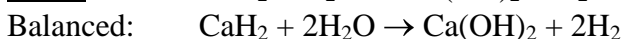
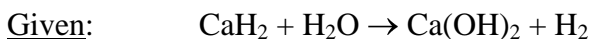
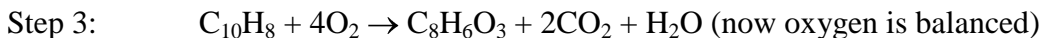


Thinking process:

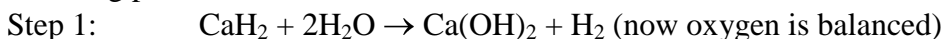


Thinking process:



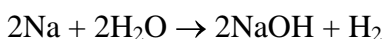


Thinking process:



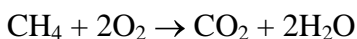
Calculations Using Mol Equivalencies in Equations

ALWAYS BALANCE EQUATIONS FIRST!! (These equations are already balanced.)



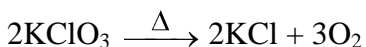
$$2 \text{ mol Na} \approx 2 \text{ mol H}_2\text{O} \approx 2 \text{ mol NaOH} \approx 1 \text{ mol H}_2$$

Calculate the mol O_2 needed to react with .75 mol CH_4 .



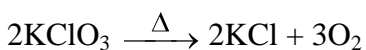
$$\begin{array}{l} 1 \text{ mol CH}_4 \approx 2 \text{ mol O}_2 \\ .75 \text{ mol CH}_4 = x \end{array} \left. \vphantom{\begin{array}{l} 1 \text{ mol CH}_4 \approx 2 \text{ mol O}_2 \\ .75 \text{ mol CH}_4 = x \end{array}} \right\} \frac{.75 \text{ mol CH}_4 \times 2 \text{ mol O}_2}{1 \text{ mol CH}_4} = \mathbf{1.5 \text{ mol O}_2}$$

Calculate the mol O_2 formed from decomposing 2.5 mol $KClO_3$.



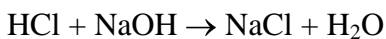
$$\begin{array}{l} 2 \text{ mol KClO}_3 \approx 3 \text{ mol O}_2 \\ 2.5 \text{ mol KClO}_3 = x \end{array} \left. \vphantom{\begin{array}{l} 2 \text{ mol KClO}_3 \approx 3 \text{ mol O}_2 \\ 2.5 \text{ mol KClO}_3 = x \end{array}} \right\} \frac{2.5 \text{ mol KClO}_3 \times 3 \text{ mol O}_2}{2 \text{ mol KClO}_3} = \mathbf{3.75 \text{ mol O}_2}$$

Calculate the g O_2 formed from decomposing 4 g $KClO_3$.



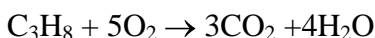
$$\begin{array}{l} 2 \times 122.6 \text{ g KClO}_3 \approx 3 \times 32 \text{ g ox.} \\ 4 \text{ g KClO}_3 = x \end{array} \left. \vphantom{\begin{array}{l} 2 \times 122.6 \text{ g KClO}_3 \approx 3 \times 32 \text{ g ox.} \\ 4 \text{ g KClO}_3 = x \end{array}} \right\} \frac{4 \text{ g KClO}_3 \times 3 \times 32 \text{ g ox.}}{2 \times 122.6 \text{ g KClO}_3} = \mathbf{1.5 \text{ g ox.}}$$

Calculate the g NaOH needed to react with 1.5 g HCl.



$$\begin{array}{l} 36.5 \text{ g HCl} \approx 40 \text{ g NaOH} \\ 1.5 \text{ g HCl} = x \end{array} \left. \vphantom{\begin{array}{l} 36.5 \text{ g HCl} \approx 40 \text{ g NaOH} \\ 1.5 \text{ g HCl} = x \end{array}} \right\} \frac{1.5 \text{ g HCl} \times 40 \text{ g NaOH}}{36.5 \text{ g HCl}} = \mathbf{1.64 \text{ g NaOH}}$$

Calculate the g CO_2 formed from the reaction of .5 mol C_3H_8 in excess of ox.



$$\begin{array}{l} 1 \text{ mol C}_3\text{H}_8 \approx 3 \times 44 \text{ g CO}_2 \\ .5 \text{ mol C}_3\text{H}_8 = x \end{array} \left. \vphantom{\begin{array}{l} 1 \text{ mol C}_3\text{H}_8 \approx 3 \times 44 \text{ g CO}_2 \\ .5 \text{ mol C}_3\text{H}_8 = x \end{array}} \right\} \frac{.5 \text{ mol C}_3\text{H}_8 \times 3 \times 44 \text{ g CO}_2}{1 \text{ mol C}_3\text{H}_8} = \mathbf{66 \text{ g CO}_2}$$

Or

$$\begin{array}{l} 1 \text{ mol C}_3\text{H}_8 \approx 3 \text{ mol CO}_2 \\ .5 \text{ mol C}_3\text{H}_8 = x \end{array} \left. \vphantom{\begin{array}{l} 1 \text{ mol C}_3\text{H}_8 \approx 3 \text{ mol CO}_2 \\ .5 \text{ mol C}_3\text{H}_8 = x \end{array}} \right\} \frac{.5 \text{ mol C}_3\text{H}_8 \times 3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} = 1.5 \text{ mol CO}_2$$

$$\text{mol} = \frac{\text{g}}{\text{mw}}$$

$$\text{g} = \text{mol} \times \text{mw}$$

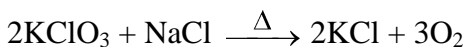
$$= 1.5 \text{ mol CO}_2 \times \frac{44 \text{ g CO}_2}{1 \text{ mol CO}_2} = \mathbf{66 \text{ g CO}_2}$$

The loss in the weight of a test tube containing KClO_3 is .45 g. Calculate the weight of KClO_3 in the test tube.



$$\begin{array}{l} 2 \times 122.6 \text{ g KClO}_3 \approx 3 \times 32 \text{ g ox.} \\ x = .45 \text{ g} \end{array} \left. \vphantom{\begin{array}{l} 2 \times 122.6 \text{ g KClO}_3 \approx 3 \times 32 \text{ g ox.} \\ x = .45 \text{ g} \end{array}} \right\} \frac{2 \times 122.6 \text{ g KClO}_3 \times .45 \text{ g ox.}}{3 \times 32 \text{ g ox.}} = \mathbf{1.2 \text{ g KClO}_3}$$

A test tube contains KClO_3 and NaCl . The weight of the mixture before heating is 3.75 g. After heating the weight of the mixture in the test tube is 3.62 g. Calculate (a) the weight of NaCl in the mixture and (b) the % of KClO_3 in the mixture.

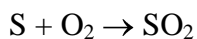


$$(a) \begin{array}{l} 2 \times 122.6 \text{ g KClO}_3 \approx 3 \times 32 \text{ g ox.} \\ x = .13 \text{ g ox.} \end{array} \left. \vphantom{\begin{array}{l} 2 \times 122.6 \text{ g KClO}_3 \approx 3 \times 32 \text{ g ox.} \\ x = .13 \text{ g ox.} \end{array}} \right\} \frac{.13 \text{ g ox.} \times 2 \times 122.6 \text{ g KClO}_3}{3 \times 32 \text{ g ox.}} = \mathbf{.33 \text{ g KClO}_3}$$

(Note: The .13 g ox. is the loss in weight of test tube after heating.)

$$(b) \begin{array}{l} \text{NaCl} + \text{KClO}_3 = 3.75 \text{ g} \\ \text{KClO}_3 = .33 \text{ g} \end{array} \left. \vphantom{\begin{array}{l} \text{NaCl} + \text{KClO}_3 = 3.75 \text{ g} \\ \text{KClO}_3 = .33 \text{ g} \end{array}} \right\} \frac{.33}{3.75} \times 100 = \mathbf{8.8 \% \text{KClO}_3}$$

Calculate the mass of 80% pure S needed to prepare 6 g SO_2 .



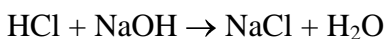
$$\begin{array}{l} 32 \text{ g S} \approx 64 \text{ g SO}_2 \\ x = 6 \text{ g SO}_2 \end{array} \left. \vphantom{\begin{array}{l} 32 \text{ g S} \\ x \end{array}} \right\} \frac{6 \text{ g SO}_2 \times 32 \text{ g S}}{64 \text{ g SO}_2} = 3 \text{ g S}$$

$$\begin{array}{l} 100 \text{ g impure S} = 80 \text{ g pure S} \\ x = 3 \text{ g pure S} \end{array} \left. \vphantom{\begin{array}{l} 100 \text{ g impure S} \\ x \end{array}} \right\} \frac{100 \text{ g impure S} \times 3 \text{ g pure S}}{80 \text{ g pure S}} = 3.75 \text{ g imp. S}$$

Excess and Limiting Reagents in Calculations

(Hint: Problems will give the masses of both the reagents.)

ALWAYS USE LIMITING REAGENT TO CALCULATE WEIGHT OF PRODUCT!!



$$36.5 \text{ g HCl} \approx 40 \text{ g NaOH} \approx 58.5 \text{ g NaCl} \approx 18 \text{ g H}_2\text{O}$$

If 40 g HCl mixed with 40 g NaOH:
HCl not completely consumed – in excess.
NaOH consumed – limiting reagent.

If 2.2 g HCl is mixed with 2 g NaOH (a) Which is in excess? (b) Calculate the weight of NaCl formed in the reaction.

(a) Step 1: Assume 2.2 g HCl was completely consumed to compare the mass of NaOH.

$$\begin{array}{l} 36.5 \text{ g HCl} \approx 40 \text{ g NaOH} \\ 2.2 \text{ g HCl} = x \end{array} \left. \vphantom{\begin{array}{l} 36.5 \text{ g HCl} \\ 2.2 \text{ g HCl} \end{array}} \right\} \frac{40 \text{ g NaOH} \times 2.2 \text{ g HCl}}{36.5 \text{ g HCl}} = 2.4 \text{ g NaOH}$$

Since only 2 g NaOH was given, this is the wrong assumption.

Step 2: Assume 2 g NaOH was consumed.

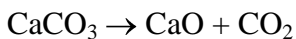
$$\begin{array}{l} 36.5 \text{ g HCl} \approx 40 \text{ g NaOH} \\ x = 2 \text{ g NaOH} \end{array} \left. \vphantom{\begin{array}{l} 36.5 \text{ g HCl} \\ x \end{array}} \right\} \frac{2 \text{ g NaOH} \times 36.5 \text{ g HCl}}{40 \text{ g NaOH}} = 1.83 \text{ g HCl}$$

Therefore, **HCl is in excess by .37 g** (2.2 – 1.83).

$$\begin{array}{l} 40 \text{ g NaOH} \approx 58.5 \text{ g NaCl} \\ 2 \text{ g NaOH} = x \\ \text{Or} \end{array} \left. \vphantom{\begin{array}{l} 40 \text{ g NaOH} \\ 2 \text{ g NaOH} \end{array}} \right\} \frac{2 \text{ g NaOH} \times 58.5 \text{ g NaCl}}{40 \text{ g NaOH}} = 2.93 \text{ g NaCl}$$

$$\begin{array}{l} 36.5\text{g HCl} \approx 58.5\text{g NaCl} \\ 1.83\text{g HCl} = x \end{array} \left. \vphantom{\begin{array}{l} 36.5\text{g HCl} \\ 1.83\text{g HCl} \end{array}} \right\} \frac{58.5\text{g NaCl} \times 1.83\text{g HCl}}{36.5\text{g HCl}} = \mathbf{2.93\text{ g NaCl}}$$

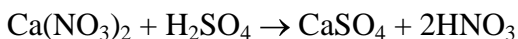
Decomposing 5 g CaCO₃ will give 2.2 g CaO. Calculate the % yield of the reaction.



$$\begin{array}{l} 100\text{g CaCO}_3 \approx 56\text{g CaO} \\ 5\text{g CaCO}_3 = x \end{array} \left. \vphantom{\begin{array}{l} 100\text{g CaCO}_3 \\ 5\text{g CaCO}_3 \end{array}} \right\} \frac{5\text{g CaCO}_3 \times 56\text{g CaO}}{100\text{g CaCO}_3} = 2.8\text{ g CaO}$$

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} = \frac{2.2}{2.8} \times 100 = \mathbf{78.8\%}$$

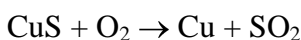
Calculate the weight of CaSO₄ formed from the reaction of 10 g H₂SO₄ if % yield is 70%.



$$\begin{array}{l} 98\text{g H}_2\text{SO}_4 \approx 136\text{g CaSO}_4 \\ 10\text{g H}_2\text{SO}_4 = x \end{array} \left. \vphantom{\begin{array}{l} 98\text{g H}_2\text{SO}_4 \\ 10\text{g H}_2\text{SO}_4 \end{array}} \right\} \frac{10\text{g H}_2\text{SO}_4 \times 136\text{g CaSO}_4}{98\text{g H}_2\text{SO}_4} = 13.8\text{ g CaSO}_4$$

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \quad .7 = \frac{\text{actual yield}}{13.8} \quad .7 \times 13.8 = \text{actual yield} = \mathbf{9.7\text{ g CaSO}_4}$$

An ore of Cu contains 80 % CuS. Calculate the tons of Cu produced from 100,000 tons of ore if the yield of the process is 60%.



$$\begin{array}{l} 100\text{ tons ore} \approx 80\text{ tons CuS} \\ 100,000\text{ tons ore} = x \end{array} \left. \vphantom{\begin{array}{l} 100\text{ tons ore} \\ 100,000\text{ tons ore} \end{array}} \right\} \frac{100,000\text{ tons ore} \times 80\text{ tons CuS}}{100\text{ tons ore}} = 80,000\text{ tons CuS}$$

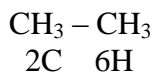
$$\begin{array}{l} 96\text{ tons CuS} \approx 64\text{ tons Cu} \\ 80,000\text{ tons CuS} = x \end{array} \left. \vphantom{\begin{array}{l} 96\text{ tons CuS} \\ 80,000\text{ tons CuS} \end{array}} \right\} \frac{80,000\text{ tons CuS} \times 64\text{ tons Cu}}{96\text{ tons CuS}} = 53,333\text{ tons Cu}$$

$$\text{actual yield} = \% \text{ yield} \times \text{theoretical yield} \quad \frac{60}{100} \times 53,333 = \mathbf{31,999.8\text{ tons Cu}}$$

Simple/Empirical Formula

Shows the lowest ratio of elements in a compound.

$$x(\text{mass of S.F}) = \text{mw}$$



E.F. Molecular Formula

$$x(\text{CH}_3) = \text{mw}$$

$$x(\text{CH}_3) = 30$$

$$15x = 30$$

$$x = \frac{30}{15} = 2$$

$$\frac{\% \text{ of each element}}{\text{at. mass of the element}} = x$$

Ex. $\frac{\% \text{ A}}{\text{at. mass A}} = .25$

$$\frac{\% \text{ B}}{\text{at. mass B}} = .5$$

Divide
both by
smallest

$$= \frac{.25}{.25} = 1$$

$$= \frac{.5}{.25} = 2$$

So E.F.
is
AB₂

A 5 g sample of an oxide of lead contains 4.53 g Pb. Determine the Empirical Formula of the compound.

$$\frac{4.53 \text{ g Pb}}{207} = \frac{.022}{.02} = 1$$

$$\frac{.47 \text{ g ox. (5 - 4.53)}}{16} = \frac{.03}{.02} = 1$$

So E.F. is **PbO**