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 (H_2) = 2 g (1 x 2) = 6.02×10^{23} hydrogen molecules

1 mol of oxygen (O_2) = 32 g (16 x 2) = 6.02×10^{23} oxygen molecules

1 mol of chlorine (Cl_2) = 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 CO_2 = 44 g CO_2 = 6.02×10^{23} CO_2 molecules

1 mol SO_2 = 64 g SO_2 = 6.02×10^{23} SO_2 molecules

1 mol C_2H_6 = 30 g C_2H_6 = 6.02×10^{23} C_2H_6 molecules

2 mol C_2H_6 = 60 g C_2H_6 = 2 x 6.02×10^{23} C_2H_6 molecules

Calculations using Avogadro's Number

Calculate the number of CO_2 molecules in 2 g of CO_2.

$$\left.\begin{array}{l} 44\,\mathrm{g\,CO_2} = 6.02 \times 10^{23}\ \mathrm{CO_2\ molecules} \\ 2\,\mathrm{g\,CO_2} = x \end{array}\right\} \frac{2\,\mathrm{g\,CO_2} \times 6.02 \times 10^{23}\ \mathrm{molecules}}{44\,\mathrm{g\,CO_2}}$$

= **2.7×10^{22} CO_2 molecules**

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

$$\left.\begin{aligned} 16\,g\,CH_4 &= 6.02\,x\,10^{23}\ CH_4\ \text{molecules} \\ x &= 3.01\,x\,10^{24}\ CH_4\ \text{molecules} \end{aligned}\right\} \frac{16\,g\,CH_4\ x\,3.01\,x\,10^{24}\ CH_4\ \text{molecules}}{6.02\,x\,10^{23}\ CH_4\ \text{molecules}}$$

= **80 CH_4 molecules**

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

$$\left.\begin{aligned} 23\,g\,Na &= 6.02\,x\,10^{23}\ Na\ \text{atoms} \\ x &= 1.2\,x\,10^{20}\ Na\ \text{atoms} \end{aligned}\right\} \frac{23\,g\,Na\ x\,1.2\,x\,10^{20}\ Na\ \text{atoms}}{6.02\,x\,10^{23}\ Na\ \text{atoms}}$$

= **4.6 x 10^{-3} g Na**

Percents of Elements in Compounds

C_2H_6 $\quad$ %C = $\frac{2\,x\,12}{(2\,x\,12)+(6\,x\,1)}$ x 100 = $\frac{24}{30}$ x 100 = **80 %C**

%H = 100 – 80 = **20 %H**

CH_3OH $\quad$ %C = $\frac{12}{32}$ x 100 = **37.5 %C**

%O = $\frac{16}{32}$ x 100 = **50 %O**

%H = 100 – (37.5 +50) = 100 – 87.5 = **12.5 %H**

Balancing Equations

Given: $Fe + O_2 \rightarrow FeO$
Balanced: $2Fe + O_2 \rightarrow 2FeO$

Given: $C_5H_{11}OH + O_2 \rightarrow CO_2 + H_2O$
Balanced: $C_5H_{11}OH + 7.5O_2 \rightarrow 5CO_2 + 6H_2O$
Or $2C_5H_{11}OH + 15O_2 \rightarrow 10CO_2 + 12H_2O$ (x 2 to get rid of the half mol)

Thinking process:
Step 1: $C_5H_{11}OH + O_2 \rightarrow 5CO_2 + H_2O$ (now carbon is balanced)
Step 2: $C_5H_{11}OH + O_2 \rightarrow 5CO_2 + 6H_2O$ (now hydrogen is balanced)
Step 3: $C_5H_{11}OH + 7.5O_2 \rightarrow 5CO_2 + 6H_2O$ (now oxygen is balanced)

Given: $C_{10}H_8 + O_2 \rightarrow C_8H_6O_3 + CO_2 + H_2O$
Balanced: $C_{10}H_8 + 4O_2 \rightarrow C_8H_6O_3 + 2CO_2 + H_2O$

Thinking process:
Step 1: $C_{10}H_8 + O_2 \rightarrow C_8H_6O_3 + 2CO_2 + H_2O$ (now carbon is balanced)
Step 2: $C_{10}H_8 + O_2 \rightarrow C_8H_6O_3 + 2CO_2 + H_2O$ (now hydrogen is balanced)

Step 3: $C_{10}H_8 + 4O_2 \rightarrow C_8H_6O_3 + 2CO_2 + H_2O$ (now oxygen is balanced)

<u>Given</u>: $CaH_2 + H_2O \rightarrow Ca(OH)_2 + H_2$
<u>Balanced</u>: $CaH_2 + 2H_2O \rightarrow Ca(OH)_2 + 2H_2$

Thinking process:
Step 1: $CaH_2 + 2H_2O \rightarrow Ca(OH)_2 + H_2$ (now oxygen is balanced)
Step 2: $CaH_2 + 2H_2O \rightarrow Ca(OH)_2 + 2H_2$ (now hydrogen is balanced)

Calculations Using Mol Equivalencies in Equations

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

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

2 mol Na $\approx$ 2 mol H_2O $\approx$ 2 mol NaOH $\approx$ 1 mol H_2

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

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$$\left.\begin{array}{r} 1\,\text{mol}\,CH_4 \approx 2\,\text{mol}\,O_2 \\ .75\,\text{mol}\,CH_4 = x \end{array}\right\} \quad \frac{.75\,\text{mol}\,CH_4 \times 2\,\text{mol}\,O_2}{1\,\text{mol}\,CH_4} = \mathbf{1.5\ mol\ O_2}$$

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

$2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$

$$\left.\begin{array}{r} 2\,\text{mol}\,KClO_3 \approx 3\,\text{mol}\,O_2 \\ 2.5\,\text{mol}\,KClO_3 = x \end{array}\right\} \quad \frac{2.5\,\text{mol}\,KClO_3 \times 3\,\text{mol}\,O_2}{2\,\text{mol}\,KClO_3} = \mathbf{3.75\ mol\ O_2}$$

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

$2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$

$$\left.\begin{array}{r} 2 \times 122.6\,\text{g}\,KClO_3 \approx 3 \times 32\,\text{g ox.} \\ 4\,\text{g}\,KClO_3 = x \end{array}\right\} \quad \frac{4\,\text{g}\,KClO_3 \times 3 \times 32\,\text{g ox.}}{2 \times 122.6\,\text{g}\,KClO_3} = \mathbf{1.5\ g\ ox.}$$

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

$HCl + NaOH \rightarrow NaCl + H_2O$

$$\left.\begin{array}{r} 36.5\,g\ HCl \approx 40\,g\ NaOH \\ 1.5\,g\ HCl = x \end{array}\right\} \quad \frac{1.5\,g\ HCl \times 40\,g\ NaOH}{36.5\,g\ HCl} = \mathbf{1.64\ g\ NaOH}$$

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

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

$$\left.\begin{array}{r} 1\,mol\,C_3H_8 \approx 3 \times 44\,g\,CO_2 \\ .5\,mol\,C_3H_8 = x \end{array}\right\} \quad \frac{.5\,mol\,C_3H_8 \times 3 \times 44\,g\,CO_2}{1\,mol\,C_3H_8} = \mathbf{66\ g\ CO_2}$$

Or

$$\left.\begin{array}{r} 1\,mol\,C_3H_8 \approx 3\,mol\,CO_2 \\ .5\,mol\,C_3H_8 = x \end{array}\right\} \quad \frac{.5\,mol\,C_3H_8 \times 3\,mol\,CO_2}{1\,mol\,C_3H_8} = 1.5\ mol\ CO_2$$

$$mol = \frac{g}{mw} \qquad g = mol \times mw$$

$$= 1.5\ mol\ CO_2 \times \frac{44\,g\,CO_2}{1\,mol\,CO_2} = \mathbf{66\ g\ CO_2}$$

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.

$2KClO_{3\,(s)} \rightarrow 2KCl_{\,(s)} + 3O_{2\,(g)}$

$$\left.\begin{array}{r} 2 \times 122.6\,g\,KClO_3 \approx 3 \times 32\,g\,ox. \\ x = .45\,g \end{array}\right\} \quad \frac{2 \times 122.6\,g\,KClO_3 \times .45\,g\,ox.}{3 \times 32\,g\,ox.} = \mathbf{1.2\ g\ KClO_3}$$

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.

$2KClO_3 + NaCl \xrightarrow{\Delta} 2KCl + 3O_2$

(a) $$\left.\begin{array}{r} 2 \times 122.6\,g\,KClO_3 \approx 3 \times 32\,g\,ox. \\ x = .13\,g\,ox. \end{array}\right\} \quad \frac{.13\,g\,ox. \times 2 \times 122.6\,g\,KClO_3}{3 \times 32\,g\,ox.} = \mathbf{.33\ g\ KClO_3}$$

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

(b) $$\left.\begin{array}{r} NaCl + KClO_3 = 3.75\,g \\ KClO_3 = .33\,g \end{array}\right\} \quad \frac{.33}{3.75} \times 100 = \mathbf{8.8\ \%KClO_3}$$

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

$S + O_2 \rightarrow SO_2$

$$\left.\begin{array}{r} 32\,\text{g S} \approx 64\,\text{g } SO_2 \\ x = 6\,\text{g } SO_2 \end{array}\right\} \quad \frac{6\,\text{g } SO_2 \text{ x } 32\,\text{g S}}{64\,\text{g } SO_2} = 3\text{ g S}$$

$$\left.\begin{array}{r} 100\,\text{g impure S} = 80\,\text{g pure S} \\ x = 3\,\text{g pure S} \end{array}\right\} \quad \frac{100\,\text{g impure S x } 3\,\text{g pure S}}{80\,\text{g pure S}} = \mathbf{3.75\text{ g imp. S}}$$

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!!

$HCl + NaOH \rightarrow NaCl + H_2O$
36.5 g HCl $\approx$ 40 g NaOH $\approx$ 58.5 g NaCl $\approx$ 18 g H_2O

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.

$$\left.\begin{array}{r} 36.5\,\text{g HCl} \approx 40\,\text{g NaOH} \\ 2.2\,\text{g HCl} = x \end{array}\right\} \quad \frac{40\,\text{g NaOH x } 2.2\,\text{g HCl}}{36.5\,\text{g HCl}} = 2.4\text{ g NaOH}$$

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

Step 2: Assume 2 g NaOH was consumed.

$$\left.\begin{array}{r} 36.5\,\text{g HCl} \approx 40\,\text{g NaOH} \\ x = 2\,\text{g NaOH} \end{array}\right\} \quad \frac{2\,\text{g NaOH x } 36.5\,\text{g HCl}}{40\,\text{g NaOH}} = 1.83\text{ g HCl}$$

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

(b)
$$\left.\begin{array}{r} 40\,\text{g NaOH} \approx 58.5\,\text{g NaCl} \\ 2\,\text{g NaOH} = x \end{array}\right\} \quad \frac{2\,\text{g NaOH x } 58.5\,\text{g NaCl}}{40\,\text{g NaOH}} = \mathbf{2.93\text{ g NaCl}}$$

Or

$$36.5\,g\,HCl \approx 58.5\,g\,NaCl$$
$$1.83\,g\,HCl = x$$
$$\frac{58.5\,g\,NaCl \times 1.83\,g\,HCl}{36.5\,g\,HCl} = \mathbf{2.93\ g\ NaCl}$$

Decomposing 5 g $CaCO_3$ will give 2.2 g CaO. Calculate the % yield of the reaction.

$CaCO_3 \rightarrow CaO + CO_2$

$$100\,g\,CaCO_3 \approx 56\,g\,CaO$$
$$5\,g\,CaCO_3 = x$$
$$\frac{5\,g\,CaCO_3 \times 56\,g\,CaO}{100\,g\,CaCO_3} = 2.8\ g\ CaO$$

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

Calculate the weight of $CaSO_4$ formed from the reaction of 10 g H_2SO_4 if % yield is 70%.

$Ca(NO_3)_2 + H_2SO_4 \rightarrow CaSO_4 + 2HNO_3$

$$98\,g\,H_2SO_4 \approx 136\,g\,CaSO_4$$
$$10\,g\,H_2SO_4 = x$$
$$\frac{10\,g\,H_2SO_4 \times 136\,g\,CaSO_4}{98\,g\,H_2SO_4} = 13.8\ g\ CaSO_4$$

$$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \qquad .7 = \frac{actual\ yield}{13.8} \qquad .7 \times 13.8 = actual\ yield = \mathbf{9.7\ g\ CaSO_4}$$

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%.

$CuS + O_2 \rightarrow Cu + SO_2$

$$100\ tons\ ore \approx 80\ tons\ CuS$$
$$100{,}000\ tons\ ore = x$$
$$\frac{100{,}000\ tons\ ore \times 80\ tons\ CuS}{100\ tons\ ore} = 80{,}000\ tons\ CuS$$

$$96\ tons\ CuS \approx 64\ tons\ Cu$$
$$80{,}000\ tons\ CuS = x$$
$$\frac{80{,}000\ tons\ CuS \times 64\ tons\ Cu}{96\ tons\ CuS} = 53{,}333\ tons\ Cu$$

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

Simple/Empirical Formula
Shows the lowest ratio of elements in a compound.

x(mass of S.F) = mw

$CH_3 - CH_3$

2C 6H

$$\underset{\text{E.F}}{x(CH_3)} = \underset{\text{Molecular Formula}}{mw}$$

$$x(CH_3) = 30$$

$$15x = 30$$

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

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

Ex.

$$\left.\begin{aligned} \frac{\%A}{\text{at. mass}A} &= .25 \\ \frac{\%B}{\text{at. mass}B} &= .5 \end{aligned}\right\} \begin{array}{c}\text{Divide}\\ \text{both by}\\ \text{smallest}\end{array} \qquad \left.\begin{aligned} &= \frac{.25}{.25} = 1 \\ &= \frac{.5}{.25} = 2 \end{aligned}\right\} \begin{array}{c}\text{So E.F}\\ \text{is}\\ \mathbf{AB_2}\end{array}$$

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

$$\left.\begin{aligned} \frac{4.53\text{g Pb}}{207} &= \frac{.022}{.02} = 1 \\ \frac{.47\text{g ox.}(5-4.53)}{16} &= \frac{.03}{.02} = 1 \end{aligned}\right\} \text{So E.F. is } \mathbf{PbO}$$