Unit 7 Stoichiometry

Stoichiometry- calculating quantities in a chemical reaction.

$$\mathbf{H_2\,(g) + I_2\,(g) \rightarrow 2HI\,(g)}$$

particles (atoms)	2	+	2	=	4
mass (molar)	2g	+	254g	=	256g
molar volume	22.4L	+	22.4L	=	44.8L
moles	1	+	1	=	2

$$\mathbf{N_2(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)}$$

particles (atoms)	2	+	6	=	8
mass (molar)	28g	+	6g	=	34g
molar volume	22.4L	+	67.2L	≠	44.8L
moles	1	+	3	≠	2

Note: only mass and particles (number of atoms) are conserved in ALL reactions.

$$\mathbf{H_2\,(g) + I_2\,(g) \rightarrow 2HI\,(g)}$$

moles 1 + 1 = 2

mole ratios $\dfrac{1 \text{ mol } H_2}{1 \text{ mol } I_2}$ or $\dfrac{2 \text{ mol HI}}{1 \text{ mol } H_2}$ or $\dfrac{1 \text{ mol } I_2}{2 \text{ mol HI}}$

$$\mathbf{N_2(g) + 3H_2\ (g) \rightarrow 2NH_3\ (g)}$$

moles $\quad 1 \quad + \quad 3 \quad \neq \quad 2$

mole ratios $\quad \dfrac{1\ \text{mol}\ N_2}{3\ \text{mol}\ H_2}$ or $\dfrac{2\ \text{mol}\ NH_3}{1\ \text{mol}\ N_2}$ or $\dfrac{3\ \text{mol}\ H_2}{2\ \text{mol}\ NH_3}$

Problem:
How many moles of H_2 are needed to produce 6 moles of NH_3 ?

$$\frac{6\ \text{mol}\ NH_3}{} \ \text{X}\ \frac{3\ \text{mol}\ H_2}{2\ \text{mol}\ NH_3} = 9\ \text{mol}\ H_2$$

Unit 7 Stoichiometry More sample problems

Example 1. How many moles of ammonia are produced when 0.6 moles of nitrogen reacts with hydrogen?

$$\mathbf{N_2(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)}$$

$$\frac{0.6\ \cancel{\text{mol } N_2}}{} \times \frac{2\ \text{mol } NH_3}{1\ \cancel{\text{mol } N_2}} = 1.2\ \text{mol } NH_3$$

Example 2. How many moles of aluminum are needed to form 3.7 moles of aluminum trioxide?

$$\mathbf{4Al\,(s) + 3O_2\,(g) \rightarrow 2Al_2O_3\,(s)}$$

$$\frac{3.7\ \cancel{\text{mol } Al_2O_3}}{} \times \frac{4\ \text{mol Al}}{2\ \cancel{\text{mol } Al_2O_3}} = 7.4\ \text{mol Al}$$

Example 3. Calculate the number of grams of ammonia produced by the reaction of 5.4 g of hydrogen with an excess of nitrogen.

$$\mathbf{N_2(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)}$$

Step 1. Convert grams to moles.

$$\frac{5.4\ \cancel{\text{g } H_2}}{} \times \frac{1\ \text{mol } H_2}{2\ \cancel{\text{gram } H_2}} = 2.7\ \text{moles } H_2$$

Step 2. Convert moles reactant to moles product.

$$\frac{2.7\ \cancel{\text{mol } H_2}}{} \times \frac{2\ \text{mol } NH_3}{3\ \cancel{\text{mol } H_2}} = 1.8\ \text{mol } NH_3$$

Step 3. Convert moles back to grams.

$$\frac{1.8\ \cancel{\text{mol } NH_3}}{} \times \frac{17\ \text{g } NH_3}{1\ \cancel{\text{mol } NH_3}} = 31\ \text{g } NH_3$$

Example 4. How many molecules of oxygen are produced when a sample of 29.2g of water is decomposed by electrolysis?

$$2H_2O\ (l) \rightarrow 2H_2\ (g) + O_2\ (g)$$

Step 1. Convert grams to moles.

$$\frac{29.2\ \cancel{g\ H_2O}}{} \times \frac{1\ \text{mole}\ H_2O}{18\ \cancel{g\ H_2O}} = 1.62\ \text{mol}\ H_2O$$

Step 2. Convert moles reactant to moles product.

$$\frac{1.62\ \cancel{mol\ H_2O}}{} \times \frac{1\ \text{mole}\ O_2}{2\ \cancel{mol\ H_2O}} = 0.81\ \text{mol}\ O_2$$

Step 3. Convert moles product to molecules product.

$$\frac{0.81\ \cancel{mol\ O_2}}{} \times \frac{6.02 \times 10^{23}\ \text{molecules}\ O_2}{1\ \cancel{mol\ O_2}} = 4.9 \times 10^{23}\text{molecules}\ O_2$$

Example 5. Assuming STP, how many liters of oxygen are needed to produce 19.8 liters SO_3 ?

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

Step 1. Convert volume to moles

$$\frac{19.8\ l\ SO_3}{} \times \frac{1\ \text{mol}}{22.4\ l} = 0.884\ \text{mol}\ SO_3$$

Step 2. Convert moles product to moles reactant

$$\frac{0.884\ \cancel{mol\ SO_3}}{} \times \frac{1\ \text{mol}\ O_2}{2\ \cancel{mol\ SO_3}} = 0.442\ \text{mol}\ O_2$$

Step 3. Convert moles to volume

$$\frac{0.442\ \cancel{mol}\ O_2}{} \times \frac{22.4\ l}{1\ \cancel{mol}} = 9.9\ l\ O_2$$

Example Problems done in one line/one step

Example 3.
Calculate the number of grams of ammonia produced by the reaction of 5.4 g of hydrogen with an excess of nitrogen.

$$\mathbf{N_2(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)}$$

$$\frac{5.4\text{ g }H_2}{} \times \frac{1\text{ mol }H_2}{2\text{ gram }H_2} \times \frac{2\text{ mol }NH_3}{3\text{ mol }H_2} \times \frac{17\text{ g }NH_3}{1\text{ mol }NH_3} = 31\text{ g }NH_3$$

Example 4.
How many molecules of oxygen are produced when a sample of 29.2 g of water is decomposed by electrolysis?

$$\mathbf{2H_2O\,(\mathit{l}) \rightarrow 2H_2\,(\mathit{g}) + O_2\,(\mathit{g})}$$

$$\frac{29.2\text{g }H_2O}{} \times \frac{1\text{ mole }H_2O}{18\text{ g }H_2O} \times \frac{1\text{ mole }O_2}{2\text{ mol }H_2O} \times \frac{6.02\text{x}10^{23}\text{ molecules }O_2}{1\text{ mol }O_2} = 4.9\text{x}10^{23}\text{ molecules }O_2$$

Example 5.
Assuming STP, how many liters of oxygen are needed to produce 19.8 liters SO_3 ?

$$\mathbf{2SO_2\,(\mathit{g}) + O_2\,(\mathit{g}) \rightarrow 2SO_3\,(\mathit{g})}$$

$$\frac{19.8\ l\ SO_3}{} \times \frac{1\text{ mol}}{22.4\ l} \times \frac{1\text{ mol }O_2}{2\text{ mol }SO_3} \times \frac{22.4\ l}{1\text{ mol}} = 9.9\ l\ O_2$$

Limiting Reagent and Percent Yield

Theoretical yield- amount of product (maximum) that could be formed in a reaction.

Actual yield- how much product is actually produced.

Percent yield- how much product is produced compared to how much was expected.

$$\textbf{Percent Yield} = \frac{\textbf{(Actual yield)}}{\textbf{(Theoretical yield)}} \ \textbf{x 100}$$

Limiting Reagent- the first substance used up in an experiment; i.e. completely reacted.

Excess reagent- the substance left over in an experiment, i.e. did not completely react.

Example Problem: If we start with 6.70 mol Na and 3.20 mol Cl_2 what is the limiting reagent?

$$2Na\ (s) + Cl_2 \rightarrow 2\ NaCl$$

$$\frac{6.70 \text{ mol Na}}{} \ \text{X} \ \frac{1 \text{ mol } Cl_2}{2 \text{ mol Na}} = 3.35 \text{ mol } Cl_2$$

Since 3.35 mol Cl_2 is required to react with 6.70 mol Na, Cl_2 is the limiting reagent and Na is the excess reagent.

$$\frac{3.20 \text{ mol } Cl_2}{} \ \text{X} \ \frac{2 \text{ mol Na}}{1 \text{ mol } Cl_2} = 6.40 \text{ mol Na}$$

Since Cl_2 is limiting, the theoretical yield is 6.40 mol NaCl = 374.4g.

If only 337g was produced, what is the percent yield?

$$\textbf{Percent Yield} = \frac{\textbf{337g}}{\textbf{374.4g}} \ \textbf{x 100 = 90\%}$$