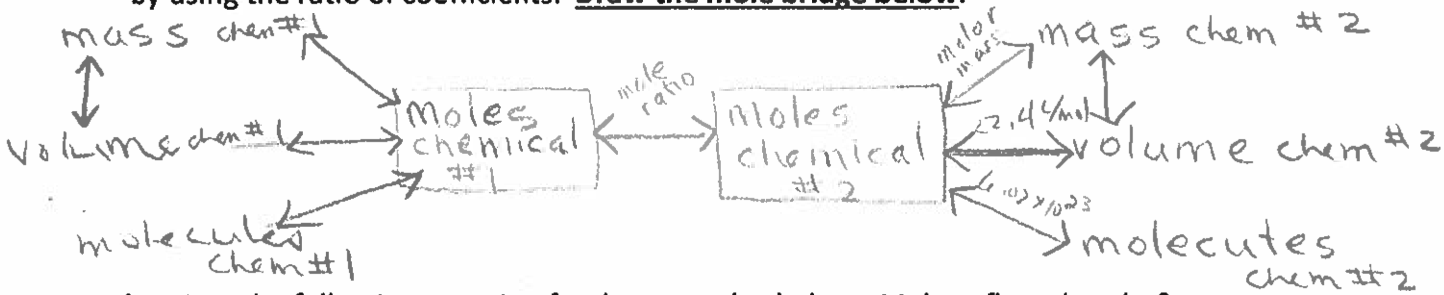
KEY

Stoichiometry - the relationship between the amount of reactants used and the amount of products produced in a chemical reaction.

Coefficients in a balanced reaction show the mole ratio between reactants and products

- Sometimes we say that a "mole bridge" connects any chemicals in a reaction. If you know the moles of one chemical, you can find the moles of any of the other chemicals by using the ratio of coefficients. **Draw the mole bridge below**:



Examples: Use the following equation for the examples below. Make a flow chart before starting your calculations.

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

$12+2(16)=$

1. What mass of CO_2 is produced by reacting 2.00 moles of O_2 ? mol O_2 → mol CO_2 → g CO_2

$$2.00\ \cancel{mol\ O_2} \times \frac{3\ \cancel{mol\ CO_2}}{5\ \cancel{mol\ O_2}} \times \frac{44.0\ g\ CO_2}{1\ \cancel{mol\ CO_2}} = 52.8\ g\ CO_2$$

$3(12)+8=$

2. What mass of C_3H_8 is required to produce 100.0g of H_2O? mass H_2O → mol H_2O → mol C_3H_8 → g

$$100.0\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.0\ g\ H_2O} \times \frac{1\ mol\ C_3H_8}{4\ mol\ H_2O} \times \frac{44.0\ g\ C_3H_8}{1\ mol\ C_3H_8} = 61.1\ g\ C_3$$

3. If a sample of propane is burned what mass of $H_2O_{(l)}$ is produced if the reaction also produced 50.0L of $CO_{2(g)}$ at STP? L CO_2 → mol CO_2 → mol H_2O → g H_2O

$$50.0\ L\ CO_2 \times \frac{1\ mol\ CO_2}{22.4\ L\ CO_2} \times \frac{4\ mol\ H_2O}{3\ mol\ CO_2} \times \frac{18.0\ g\ H_2O}{1\ mol\ H_2O} = 53.6\ g\ H_2O$$

4. A sample of porous, gas-bearing rock is crushed and 1.35×10^{-6}g of C_3H_8 is extracted from the powdered rock. How many molecules of CO_2 are produced if the gas sample is burned in the presence of an excess of O_2? g C_3H_8 → mol C_3H_8 → mol O_2 → molec O_2

$$1.35\times10^{-6}\ g\ C_3H_8 \times \frac{1\ mol\ C_3H_8}{44.0\ g} \times \frac{3\ mol\ CO_2}{1\ mol\ C_3H_8} \times \frac{6.02\times10^{23}\ molec\ CO_2}{1\ mol} = 5.54\times10^{16}\ molec.\ CO_2$$

Try #6-8 on page 127 in Hebden

% Yield in Chemical Reactions

Sometimes 100% of the expected amount of products cannot be obtained from a reaction. Percent yield is the amount of product actually obtained as a percentage of the expected amount.

Possible Reasons:

1. Reactants may not all react
2. Some of the products may be lost during solvent extraction, filtration or evaporation.

Formula:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Actual yield = mass of product actually obtained in a lab. (measured on a scale)

Theoretical yield = mass of product that is expected based on amount of reactant that was consumed. (calculated using stoichiometry)

Example:

$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$

meaning this will be left over.

35.5
12
3
50.5

When 15.0g of CH_4 is reacted with an excess of Cl_2 a total of 29.7g of CH_3Cl is formed. What is the percent yield?

experimental

Step1- Find the mass of CH_3Cl expected. g CH_4 → mol CH_4 → mol CH_3Cl → g CH_3Cl

$$15.0\text{g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4} \times \frac{1 \text{ mol } CH_3Cl}{1 \text{ mol } CH_4} \times \frac{50.5 \text{ g } CH_3Cl}{1 \text{ mol } CH_3Cl} = 47.3\text{g}$$

theoretical

Step 2 – Find the % Yield of CH_3Cl.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{29.7\text{g}}{47.3\text{g}} \times 100\% = 63\% \text{ Yield}$$

fix.

can be over 100% or under

c n V

Stoichiometry Calculations involving Molarity

$V = \frac{n}{c}$

Molar concentration calculations use the formula: $c = n/V$ or $n = c \times V$ or ~~$V = n \times c$~~

Examples:

Tums is an antacid primarily composed of calcium carbonate, and stomach acid is a dilute form of Hydrochloric acid. They react is the following equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

1. A tablet of Tums has a mass of 0.750g. What volume of stomach acid having [HCl] = 0.0010M does it neutralize? $40.1 + 12 + 3(16.0) = 100.1$

   Plan: *mass CaCO3 → moles CaCO3 → moles HCl*

   ..... then use v= n/c to find volume

   $$0.750\,g \times \frac{1\,mol}{100.1\,g} \times \frac{2\,mol\ HCl}{1\,mol\ CaCO_3} = 0.0150\ mol\ HCl$$

   $$V = \frac{n}{c} = \frac{0.0150\,mol}{0.0010\,M} = 15\,L$$

2. What volume of $CO_2(g)$ at STP is produced if 1.25L of 0.0055 M HCl reacts with an excess of $CaCO_3$?

   Plan: use n=c x V to find moles of HCl

   .... then mol HCl → mol CO_2 → Volume ~~HCl~~ CO_2

   $$n = c \times V = 0.0055\,M \times 1.25\,L = 0.0069\ mol\ HCl$$

   $$0.0069\ mol\ HCl \times \frac{1\,mol\ CO_2}{2\,mol\ HCl} \times \frac{22.4\,L}{1\,mol} = 0.077\,L$$

   or 77 mL CO_2

Acid Base Titrations: Use an equivalence point to determine an unknown concentration

Example:

HCl (aq) +	NaOH (aq)	→	NaCl(aq) + H_2O
? M	0.5 M		
10.0 mL	0.90 mL (avg)		
0.0100 L	0.00090 L		

Steps to find the unknow molarity of HCl:

1. Find moles of NaOH (n = c x v)

   $$n = 0.5\,\frac{mol}{L} \times 0.00090\,L = 0.00045\ mol\ NaOH$$

2. Find moles of HCl (multiply by ratio 1:1)

   $$0.00045\ mol\ NaOH \times \frac{1\,mol\ HCl}{1\,mol\ NaOH} = 0.00045\ mol\ HCl$$

3. Find [HCl] (c = n/v)

   $$c = \frac{0.00045\,mol}{0.0100\,L} = 0.045\,M\ HCl$$

$\boxed{\text{Try \#17-20 on page 131 Hebden}}$

Excess and Limiting Reagents

A reaction can only produce as much product as is allowed by the amount of reactants used.

Excess reagent = the reactant that will have some remaining when the reaction is complete

Limiting reagent = the reactant that is completely used up during the chemical reaction. The mass of this reactant must be used when calculating the theoretical mass of products!

$Ba(OH)_2 = 137.3 + 2(16.0) + 2(1.0)$
$= 171.3$

Example problems are based on the following reaction:

$3Ba(OH)_2 + 2AlCl_3 \rightarrow 3BaCl_2 + 2Al(OH)_2$

$BaCl_2 = 137.3 + 2(35.5)$

$AlCl_3 = \cancel{27}.0 + 3(35.5)$
$= 133.5$

1. If 200.0g of $Ba(OH)_2$ are combined with 451.0g of $AlCl_3$ which reactant is in excess?

must choose the same product.

STEP 1 – Using 200.0g of $Ba(OH)_2$ find the ~~mass of a product (or~~ mol)

$200.0g \times \frac{1}{171.3} \times \frac{3}{3} = 1.168 \text{ mol } BaCl_2$

STEP 2 – Using 451.0g of $AlCl_3$ find the ~~mass of a product~~ mol

$451.0 \times \frac{1}{133.5} \times \frac{3}{2} = 5.067 \text{ mol } BaCl_2$

STEP 3 – The reactant that leads to the higher mass is in excess!

$AlCl_3$ is in excess!
$Ba(OH)_2$ is limiting!

2. How much of the excess reagent is leftover when the reaction is complete?

STEP 1 – Starting with the limiting reagent, find the mass of excess (other) reagent.

$200.0g \times \frac{1}{171.3} \times \frac{2}{3} \times \frac{133.5}{1} = 103.9g\ AlCl_3$ actually react

STEP 2 – Subtract this amount from the amount of excess reagent available.

$451.0 - 103.9 = 347.1g\ AlCl_3$ is left over.

3. If there is a 70% yield of BaCl2, what was the actual, experimental yield?

STEP 1 – Find the theoretical mass of BaCl2 (using limiting reagent)

$1.168 \text{ mol } BaCl_2 \times \frac{208.3g}{1 \text{ mol}} = 243.3g$

STEP 2 – Use the equation: Experimental / Theoretical x 100 Solve for Experimental!

$70\% = \frac{E}{243.3} \times 100$

$0.70 = \frac{x}{243.3}$

$x = 170.3g$

Try #26, 30, 31, 35, 36 on pages 133-137 Hebden