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12:13 PM

Stoich

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Stoichiometry

Stoichiometry (greek): **Stoicheion** – element, **metry** – to measure

Balanced Chemical equation:

Skills –formula writing
–balancing equations

Tells –substances involved in the chemical rxn
–relationship between the # of particles of each substance

Ex. $N_{2\,(g)} + 3\,H_{2\,(g)} \rightarrow 2\,NH_{3\,(g)}$

1 molecule N_2 + 3 molecules of $H_2 \rightarrow$ 2 molecules of NH_3
Or… 1 mole N_2 + 3 moles of $H_2 \rightarrow$ 2 moles of NH_3

If we use 2 moles of N_2 we could make 4 moles of NH_3

Equation: $2\text{ mol }N_2\left(\frac{2\text{ mol }NH_3}{1\text{ mol }N_2}\right) = 4\text{ mol }NH_3$

Ex. How many moles of oxygen gas are required to produce 5.0 moles of MgO when $Mg_{(s)}$ undergoes combustion? excess

$2\,Mg_{(s)} + O_{2\,(g)} \rightarrow 2MgO_{(s)}$

$$5.0\text{ mol }\cancel{MgO}\left(\frac{1\text{ mol }O_2}{2\text{ mol }\cancel{MgO}}\right) = 2.5\text{ mol }O_2$$

Ex. Aluminum metal will undergo combustion in $O_{2(g)}$ How many moles of aluminum oxide would be produced from 3.70 moles of aluminum?

$4\,Al_{(s)} + 3\,O_{2\,(g)} \rightarrow 2Al_2O_{3\,(s)}$

$$3.70\text{ mol Al}\left(\frac{2\text{ mol }Al_2O_3}{4\text{ mol Al}}\right) = 1.85\text{ mol }Al_2O_3$$

Ex. How many moles of ammonia are produced when you react 12 grams of hydrogen gas with an excess of nitrogen gas?

$N_{2\,(g)} + 3\,H_{2\,(g)} \rightarrow 2\,NH_{3\,(g)}$

$$12\text{ g H}_2\left(\frac{1\text{ mol H}_2}{2.0\text{ g}}\right) = 6.0\text{ mol H}_2$$

$$6.0\text{ mol H}_2\left(\frac{2\text{ mol NH}_3}{3\text{ mol H}_2}\right) = 4.0\text{ mol NH}_3$$

Ex. How many grams of hydrogen gas are required to burn with 24.0 grams of oxygen when making water by combustion?

$2\,H_{2\,(g)} + O_{2\,(g)} \rightarrow 2\,H_2O_{(g)}$ $g\,O_2 \rightarrow mol\,O_2 \rightarrow mol\,H_2 \rightarrow g\,H_2$

$$24.0\text{ g}\cancel{\text{ O}_2}\left(\frac{1\text{ mol O}_2}{32.0\cancel{\text{ g}}}\right) = 0.750\text{ mol O}_2$$

$$0.750\cancel{\text{ mol O}_2}\left(\frac{2\text{ mol H}_2}{1\cancel{\text{ mol O}_2}}\right) = 1.50\text{ mol H}_2$$

$$1.50\cancel{\text{ mol H}_2}\left(\frac{2.0\text{ g H}_2}{1\cancel{\text{ mol H}_2}}\right) = 3.0\text{ g H}_2$$

Pg 124 #1-5

Reactions involving a Limiting Reactant

When chemicals are mixed together in reactions there are two possibilities:

a) Stoichiometric quantities: exactly correct amounts used, so each reactant runs out at the same time (not too likely)

b) One or more reactants are in excess: more of one reactant then is required, the reaction will proceed until the limiting reactant is all used up

* Sandwiches !!

Steps to Follow – A strategy

1) Write and balance the equation for the reaction
2) Convert known masses of substances into moles
3) Calculate the theoretical number of moles of a product that each of the reactants could form
4) Take the smaller amount to be your limiting reactant
5) Solve for final units (grams, concentration, etc…)

Ex. If 75.0g of mercury react with 50.0g of sulphur, how much HgS will be produced?

1) Hg + S → HgS

2) PT

$$\text{Moles of Hg} = 75.0\text{g}\left(\frac{1\text{ mol Hg}}{200.6\text{ g}}\right) = 0.374\text{ mol}$$

$$\text{Moles of S} = 50.0\text{ g}\left(\frac{1\text{ mol S}}{32.1\text{ g}}\right) = 1.56\text{ mol S}$$

3)

$$\text{From } \mathbf{Hg}\text{: } 0.374\text{ mol Hg}\left(\frac{1\text{ mol HgS}}{1\text{ mol Hg}}\right) = \underline{0.374\text{ HgS}}$$ less

$$\text{From } \mathbf{S}\text{: } 1.56\text{ mol S}\left(\frac{1\text{ mol HgS}}{1\text{ mol S}}\right) = 1.56\text{ mol HgS}$$

4) So mercury only gives 0.374 mol of product, to sulphurs 1.56 mol of product, and therefore mercury is the limiting reagent. Now you know that 0.374 mol of HgS will be produced

5) So just calculate the amount of HgS produced. (in g)

$$0.374\text{ mol HgS} \times \left(\frac{232.7\text{g HgS}}{1\text{ mol HgS}}\right) = 87.0\text{ g of HgS produced}$$

Ex. If 15.5g of aluminum react with 46.7g of chlorine how much aluminum chloride will be produced?

$2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$

$$\text{From } \mathbf{Al}\text{: } 15.5\text{g}\left(\frac{1\text{ mol Al}}{27.0\text{ g}}\right)\times\left(\frac{2\text{ mol AlCl}_3}{2\text{ mol Al}}\right) = 0.574 \overset{\text{mol}}{\wedge} \text{AlCl}_3$$

$$\text{From } \mathbf{Cl_2}\text{: } 46.7\text{ g}\left(\frac{1\text{ mol Cl}_2}{71.0\text{ g}}\right)\times\left(\frac{2\text{ mol AlCl}_3}{3\text{ mol Cl}_2}\right) = 0.439\text{ mol AlCl}_3$$ ☆ limiting

So because Cl_2 produces less aluminum chloride it's the limiting reactant! And only 0.439 moles of $AlCl_3$ can be produced.

$$0.439\text{ mol AlCl}_3 \times\left(\frac{133.5\text{g AlCl}_3}{1\text{ mol AlCl}_3}\right) = 58.6\text{ g of AlCl}_3\text{ produced}$$

Ex. How many grams of solid aluminum oxide are produced if 270g of aluminum are combined with 256g of oxygen gas?

$4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3$

$$\text{Moles from Al} = 270\text{g}\left(\frac{1\text{ mol Al}}{27.0\text{ g}}\right)\times\left(\frac{2\text{ mol Al}_2\text{O}_3}{4\text{ mol Al}}\right) = 5.0\text{ mol Al}_2\text{O}_3$$

$$\text{Moles from O}_2 = 256\text{ g}\left(\frac{1\text{ mol O}_2}{32.0\text{ g}}\right)\times\left(\frac{2\text{ mol Al}_2\text{O}_3}{3\text{ mol O}_2}\right) = 5.33\text{ mol Al}_2\text{O}_3$$

Aluminum is limiting so…

$$5.0\text{ mol Al}_2\text{O}_3 \times\left(\frac{102.0\text{g Al}_2\text{O}_3}{1\text{ mol Al}_2\text{O}_3}\right) = 510\text{ g of Al}_2\text{O}_3\text{ produced}$$

Amount of excess

How many liters of Oxygen are left over from the above question?

Moles of reactant – moles required for reaction = moles remaining (excess)

$$\text{moles O}_2\text{excess} = \underbrace{256\text{ g}\left(\frac{1\text{ mol O}_2}{32.0\text{ g}}\right)}_{\text{have}} - \underbrace{5.0\text{ mol Al}_2\text{O}_3\left(\frac{3\text{ mol O}_2}{2\text{ mol Al}_2\text{O}_3}\right)}_{\text{used}} = 0.50\text{mol O}_2$$

$$\text{Liters of O}_2\text{remaining} = \underbrace{(8.0\text{ moles O}_2 - 7.5\text{ moles O}_2)}_{0.5}\times\left(\frac{22.4\text{ L}}{1\text{ mol O}_2}\right) = 11.2\text{ L O}_2$$

Percent Yield

Actual experimental results rarely follow theory (due to side reactions, and impurities, and experimental error etc...) percent yield is a measure of how efficient a particular chemical reaction is in practice.

$$\% \text{ Yield} = \frac{\text{actual yield (usually given in question)}}{\text{theoretical yield (calculated)}} \times 100\%$$

Ex. If 4.0g of hydrogen gas are burned in excess oxygen and 32.0g of water are produced, what is the yield?

$2\ H_2 + O_2 \rightarrow 2\ H_2O$

$$\text{mol of } H_2 = 4.0\text{g}\left(\frac{1 \text{ mol } H_2}{2.0 \text{ g}}\right) = 2.0 \text{ mol } H_2$$

$$2.0 \text{ mol } H_2 \times \left(\frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2}\right) = 2.0 \text{ mol } H_2O$$

$$2.0 \text{ mol } H_2O \times \left(\frac{18.0 \text{ g}}{1 \text{ mol } H_2O}\right) = 36 \text{ g } H_2O$$

$$\% \text{ Yield} = \frac{32\text{g } H_2O}{36\text{g } H_2O} \times 100\% = 89\ \%$$

Percent Composition (purity)

Very similar to % yield, but is concerned with percentage of active ingredient in a sample; such as the % acetic acid in vinegar, or % of Al in bauxite ore.

Ex. Calculate the % composition of copper metal in covellite (CuS) if 37.0kg sample of covellite yields 22.2kg of copper?

$8\ CuS \rightarrow 8\ Cu + S_8$

$$\text{mol of covellite} = 37.0\text{kg}\left(\frac{1000\text{g}}{1\text{ kg}}\right)\left(\frac{1 \text{ mol CuS}}{95.6\text{g CuS}}\right) = 387 \text{ mol CuS}$$

$$387 \text{ mol CuS} \times \left(\frac{8 \text{ mol Cu}}{8 \text{ mol CuS}}\right) \times \left(\frac{63.55 \text{ g}}{1 \text{ mol Cu}}\right) = 24.5\text{ kg Cu}$$

$$\% \text{ } \cancel{\text{Yield}} \text{ Comp} = \frac{22.2\text{kg } \cancel{H_2O}\text{ Cu}}{24.5\text{ kg } \cancel{H_2O}\text{ Cu}} \times 100\% = 90.6\%$$

Pg 137/138 #33-38

lab 6B

Stoich problems with Solutions

Stoich problems with Solutions

Recall: $M = n/V$ and $n = M \times V$

Where M = molarity, n = number of moles, V = volume

Ex. 100.0 mL of 0.50 M $Pb(NO_3)_2$ are mixed with 100.0 mL of 0.50 M KI. How much PbI_2 is precipitated? (assume PbI_2 is totally insoluble)

$$Pb(NO_3)_{2(aq)} + 2\ KI_{(aq)} \rightarrow PbI_{2(s)} + 2\ KNO_{3(aq)}$$

limiting

$$Pb(NO_3)_2 \Rightarrow 0.50\ M \times 0.100\ L \times \left(\frac{1\ mol\ PbI_2}{1\ mol\ Pb(NO_3)_2}\right) = 0.050\ mol\ PbI_2$$

excess

$$KI \Rightarrow 0.50\ M \times 0.100\ L \times \left(\frac{1\ mol\ PbI_2}{2\ mol\ KI}\right) = 0.025\ mol\ PbI_2$$

Since KI produces less it's the limiting reactant…

$$0.025\ mol\ PbI_2 \times \left(\frac{461.0\ g\ PbI_2}{1\ mol\ PbI_2}\right) = 12\ g\ of\ PbI_2\ produced$$

Ex. How many grams of copper will react to completely replace the silver from 208 mL of a 0.100M solution of silver nitrate?

$$2\ AgNO_{3\ (aq)} + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2\ (aq)}$$

Here, we can assume that silver is the limiting reactant because it asks how much copper is needed

$$AgNO_3 \Rightarrow 0.100\ M \times 0.208\ L \times \left(\frac{1\ mol\ Cu}{2\ mol\ AgNO_3}\right) = 0.0104\ mol\ Cu$$

$$0.0104\ mol\ Cu \times \left(\frac{63.55\ g\ Cu}{1\ mol\ Cu}\right) = 0.661\ g\ of\ Cu\ produced$$

Stoich problems with Gases

Recall: 1 mole of any gas at STP has a volume of 22.4L

Ex. 33.6 L of N_2 gas reacts with 44.8 L of H_2 gas. What volume of NH_3 will be produced at STP?

$N_{2\,(g)} + 3\ H_{2\,(g)} \rightarrow 2\ NH_{3\,(g)}$

limit

excess

$$N_2 \Rightarrow 33.6\ \text{L} \times \left(\frac{1\ \text{mol gas}}{22.4\ \text{L}}\right) \times \left(\frac{2\ \text{mol NH}_3}{1\ \text{mol N}_2}\right) = \underline{3.00\ \text{mol NH}_3}$$

$$H_2 \Rightarrow 44.8\ \text{L} \times \left(\frac{1\ \text{mol gas}}{22.4\ \text{L}}\right) \times \left(\frac{2\ \text{mol NH}_3}{3\ \text{mol H}_2}\right) = \underline{1.33\ \text{mol NH}_3}$$

Since H_2 produces less it's the limiting reactant…

$$1.33\ \text{mol NH}_3 \times \left(\frac{22.4\ \text{L NH}_3}{1\ \text{mol NH}_3}\right) = 29.9\ \text{L of NH}_3\ \text{produced}$$

Titrations

An important analytical technique, titrations are used to determine the concentration of a substance in solution by comparing it to a solution of a known concentration; called a **standard**

The reaction between the unknown and the standard is allowed to proceed until mole ratios like those in the balanced chemical equation are reached.

This stoichiometric point (sometimes called the equivalence point or end point) is determined using some sort of chemical indicator

Often these reactions are used with acids/bases, for acid/base titrations the end point is reached when:

Moles of OH^- from the base= moles of H^+ from the acid

Titrations consist of several trials; the first is an estimate, usually done quickly to get an approximate endpoint, almost always the first trial will be over the endpoint (overshot) The average volume of the second and subsequent trials are used in calculations:

Steps to follow ← a strategy

1) Write the balanced equation (as always – you can't go wrong with this)
2) Calculate the moles of standard used (n = M x V)
3) Determine the moles of unknown using the mole ratio
4) Calculate the unknown concentration (M = n/V)

Ex. If 10.0mL of HCl is titrated with 20.0mL of 0.40M NaOH, what is [HCl]?

1 $NaOH_{(aq)} + HCl_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

2 Moles of standard (in this case NaOH) = 0.40 M x 0.0200 L = 8.0×10^{-3} mol

3 Moles of unknown (HCl) =

$$8.0 \times 10^{-3} \text{mol NaOH} \left(\frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}}\right) = 8.0 \times 10^{-3} \text{mol HCl}$$

4 [HCl] =

$$\left(\frac{8.0 \times 10^{-3} \text{mol HCl}}{0.0100 \text{ L}}\right) = \mathbf{0.80 \text{ M HCl}}$$

- Some acids or bases produce two (or more) H^+ or OH^- for each formula unit ($H_3PO_4 \rightarrow$ **3 H^+** + PO_4^{3-})

Ex. An average volume of 24.90mL of 0.100 M NaOH was required to neutralize 15.00mL of H_2SO_4. What is the concentration of the acid?

$2\ NaOH_{(aq)} + H_2SO_{4\,(aq)} \rightarrow 2\ H_2O_{(l)} + Na_2SO_{4\,(aq)}$

Moles of standard (in this case NaOH) = 0.100 M x 0.02490 L = 2.49×10^{-3} mol NaOH

Moles of unknown (~~HCl~~ H_2SO_4) =

$$2.49 \times 10^{-3} \text{mol NaOH} \left(\frac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}}\right) = 1.25 \times 10^{-3} \text{mol } H_2SO_4$$

[~~HCl~~] $[H_2SO_4]$ =

$$\left(\frac{1.25 \times 10^{-3} \text{mol } H_2SO_4}{0.01500 \text{ L}}\right) = \mathbf{0.0833 \text{ M } H_2SO_4}$$

HW Prep Lab 20C
Hand out: titrations