CHE1010 Lecture 5

Balancing equations

Partially balanced equation:

$2NH_3(g) + O_2(g) \rightarrow 2NO + 3H_2O$

No atoms are created or destroyed.

Number of N atoms are same.

2 oxygens on left & 5 oxygens on right.

$2NH_3 + 5O_2 \rightarrow 4\ NO + 6H_2O$

$2NH_3 + 5O_2 \rightarrow 4\ NO + 6H_2O$

10 Oxygens so double the oxygen on right

6 H's on left add coefficient 2 12 on right

$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

<u>Limiting Reagent</u>

Consider the balanced equation

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

3:1 ratio, proportion, mole-mole relationship

1 mol of N_2 requires 3 moles of H_2 for completing the reaction.

If 1:2 H_2 will be consumed first → (limiting reagent) → reaction stops.

N_2 remain unused → (excess reagent)

Limiting reagent limits the amount of product.

Percent yield (%yield)

$$\frac{Actual\ yield\ x\ 100}{Theoretical\ yield} = \%yield$$

The amount of product formed when limiting reagent is completely consumed is theoretical yield.

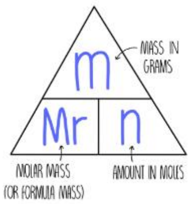
But actual yield ≠ theoretical yield

Because of side reactions or incomplete reactions, actual yield will be lower than the theoretical yield.

Q) 68.5 kg $CO_{(g)}$ reacts with 8.60 kg $H_{2\,(g)}$ to form 3.57×10^4 g CH_3OH. Calculate the % yield.

$2H_{2\,(g)} + CO_{(g)} \rightarrow CH_3OH$

Key: Identify limiting reagent



n, CO = $\frac{68.5}{28} \; x\, 1000 = 2.44\; x\; 10^3\; mols\; CO\; (available)$

$$\frac{8.60}{2.016} x\; 1000 = 4.27\; x\; 10^3\; mol\; H_2$$

Mole ratio : $\frac{2\; mol\; H_2}{1\; mol\; CO}$

Moles of CO required to react with 4.27×10^3 mol H_2 in this ratio.

$$\frac{1 mol\; CO\; x\; 4.27\; x\; 10^3\;\; H_2}{2\; mol\; H_2} = 2.135\; x\; 10^3\; mol\; CO \qquad (consumed)$$

2.44×10^3 mol CO (available)

i.e CO is in excess and H_2 is limiting

To calculate theoretical yield of CH_3OH

Consider the limiting reagent H_2.

H_2 : CH_3OH

2 : 1

$$\frac{4.27\; x\; 10^3 x\; 1\; molCH_3OH}{2\; mol\; H_2} = 2.14\; x\; 10^3\; mol\; CH_3OH\; formed$$

Convert to mass

$2.14 \times 10^3 \times 32.04$ g = 6.86×10^4 g CH_3OH

(Theoretical yield obtained through calculation)

Actual yield is only 3.57×10^4 g

% yield = $\frac{3.57 \times 10^4 \times 100}{6.86 \times 10^4} = 52.0\%\ only$

Lecture 6

<u>Solution Stoichiometry</u>

expressing concentration

Molarity, M = moles of solute in litre of solution

2.00 M = 2 moles dissolved in 1 litre

<u>Dilution</u>

Stock solution : Concentrated solution prepared to save space, which can be diluted to a desired conc.

Add water →

We add water to dilute, no solute is added.

moles of solute before dilution = moles of solute after dilution

n = CV

$M_1V_1 = M_2V_2$

(M_2 may be unknown)

Standard solution

A std solution is a solution whose conc is accurately known.

e.g 0.20 moles dissolved in 1.00 L

Primary std: It is a reagent which can be weighed easily and which is so pure that its mass truly represents the moles of substance contained.

When we dissolve a known mass in known volume we get std soln.

mass/Mr = moles mols/L = concentration

Primary standard should be unreactive, should not absorb moisture.

Then only concentration calculated will be accurate.

e.g Potassium Hydrogen phthalate, KHP, anhydrous Na_2CO_3.

Secondary standard: Not available in pure form. It has to be standardised / titrated against a solution of known concentration (primary std). Stoichiometry is used to establish the strength of a titrant.

e.g $KMnO_4$ (not pure, contains MnO_2)

NaOH (absorbs moisture)

Tertiary standard

Acid base reactions

$HCl + NaOH \rightarrow NaOH + H_2O$ (neutralisation)

$2HCl + Na_2CO_3 \rightarrow NaCl + CO_2 (g) + H_2O$ (gas evolved)

Precipitation reaction (metathesis)

When 2 solutions are mixed, a precipitate is formed.

$K_2CrO_4(aq) + Ba(NO_3)_2 (aq) \rightarrow BaCrO_4 (s) + 2KNO_3 (aq)$

Pot chromate (ppt)

(double displacement)

Ionic equation for the above reaction

$2K^+ (aq) + CrO_4^{2-} (aq) + Ba^{2+} + 2 NO_3^- {}_{(aq)} \rightarrow BaCrO_4(s) + 2K^+(aq) + 2NO_3^- (aq)$

Cancel the common ions (spectator ions), no role

Net ionic equation,

$Ba^{2+}{}_{(aq)} + CrO_4^{2-} (aq) \rightarrow BaCrO_4 (s)$

Q) When aqueous solution of Na_2SO_4 and $Pb(NO_3)_2$ are mixed, $PbSO_4$ precipitates. Calculate the mass of $PbSO_4$ formed when 1.25 L of 0.0500 M $Pb(NO_3)_2$ and 2.00L of 0.0250 M Na_2SO_4 are mixed.

Key: $Pb^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_4{}_{(s)}$

moles of reactants, n = CV

0.050 mol/L Pb^{2+} x 1.25 L = 0.0625 mol Pb^{2+}

0.0250 mol/L SO_4^{2-} x 2.00 L = 0.0500 mol SO_4^{2-}

[Using these moles, we can predict the mass of product but use moles of limiting reagent]

Pb^{2+} : SO_4^{2-} react in 1:1 ratio

Just pick the solution with smaller number of moles.

SO_4^{2-} is limiting. SO_4^{2-} will decide the yield.

SO_4^{2-} :$PbSO_4$ 1:1 ratio

So, moles of $PbSO_4$ = 0.0500 moles

mass = 0.0500 x 303.3 g = 15.2 g $PbSO_4$

<u>Oxidation – Reduction reactions</u> (redox reaction)

$$\overset{0}{2\,Na}\,(s) + \overset{0}{Cl_2}\,_{(g)} \rightarrow 2\,\overset{+}{Na}\overset{-}{Cl}\,_{(s)}$$

Na is losing e and Cl is accepting e . Electrons are transferred.

There involves change in oxidation state.

<u>Oxidation</u>	***<u>Reduction</u>***
e lost	e gained
OX state increases	OX state decreases
Na is the e donor	Cl is the e acceptor
Reducing agent	OX agent
Reductant	oxidant

Here Na is oxidised and Cl is reduced

Oxidation & reduction is taking place : So it's a redox reaction

<u>Assigning oxidation states</u>

Pure element	O
Monoatomic ion	charge on the ion
Na^+	+1
F in a compound	-1
O in a compound	-2
Cl_2 (diatomic)	0