Section 7.4: Calculations Involving Limiting Reagents

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1. Given: $n_{\text{HNO}_3} = 2.3\text{ mol}$; $n_{\text{NaCO}_3} = 2.0\text{ mol}$

Required: amount of water, $n_{\text{H}_2\text{O}}$

Solution:

Step 1. List the given values and the required value.

$HNO_3(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaNO_3(aq)$

2.3 mol 2.0 mol $n_{\text{H}_2\text{O}}$

Step 2. Determine the limiting reagent.

$$n_{\text{NaHCO}_3} = 2.3\ \cancel{\text{mol}_{\text{HNO}_3}} \times \frac{1\ \text{mol}_{\text{NaHCO}_3}}{1\ \cancel{\text{mol}_{\text{HNO}_3}}}$$

$$n_{\text{NaHCO}_3} = 2.3\text{ mol}$$

Since the amount of sodium hydrogen carbonate present initially is less than the required amount, sodium hydrogen carbonate is the limiting reagent.

Step 3. Convert amount of sodium hydrogen carbonate to amount of water.

$$n_{\text{H}_2\text{O}} = 2.0\ \cancel{\text{mol}_{\text{NaHCO}_3}} \times \frac{1\ \text{mol}_{\text{H}_2\text{O}}}{1\ \cancel{\text{mol}_{\text{NaHCO}_3}}}$$

$$n_{\text{H}_2\text{O}} = 2.0\text{ mol}$$

Statement: When 2.3 mol of nitric acid is combined with 2.0 mol of sodium hydrogen carbonate, 2.0 mol of water is produced.

2. Given: $n_{\text{HCl}} = 5.2\text{ mol}$; $n_{\text{MnO}_2} = 1.5\text{ mol}$

Required: amount of chlorine, n_{Cl_2}

Solution:

Step 1. List the given values and the required value.

$4\ HCl(aq) + MnO_2(s) \rightarrow Cl_2(g) + 2\ H_2O(l) + MnCl_2(aq)$

5.2 mol 1.5 mol n_{Cl_2}

Step 2. Determine the limiting reagent.

$$n_{\text{MnO}_2} = 5.2\ \cancel{\text{mol}_{\text{HCl}}} \times \frac{1\ \text{mol}_{\text{MnO}_2}}{4\ \cancel{\text{mol}_{\text{HCl}}}}$$

$$n_{\text{MnO}_2} = 1.3\text{ mol}$$

Since the amount of manganese dioxide present is greater than the required amount, hydrochloric acid is the limiting reagent.

Step 3. Convert amount of hydrochloric acid to amount of chlorine.

$$n_{Cl_2} = 5.2\ \cancel{mol}_{HCl} \times \frac{1\ mol_{Cl_2}}{4\ \cancel{mol}_{HCl}}$$

$$n_{Cl_2} = 1.3\ mol$$

Statement: 1.3 mol of chlorine can be made from 5.2 mol of hydrochloric acid and 1.5 mol of manganese dioxide.

3. (a) $2\ Al(s) + 3\ I_2(s) \rightarrow 2\ AlI_3(s)$

(b) Given: $n_{Al} = 0.50\ mol$; $n_{I_2} = 0.6\ mol$

Required: amount of product, n_{AlI_3}

Solution:

Step 1. List the given values and the required value.

$2\ Al(s)$	+	$3\ I_2(s)$	→	$2\ AlI_3(s)$
0.50 mol		0.60 mol		n_{AlI_3}

Step 2. Determine the limiting reagent.

$$n_{I_2} = 0.50\ \cancel{mol}_{Al} \times \frac{3\ mol_{I_2}}{2\ \cancel{mol}_{Al}}$$

$$n_{I_2} = 0.75\ mol$$

Since the amount of iodine present is less than the required amount, iodine is the limiting reagent.

Step 3. Convert amount of iodine to amount of product.

$$n_{AlI_3} = 0.60\ \cancel{mol}_{I_2} \times \frac{2\ mol_{AlI_3}}{3\ \cancel{mol}_{I_2}}$$

$$n_{AlI_3} = 0.4\ mol$$

Statement: The amount of product that can be made from 0.50 mol of aluminum and 0.60 mol of iodine is 0.4 mol.

4. (a) $2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$

(b) Given: $n_{Al} = 0.26\ mol$; $n_{Fe_2O_3} = 0.10\ mol$

Required: amount of iron, n_{Fe}

Solution:

Step 1. List the given values and the required value.

$2\ Al(s)$	+	$Fe_2O_3(s)$	→	$2\ Fe(s) + Al_2O_3(s)$
0.26 mol		0.10 mol		n_{Fe}

Step 2. Determine the limiting reagent.

$$n_{Fe_2O_3} = 0.26\ \cancel{mol}_{Al} \times \frac{1\ mol_{Fe_2O_3}}{2\ \cancel{mol}_{Al}}$$

$$n_{Fe_2O_3} = 0.13\ mol$$

Since the amount of iron(III) oxide present is less than the required amount, iron(III) oxide is the limiting reagent.

Step 3. Convert amount of iron(III) oxide to amount of iron.

$$n_{Fe} = 0.10\ \cancel{mol_{Fe_2O_3}} \times \frac{2\ mol_{Fe}}{1\ \cancel{mol_{Fe_2O_3}}}$$

$$n_{Fe} = 0.20\ mol$$

Statement: When 0.26 mol of aluminum is combined with 0.10 mol of iron(III) oxide, 0.20 mol of iron is expected.

(c) Given: $n_{Al} = 0.26\ mol$; $n_{Fe_2O_3} = 0.10\ mol$

Required: amount of aluminum oxide, $n_{Al_2O_3}$

Solution:

Step 1. List the given values and the required value.

$2\ Al(s)$	+	$Fe_2O_3(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$	
0.26 mol		0.10 mol	$n_{Al_2O_3}$

Step 2. Convert amount of iron(III) oxide to amount of aluminum oxide.

$$n_{Al_2O_3} = 0.10\ \cancel{mol_{Fe_2O_3}} \times \frac{1\ mol_{Al_2O_3}}{1\ \cancel{mol_{Fe_2O_3}}}$$

$$n_{Al_2O_3} = 0.20\ mol$$

Statement: When 0.26 mol of aluminum is combined with 0.10 mol of iron(III) oxide, 0.10 mol of aluminum oxide is expected.

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1. Given: $m_{SiO_2} = 10.0\ g$; $m_C = 7.00\ g$

Required: mass of silicon carbide, m_{SiC}

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

$SiO_2(s)$ +	$3\ C(s)$	$\rightarrow$	$SiC(s) + 2\ CO_2(g)$
10.0 g	7.00 g		m_{SiC}
60.09 g/mol	12.01 g/mol		40.10 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{SiO_2} = 10.0\ \cancel{g} \times \frac{1\ mol_{SiO_2}}{60.09\ \cancel{g}}$$

$$n_{SiO_2} = 0.166\ 42\ \text{mol [two extra digits carried]}$$

$$n_C = 7.00\ \cancel{g} \times \frac{1\ mol_C}{12.01\ \cancel{g}}$$

$$n_C = 0.582\ 85\ \text{mol [two extra digits carried]}$$

Step 3. Determine the limiting reagent.

$$n_C = 0.166\ 42\ \cancel{mol_{SiO_2}} \times \frac{3\ mol_C}{1\ \cancel{mol_{SiO_2}}}$$

$$n_C = 0.499\ 26\ \text{mol}$$

Since the amount of carbon present is greater than the amount required, silicon dioxide is the limiting reagent.

Step 4. Convert amount of silicon dioxide to amount of silicon carbide.

$$n_{SiC} = 0.166\ 42\ \cancel{mol_{SiO_2}} \times \frac{1\ mol_{SiC}}{1\ \cancel{mol_{SiO_2}}}$$

$$n_{SiC} = 0.166\ 42\ \text{mol}$$

Step 5. Convert amount of silicon carbide to mass of silicon carbide.

$$m_{SiC} = (0.166\ 42\ \cancel{mol})\left(\frac{40.01\ g}{1\ \cancel{mol}}\right)$$

$$m_{SiC} = 6.67\ g$$

Statement: When 10.0 g of silicon dioxide is combined with 7.00 g carbon, 6.67 g silicon carbide is expected.

2. Given: $m_{Fe} = 5.00\ g$; $m_{Cl_2} = 9.00\ g$

Required: mass of iron(III) chloride, m_{FeCl_3}

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

2 Fe(s)	+	3 Cl_2(g)	→	2 $FeCl_3$(s)
5.00 g		9.00g		m_{FeCl_3}
55.85 g/mol		70.90 g/mol		162.20 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{Fe} = 5.00\ \cancel{g} \times \frac{1\ mol_{Fe}}{55.85\ \cancel{g}}$$

$$n_{Fe} = 0.089\ 526\ \text{mol [two extra digits carried]}$$

$$n_{Cl_2} = 9.00\ \cancel{g} \times \frac{1\ mol_{Cl_2}}{70.90\ \cancel{g}}$$

$n_{Cl_2} = 0.126\ 94$ mol [two extra digits carried]

Step 3. Determine the limiting reagent.

$$n_{Cl_2} = 0.089\ 526\ \cancel{mol_{Fe}} \times \frac{3\ mol_{Cl_2}}{2\ \cancel{mol_{Fe}}}$$

$n_{Cl_2} = 0.134\ 289$ mol

Since the amount of chlorine present is less than the required amount, chlorine is the limiting reagent.

Step 4. Convert amount of chlorine to amount of iron(III) chloride.

$$n_{FeCl_3} = 0.126\ 94\ \cancel{mol_{Cl_2}} \times \frac{2\ mol_{FeCl_3}}{3\ \cancel{mol_{Cl_2}}}$$

$n_{FeCl_3} = 0.084\ 627$ mol

Step 5. Convert amount of iron(III) chloride to mass of iron(III) chloride.

$$m_{FeCl_3} = (0.084\ 627\ \cancel{mol})\left(\frac{162.2\ g}{1\ \cancel{mol}}\right)$$

$m_{FeCl_3} = 13.7$ g

Statement: When 5.00 g of iron is combined with 9.00 g carbon, 13.7 g iron(III) chloride is expected.

3. (a) Given: $m_{NH_3} = 0.34 g$; $m_{O_2} = 1.00$ g

Required: limiting reagent

Solution:

Step 1. List the given values and the corresponding molar masses.

4 $NH_3(g)$	+	5 $O_2(g) \rightarrow$ 4 $NO(g)$ + 6 $H_2O(g)$
0.34 g		1.00 g
17.04 g/mol		32.00 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{NH_3} = 0.34\ \cancel{g} \times \frac{1\ mol_{NH_3}}{17.01\ \cancel{g}}$$

$n_{NH_3} = 0.019\ 99$ mol [two extra digits carried]

$$n_{O_2} = 1.00\ \cancel{g} \times \frac{1\ mol_{O_2}}{32.00\ \cancel{g}}$$

$n_{O_2} = 0.031\ 25$ mol [two extra digits carried]

Step 3. Determine the limiting reagent.

$$n_{O_2} = 0.019\ 99\ \cancel{\text{mol}}_{NH_3} \times \frac{5\ \text{mol}_{O_2}}{4\ \cancel{\text{mol}}_{NH_3}}$$

$$n_{O_2} = 0.024\ 99\ \text{mol}$$

Statement: Since the amount of oxygen present is greater than the required amount, ammonia is the limiting reagent.

(b) Given: $m_{NH_3} = 0.34\text{g}$; $m_{O_2} = 1.00\ \text{g}$; $n_{NH_3} = 0.019\ 99\ \text{mol}$

Required: m_{NO}; m_{H_2O}

Solution:

Step 1. List the given values, the required values, and the corresponding molar masses.

$4\ NH_3(g)$	+ $5\ O_2(g)$	→ $4\ NO(g)$	+ $6\ H_2O(g)$
0.01999 mol		n_{NO}	n_{H_2O}
0.34 g	1.00 g	m_{NO}	m_{H_2O}
17.04 g/mol	32.00 g/mol	30.01 g/mol	18.02 g/mol

Step 2. Convert amount of ammonia to amount of nitrogen monoxide.

$$n_{NO} = 0.019\ 99\ \cancel{\text{mol}}_{NH_3} \times \frac{4\ \text{mol}_{NO}}{4\ \cancel{\text{mol}}_{NH_3}}$$

$$n_{NO} = 0.019\ 99\ \text{mol}$$

Step 3. Convert amount of ammonia to amount of water.

$$n_{H_2O} = 0.019\ 99\ \cancel{\text{mol}}_{NH_3} \times \frac{6\ \text{mol}_{H_2O}}{4\ \cancel{\text{mol}}_{NH_3}}$$

$$n_{H_2O} = 0.029\ 99\ \text{mol}$$

Step 4. Convert amount of nitrogen monoxide to mass of nitrogen monoxide.

$$m_{NO} = (0.019\ 99\ \cancel{\text{mol}})\left(\frac{30.01\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$m_{NO} = 0.60\ \text{g}$$

Step 5. Convert amount of water to mass of water.

$$m_{H_2O} = (0.029\ 99\ \cancel{\text{mol}})\left(\frac{18.02\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$m_{H_2O} = 0.54\ \text{g}$$

Statement: When 0.34 g of ammonia combines with 1.00 g of oxygen, 0.60 g of nitrogen monoxide and 0.54 g water are produced in the reaction.

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1. Table 1 Amounts Involved in the Synthesis of Water

$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$			
Amount of hydrogen (mol)	**Amount of oxygen (mol)**	**Amount of water (mol)**	**Amount of excess reagent remaining (mol)**
2	2	**2**	**1 mol O_2**
6	2	**4**	**2 mol H_2**
0.4	0.8	**0.4**	**0.6 mol O_2**
8	**2.5**	5	3 mol H_2

2. Table 2 Amounts Involved in the Synthesis of Ammonia

$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$			
Amount of nitrogen (mol)	**Amount of hydrogen (mol)**	**Amount of ammonia (mol)**	**Amount of excess reagent remaining (mol)**
4	13	**8**	**1 mol H_2**
0.90	0.25	**0.17**	**0.82 mol N_2**
0.30	**0.24**	0.16	0.22 mol N_2
1.5	**5.0**	3.0	0.50 mol H_2
1.4	**5.0**	**2.8**	0.80 mol H_2

3. (a) Given: $n_{\text{Mg}} = 0.58\ \text{mol}$; $n_{\text{N}_2} = 0.20\ \text{mol}$

Required: limiting reagent; excess reagent

Solution:

Step 1. List the given values.

$3\ \text{Mg(s)} + \text{N}_2\text{(g)} \rightarrow \text{Mg}_3\text{N}_2\text{(s)}$

0.58 mol 0.20 mol

Step 2. Determine the limiting reagent.

$$n_{\text{N}_2} = 0.58\ \cancel{\text{mol}}_{\text{Mg}} \times \frac{1\ \text{mol}_{\text{N}_2}}{3\ \cancel{\text{mol}}_{\text{Mg}}}$$

$$n_{\text{N}_2} = 0.19\ \text{mol}$$

Statement: Since the amount of nitrogen present is greater than the required amount, magnesium is the limiting reagent and nitrogen is the excess reagent.

(b) Given: $n_{\text{Ca}} = 5.3\ \text{mol}$; $n_{\text{AlCl}_3} = 3.8\ \text{mol}$

Required: limiting reagent; excess reagent

Solution:

Step 1. List the given values.

$3\ \text{Ca(s)} + 2\ \text{AlCl}_3\text{(aq)} \rightarrow 3\ \text{CaCl}_2\text{(aq)} + 2\ \text{Al(s)}$

5.3 mol 3.8 mol

Step 2. Determine the limiting reagent.

$$n_{\text{AlCl}_3} = 5.3\ \cancel{\text{mol}}_{\text{Ca}} \times \frac{2\ \text{mol}_{\text{AlCl}_3}}{3\ \cancel{\text{mol}}_{\text{Ca}}}$$

$$n_{\text{AlCl}_3} = 3.5\ \text{mol}$$

Statement: Since the amount of aluminum chloride present is greater than the amount required to react with calcium, calcium is the limiting reagent and aluminum chloride is the excess reagent.

(c) Given: $n_{FeS_2} = 0.10$ mol; $n_{O_2} = 0.35$ mol
Required: limiting reagent; excess reagent
Solution:
Step 1. List the given values.
$4\ FeS_2(s) + 11\ O_2(g) \rightarrow 2\ Fe_2O_3(s) + 8\ SO_2(g)$
0.10 mol 0.35 mol
Step 2. Determine the limiting reagent.

$$n_{O_2} = 0.10\ \cancel{mol}_{FeS_2} \times \frac{11\ mol_{O_2}}{4\ \cancel{mol}_{FeS_2}}$$

$$n_{O_2} = 0.28\ mol$$

Statement: Since the amount of oxygen present is greater than the required amount, iron pyrite is the limiting reagent and oxygen is the excess reagent.
4. (a) Given: $n_{Cu} = 0.24$ mol; $n_{AgNO_3} = 0.52$ mol
Required: amount of silver, n_{Ag}
Solution:
Step 1. List the given values and the required value.
$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$
0.24 mol 0.52 mol n_{Ag}
Step 2. Determine the limiting reagent.

$$n_{AgNO_3} = 0.24\ \cancel{mol}_{Cu} \times \frac{2\ mol_{AgNO_3}}{1\ \cancel{mol}_{Cu}}$$

$$n_{AgNO_3} = 0.48\ mol$$

Since the amount of silver nitrate present is greater than the amount required to react with copper, copper is the limiting reagent.
Step 3. Convert amount of copper to amount of silver.

$$n_{Ag} = 0.24\ \cancel{mol}_{Cu} \times \frac{2\ mol_{Ag}}{1\ \cancel{mol}_{Cu}}$$

$$n_{Ag} = 0.48\ mol$$

Statement: If 0.24 mol of copper were combined with 0.52 mol of silver nitrate, 0.48 mol of silver would be produced.
(b) Silver nitrate is in excess.
amount of silver nitrate present initially = 0.52 mol
amount of silver nitrate required = 0.48 mol
0.52 mol – 0.48 mol = 0.04 mol
Therefore, the amount of silver nitrate remaining would be 0.04 mol.

5. (a) Given: $n_{Al} = 0.35\ \text{mol}$; $n_{HCl} = 1.2\ \text{mol}$

Required: amount of aluminum chloride, n_{AlCl_3}

Solution:

Step 1. List the given values and the required value.

$2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$

0.35 mol   1.2 mol

Step 2. Determine the limiting reagent.

$$n_{HCl} = 0.35\ \cancel{\text{mol}}_{Al} \times \frac{6\ \text{mol}_{HCl}}{2\ \cancel{\text{mol}}_{Al}}$$

$$n_{HCl} = 1.05\ \text{mol}$$

Since the amount of hydrochloric acid present is greater than the amount required to react with aluminum, aluminum is the limiting reagent.

Step 3. Convert amount of aluminum to amount of aluminum chloride.

$$n_{AlCl_3} = 0.35\ \cancel{\text{mol}}_{Al} \times \frac{2\ \text{mol}_{AlCl_3}}{2\ \cancel{\text{mol}}_{Al}}$$

$$n_{AlCl_3} = 0.35\ \text{mol}$$

Statement: When 0.35 mol of aluminum is combined with 1.2 mol of hydrochloric acid, 0.35 mol of aluminum chloride will be produced.

(b) Hydrochloric acid is in excess.

amount of hydrochloric acid present initially = 1.2 mol

amount of hydrochloric acid required = 1.05 mol

1.2 mol – 1.05 mol = 0.15 mol

Therefore, the amount of hydrochloric acid remaining will be 0.15 mol.

6. Given: $n_{SO_2} = 5.8\ \text{mol}$; $n_{O_2} = 2.8\ \text{mol}$

Required: mass of sulfur trioxide, m_{SO_3}

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

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

5.8 mol   2.8 mol m_{SO_3}

64.07 g/mol   32.00 g/mol   80.07 g/mol

Step 2. Determine the limiting reagent.

$$n_{O_2} = 5.8\ \cancel{\text{mol}}_{SO_2} \times \frac{1\ \text{mol}_{O_2}}{2\ \cancel{\text{mol}}_{SO_2}}$$

$$n_{O_2} = 2.9\ \text{mol}$$

Since the amount of oxygen present is less than the amount required to react with sulfur dioxide, oxygen is the limiting reagent.

Step 3. Convert amount of oxygen to amount of sulfur trioxide.

$$n_{SO_3} = 2.8\ \cancel{mol_{O_2}} \times \frac{2\ mol_{SO_3}}{1\ \cancel{mol_{O_2}}}$$

$$n_{SO_3} = 5.6\ mol$$

Step 4. Convert amount of sulfur trioxide to mass of sulfur trioxide.

$$m_{SO_3} = (5.6\ \cancel{mol})\left(\frac{86.07\ g}{1\ \cancel{mol}}\right)$$

$$m_{SO_3} = 450\ g$$

Statement: When 5.8 mol of sulfur dioxide and 2.8 mol of oxygen are combined, 450 g of sulfur trioxide will be produced.

7. (a) $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

(b) Given: $m_{H_2} = 10.0\ g$; $m_{Cl_2} = 320.0\ g$

Required: mass of hydrogen chloride, m_{HCl}

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

$H_2(g)$	+	$Cl_2(g)$	$\rightarrow$	$2\ HCl(g)$
10.0 g		320.0 g		m_{HCl}
2.02 g/mol		70.90 g/mol		36.46 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{H_2} = 10.0\ \cancel{g} \times \frac{1\ mol_{H_2}}{2.02\ \cancel{g}}$$

$n_{H_2} = 4.9505$ mol [two extra digits carried]

$$n_{Cl_2} = 320.0\ \cancel{g} \times \frac{1\ mol_{Cl_2}}{70.90\ \cancel{g}}$$

$n_{Cl_2} = 4.5134$ mol [two extra digits carried]

Step 3. Determine the limiting reagent.

The mole ratio of hydrogen to chlorine is 1:1, the amount of chlorine present is less than the amount of chlorine required to react with hydrogen. Therefore, chlorine is the limiting reagent.

Step 4. Convert amount of chlorine to amount of hydrogen chloride.

$$n_{HCl} = 4.5134\ \cancel{mol_{Cl_2}} \times \frac{2\ mol_{HCl}}{1\ \cancel{mol_{Cl_2}}}$$

$n_{HCl} = 9.0268$ mol [two extra digits carried]

Step 5. Convert amount of hydrogen chloride to mass of hydrogen chloride.

$$m_{HCl} = (9.0268\ \cancel{mol})\left(\frac{36.46\ g}{1\ \cancel{mol}}\right)$$

$$m_{HCl} = 329\ g$$

Statement: If 10.0 g of hydrogen mixes with 320.0 g of chlorine, 329 g of hydrogen chloride will be produced.

8. Given: $m_{Al(OH)_3} = 0.50\ g$; $m_{HCl} = 0.60\ g$

Required: mass of aluminum chloride, m_{AlCl_3}

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

$Al(OH)_3(s)$ +	$3\ HCl(aq)$ →	$3\ H_2O(l)$ +	$AlCl_3(aq)$
0.50 g	0.60 g		m_{AlCl_3}
78.01 g/mol	36.46 g/mol		133.33 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{Al(OH)_3} = 0.50\ \cancel{g} \times \frac{1\ mol_{Al(OH)_3}}{78.01\ \cancel{g}}$$

$$n_{Al(OH)_3} = 0.006\ 409\ mol\ \text{[two extra digits carried]}$$

$$n_{HCl} = 0.60\ \cancel{g} \times \frac{1\ mol_{HCl}}{36.46\ \cancel{g}}$$

$$n_{HCl} = 0.016\ 46\ mol\ \text{[two extra digits carried]}$$

Step 3. Determine the limiting reagent.

$$n_{HCl} = 0.006\ 409\ \cancel{mol_{Al(OH)_3}} \times \frac{3\ mol_{HCl}}{1\ \cancel{mol_{Al(OH)_3}}}$$

$$n_{HCl} = 0.019\ 23\ mol\ \text{[two extra digits carried]}$$

Since the amount of hydrochloric acid present is less than the amount required to react with aluminum hydroxide, hydrochloric acid is the limiting reagent.

Step 4. Convert amount of hydrochloric acid to amount of aluminum chloride.

$$n_{AlCl_3} = 0.016\ 46\ \cancel{mol_{HCl}} \times \frac{1\ mol_{AlCl_3}}{3\ \cancel{mol_{HCl}}}$$

$$n_{AlCl_3} = 0.005\ 487\ mol\ \text{[two extra digits carried]}$$

Step 5. Convert amount of aluminum chloride to mass of aluminum chloride.

$$m_{AlCl_3} = (0.005\ 487\ \cancel{mol})\left(\frac{133.33\ g}{1\ \cancel{mol}}\right)$$

$$m_{AlCl_3} = 0.73\ g$$

Statement: If 0.50 g of aluminum hydroxide is placed in a solution containing 0.60 g of hydrochloric acid, 0.73 g of aluminum chloride will be produced.

9. Given: $m_{C_4H_{10}} = 10.0\ g$; $m_{O_2} = 30.0\ g$

Required: mass of carbon dioxide, m_{CO_2}

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

$2\ C_4H_{10}(g)\ +$	$13\ O_2(g)\ \rightarrow$	$8\ CO_2(g) + 10\ H_2O(g)$
10.0 g	30.0 g	m_{CO_2}
58.14 g/mol	32.00 g/mol	44.01 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{C_4H_{10}} = 10.0\ \cancel{g} \times \frac{1\ mol_{C_4H_{10}}}{58.14\ \cancel{g}}$$

$$n_{C_4H_{10}} = 0.172\ 00\ mol\ \text{[two extra digits carried]}$$

$$n_{O_2} = 30.0\ \cancel{g} \times \frac{1\ mol_{O_2}}{32.00\ g}$$

$$n_{O_2} = 0.937\ 50\ mol\ \text{[two extra digits carried]}$$

Step 3. Determine the limiting reagent.

$$n_{O_2} = 0.172\ 00\ \cancel{mol_{C_4H_{10}}} \times \frac{13\ mol_{O_2}}{2\ \cancel{mol_{C_4H_{10}}}}$$

$$n_{O_2} = 1.1180\ mol\ \text{[two extra digits carried]}$$

Since the amount of oxygen present is less than the amount required to react with butane, oxygen is the limiting reagent.

Step 4. Convert amount of oxygen to amount of carbon dioxide.

$$n_{CO_2} = 0.937\ 50\ \cancel{mol_{O_2}} \times \frac{8\ mol_{CO_2}}{13\ \cancel{mol_{O_2}}}$$

$$n_{CO_2} = 0.576\ 92\ mol\ \text{[two extra digits carried]}$$

Step 5. Convert amount of carbon dioxide to mass of carbon dioxide.

$$m_{CO_2} = (0.576\ 92\ \cancel{mol})\left(\frac{44.01\ g}{1\ \cancel{mol}}\right)$$

$$m_{CO_2} = 25.4\ g$$

Statement: The mass of carbon dioxide produced from the combustion of 10.0 g of butane and 30.0 g of oxygen is 25.4 g.

10. (a) Given: $m_{TiO_2} = 40.0\ g$; $m_C = 7.0\ g$; $m_{Cl_2} = 30.0\ g$

Required: limiting reagent

Solution:

Step 1. List the given values and the corresponding molar masses.

$TiO_2(s)$ +	$C(s)$ +	$2\ Cl_2(g) \rightarrow TiCl_4(g) + CO_2(g)$
40.0 g	7.0 g	30.0 g
79.87 g/mol	12.01 g/mol	70.90 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{TiO_2} = 40.0\ \cancel{g} \times \frac{1\ mol_{TiO_2}}{79.87\ \cancel{g}}$$

$$n_{TiO_2} = 0.500\ 81\ mol\ \text{[two extra digits carried]}$$

$$n_{Cl_2} = 30.0\ \cancel{g} \times \frac{1\ mol_{Cl_2}}{70.90\ \cancel{g}}$$

$$n_{Cl_2} = 0.423\ 13\ mol\ \text{[two extra digits carried]}$$

$$n_C = 7.0\ \cancel{g} \times \frac{1\ mol_C}{12.01\ \cancel{g}}$$

$$n_C = 0.5828\ mol\ \text{[two extra digits carried]}$$

Statement: According to the balanced chemical equation, the mole ratio of titanium(IV) oxide to carbon to chlorine is 1:1:2. The amount of chlorine required is twice the amount of titanium or the amount of carbon. Since the amount of chlorine is less than the minimum of these two amounts, chlorine is the limiting reagent.

(b) Given: $n_{Cl_2} = 0.42313\ mol$

Required: mass of titanium(IV) chloride, m_{TiCl_4}

Solution:

Step 1. List the given value, the required value, and the corresponding molar masses.

$TiO_2(s) + C(s) +$	$2\ Cl_2(g)$ $\rightarrow$	$TiCl_4(g) + CO_2(g)$
	0.42313 mol	m_{TiCl_4}
	70.90 g/mol	189.67 g/mol

Step 2. Convert amount of chlorine to amount of titanium(IV) chloride.

$$n_{TiCl_4} = 0.423\ 13\ \cancel{mol}_{Cl_2} \times \frac{1\ mol_{TiCl_4}}{2\ \cancel{mol}_{Cl_2}}$$

$n_{TiCl_4} = 0.211\,57$ mol [two extra digits carried]

Step 3. Convert amount of titanium(IV) chloride to mass of titanium(IV) chloride.

$$m_{TiO_4} = (0.211\ 57\ \cancel{mol})\left(\frac{189.67\ g}{1\ \cancel{mol}}\right)$$

$m_{TiO_4} = 40.1$ g

Statement: The mass of titanium(IV) chloride that can be produced from mixing 40.0 g of titanium(IV) oxide, 7.0 g of carbon, and 30.0 g of chlorine is 40.1 g.

11. Given: $m_{NH_3} = 15.0$ g ; $m_{CH_4} = 6.0$ g

Required: mass of hydrogen cyanide, m_{HCN}

Analysis: Use the equations to determine the mole ratio of ammonia to methane in the process.

$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

amount of NH_3 to amount of NO = 4:4 = 1:1

$2\ NO(g) + 2\ CH_4(g) \rightarrow 2\ HCN(g) + 2\ H_2O(g) + H_2(g)$

amount of NO to amount of CH_4 = 2:2 = 1:1

Therefore, the amounts of NH_3 and CH_4 are in the ratio 1:1.

Then, determine the mass of hydrogen cyanide produced from the given masses of ammonia and methane.

Solution:

Step 1. List the given values, the required value, and the corresponding molar masses.

$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

15.0 g

17.04 g/mol

2 NO(g) +	2 CH_4(g)	→	2 HCN(g) + 2 H_2O(g) + H_2(g)
	6.0 g		m_{HCN}
	16.05 g/mol		27.03 g/mol

Step 2. Convert mass of given substances to amount of given substances.

$$n_{NH_3} = 15.0\ \cancel{g} \times \frac{1\ mol_{NH_3}}{17.04\ \cancel{g}}$$

$n_{NH_3} = 0.880\ 28$ mol [two extra digits carried]

$$n_{CH_4} = 6.0\ \cancel{g} \times \frac{1\ mol_{CH_4}}{16.05\ \cancel{g}}$$

$n_{CH_4} = 0.3738$ mol [two extra digits carried]

Step 3. Determine the limiting reagent.

$$n_{CH_4} = 0.880\ 28\ \cancel{mol_{NH_3}} \times \frac{1\ mol_{CH_4}}{1\ \cancel{mol_{NH_3}}}$$

$$n_{CH_4} = 0.880\ 28\ mol\ \text{[two extra digits carried]}$$

Since the amount of methane present is less than the required amount, methane is the limiting reagent.

Step 4. Convert amount of methane to amount of hydrogen cyanide.

$$n_{HCN} = 0.3738\ \cancel{mol_{CH_4}} \times \frac{2\ mol_{HCN}}{2\ \cancel{mol_{CH_4}}}$$

$$n_{HCN} = 0.3738\ mol\ \text{[two extra digits carried]}$$

Step 5. Convert amount of hydrogen cyanide to mass of hydrogen cyanide.

$$m_{HCN} = (0.3738\ \cancel{mol})\left(\frac{27.03\ g}{1\ \cancel{mol}}\right)$$

$$m_{HCN} = 10\ g$$

Statement: If 15.0 g of ammonia and 6.0 g of methane are present initially, 10 g of hydrogen cyanide will be produced.