Answers to Exercise 1.4
Stoichiometry Basics

Tables like the ones in this answer key are enormously helpful at keeping you organized! Sometimes, they even suggest the next logical step in a calculation…

1.	$C_3H_8(g)$	+	$5\ O_2(g)$	→	$4\ H_2O(g)$	+	$3\ CO_2(g)$
M	44.096 g/mol		31.9988 g/mol		18.0152 g/mol		44.010 g/mol
$m_{initial}$	25 g						
$n_{initial}$	0.57 mol		2.8 mol		0 mol		0 mol
n_{change}	-0.57 mol		-2.8 mol		+2.3 mol		+1.7 mol
n_{final}	0 mol		0 mol		2.3 mol		1.7 mol
$m_{initial}$	25 g		91 g		0 g		0 g
m_{final}	0 g		0 g		41 g		75 g

Step 1: Write a balanced chemical equation for the reaction

see above

Step 2: Organize all known information

see above; this involves reporting masses, concentrations, pressures, etc. as well as calculating molar masses, checking the data sheet for relevant densities, etc.

Step 3: Calculate moles of C_3H_8 reacted ($n_{initial}$)

$$n_{C_3H_8-initial} = 25\ g\ C_3H_8 \times \frac{1mol}{44.096g} = 0.57\ mol\ C_3H_8$$

Step 4: Use mole ratio to calculate moles of O_2 required for reaction ($n_{initial}$)

$$n_{O_2-initial} = 0.57\ mol\ C_3H_8 \times \frac{5\ mol\ O_2}{1\ mol\ C_3H_8} = 2.8\ mol\ O_2$$

Step 5: Calculate mass of O_2 required for reaction ($m_{initial}$)

$$m_{O_2-initial} = 2.8\ mol\ O_2 \times \frac{31.9988g}{1\ mol} = 91\ g\ O_2$$ *answer to (a)*

Step 6: Use mole ratio to calculate moles of H_2O and CO_2 produced (n_{final})

$$n_{H_2O-final} = 0.57\ mol\ C_3H_8 \times \frac{4\ mol\ H_2O}{1\ mol\ C_3H_8} = 2.3\ mol\ H_2O$$

$$n_{CO_2-final} = 0.57\ mol\ C_3H_8 \times \frac{3\ mol\ CO_2}{1\ mol\ C_3H_8} = 1.7\ mol\ CO_2$$

Step 7: Calculate masses of H_2O and CO_2 produced (m_{final})

$$m_{H_2O-final} = 2.3\ mol\ H_2O \times \frac{18.0152g}{1\ mol} = 41\ g\ H_2O$$ *answer to (b)*

$$m_{CO_2-final} = 1.7\ mol\ CO_2 \times \frac{44.010g}{1\ mol} = 75\ g\ CO_2$$ *answer to (c)*

Step 8: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

25 g + 91 g = 41 g + 75 g = 116 g so the Law of Conservation of Mass is obeyed. Therefore, this seems like a reasonable set of answers. *answer to (d)*

2.	$4\ Fe(s)$	+	$3\ O_2(g)$	→	$2\ Fe_2O_3(s)$
M	55.847 g/mol		31.9988 g/mol		159.692 g/mol
$m_{initial}$	25 g		25 g		0 g
$n_{initial}$	0.45 mol		0.78 mol		0 mol
n_{change}	-0.45 mol		-0.34 mol		+ 0.22 mol
n_{final}	0 mol		0.45 mol		0.22 mol
m_{final}	0 g		14 g		36 g

Step 1: Write a balanced chemical equation for the reaction

see above

Step 2: Organize all known information

see above

Step 3: Calculate moles of Fe and O_2 available to react ($n_{initial}$)

$$n_{Fe-initial} = 25\ g\ Fe \times \frac{1mol}{55.847g} = 0.45\ mol\ Fe$$

$$n_{O_2-initial} = 25\ g\ O_2 \times \frac{1mol}{31.9988g} = 0.78\ mol\ O_2$$

Step 4: Identify the limiting reagent

Either calculate the moles of O_2 required to fully react with 0.45 mol Fe:

$n_{O_2} = 0.45\ mol\ Fe \times \frac{3\ mol\ O_2}{4\ mol\ Fe} = 0.34\ mol\ O_2$ *There is more O_2 than this; Fe runs out first.*

Or calculate the moles of Fe required to fully react with 0.78 mol O_2:

$n_{Fe} = 0.78\ mol\ O_2 \times \frac{4\ mol\ Fe}{3\ mol\ O_2} = 1.0\ mol\ Fe$ *There is less Fe than this; Fe runs out first.*

Or calculate the maximum moles of Fe_2O_3 that can be produced from each reactant:

$n_{Fe_2O_3} = 0.45\ mol\ Fe \times \frac{2\ mol\ Fe_2O_3}{4\ mol\ Fe} = 0.22\ mol\ Fe_2O_3$ *This is the smaller amount;*

$n_{Fe_2O_3} = 0.78\ mol\ O_2 \times \frac{2\ mol\ Fe_2O_3}{3\ mol\ O_2} = 0.52\ mol\ Fe_2O_3$ *Fe runs out first.*

Regardless of the method chosen, Fe is the limiting reagent.

Step 5: Calculate moles of Fe_2O_3 produced (n_{final}) *(if not already done in Step 4)*

$$n_{Fe_2O_3-final} = 0.45\ mol\ Fe \times \frac{2\ mol\ Fe_2O_3}{4\ mol\ Fe} = 0.22\ mol\ Fe_2O_3$$

Step 6: Calculate mass of Fe_2O_3 produced (m_{final})

$$m_{Fe_2O_3-final} = 0.22\ mol\ Fe_2O_3 \times \frac{159.,6982\ g}{1\ mol} = 36\ g\ Fe_2O_3$$

Step 7: Calculate moles of O_2 left unreacted (n_{final})

$$n_{O_2-final} = n_{O_2-initial} - n_{O_2-reacted} = 0.78\ mol\ O_2 - 0.34\ mol\ O_2 = 0.45\ mol\ O_2$$

Step 8: Calculate mass of O_2 left unreacted (m_{final})

$$m_{O_2-final} = 0.45\ mol\ O_2 \times \frac{31.9988g}{1\ mol} = 14\ g\ O_2$$

Step 9: Check your work

Does your answer seem reasonable? Are sig. fig. correct?
25 g + 25 g + 0 g = 0 g + 14 g + 36 g = 50 g so this seems like a reasonable set of answers.