Answers to Exercise 6.5
Stoichiometry Practice

1.	$4\ Li(s)$	+	$O_2(g)$	→	$2\ Li_2O(s)$
M	6.941 g/mol		31.9988 g/mol		29.8814 g/mol
m_{final}					12 g
$n_{initial}$	0.80 mol		0.20 mol		0 mol
n_{change}	-0.80 mol		-0.20 mol		+0.40 mol
n_{final}	0 mol		0 mol		0.40 mol
$m_{initial}$	5.6 g		6.4 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 Li_2O produced (n_{final})

$$n_{Li_2O-final} = 12g \times \frac{1mol}{29.8814g} = 0.40mol$$

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

$$n_{Li-initial} = 0.40\ mol\ Li_2O \times \frac{4\ mol\ Li}{2\ mol\ Li_2O} = 0.80\ mol\ Li$$

$$n_{O_2-initial} = 0.40\ mol\ Li_2O \times \frac{1\ mol\ O_2}{2\ mol\ Li_2O} = 0.20\ mol\ O_2$$

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

$$m_{Li-initial} = 0.80\ mol\ Li \times \frac{6.941g}{1\ mol} = 5.6g\ Li$$

$$m_{O_2-initial} = 0.20\ mol\ O_2 \times \frac{31.9988g}{1\ mol} = 6.4\ g\ O_2$$

Step 6: Check your work

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

5.6 g + 6.4 g = 12 g so this seems like a reasonable pair of answers.

2.	$2\ K(s)$	+	$2\ H_2O(l)$	→	$2\ KOH(aq)$	+	$H_2(g)$
M	39.0983 g/mol		18.0152 g/mol		56.1056 g/mol		2.0158 g/mol
d			1.00 g/mL				
V			100. mL				
$m_{initial}$	0.25 g		100. g				
$n_{initial}$	0.0064 mol		5.6 mol		0 mol		0 mol
n_{change}	-0.0064 mol		-0.0064 mol		+0.0064 mol		+0.0032 mol
n_{final}	0 mol		5.6 mol		0.0064 mol		0.0032 mol
c_{final}					0.064 mol/L		

Step 1: Write a balanced chemical equation for the reaction

see above

Step 2: Organize all known information

see above; values in grey are not necessary for this calculation

Step 3: Calculate moles of K and H_2O ($n_{initial}$)

$$n_{K-initial} = 0.25g \times \frac{1mol}{39.0983g} = 0.0064mol$$

$$n_{H_2O-initial} = 100.mL \times \frac{1.00g}{1mL} \times \frac{1mol}{18.0152g} = 5.6mol$$

Step 4: Identify the limiting reagent

0.0064 mol H_2O are required to react with 0.0064 mol K. Since there is more H_2O than this, the K will run out before the H_2O. K is therefore the limiting reagent.

Step 5: Use mole ratio to calculate moles of KOH produced (n_{final})

$$n_{KOH-final} = 0.0064\ mol\ K \times \frac{2\ mol\ KOH}{2\ mol\ K} = 0.0064\ mol\ KOH$$

Step 6: Calculate concentration of KOH(aq) (c_{final})

$$c_{KOH-final} = \frac{0.0064\ mol}{100.mL} \times \frac{1000mL}{1L} = 0.064\frac{mol}{L}$$

Divide moles of KOH by volume of solution (assumed to be equal to initial volume of water)

Step 7: Check your work

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

3.	$Mg(s)$	+	$2\ HCl(aq)$	$\rightarrow$	$MgCl_2(aq)$	+	$H_2(g)$
M	24.305 g/mol		36.4606 g/mol		95.210 g/mol		2.0158 g/mol
$c_{initial}$			1.25 mol/L				
V			75 mL				
m_{inital}	185 mg						
$n_{initial}$	0.00761mol		0.0938 mol		0 mol		0 mol
n_{change}	-0.00761mol		-0.0152 mol		+0.00761 mol		+0.00761 mol
n_{final}	0 mol		0.0785 mol		0.00761 mol		0.00761 mol
c_{final}			1.0 mol/L				

Step 1: Write a balanced chemical equation for the reaction

see above; you could also have written a net ionic equation

Step 2: Organize all known information

see above; values in grey are not necessary for this calculation

Step 3: Calculate moles of Mg and HCl ($n_{initial}$)

$$n_{Mg-initial} = 185mg \times \frac{1g}{1000mg} \times \frac{1mol}{24.305g} = 0.00761mol$$

$$n_{HCl-initial} = 75mL \times \frac{1L}{1000mL} \times \frac{1.25mol}{1L} = 0.0938mol$$

Step 4: Identify the limiting reagent

2(0.00761 mol) = 0.0152 mol HCl are required to react with 0.00761 mol Mg. Since there is more HCl than this, the Mg will run out before the HCl. Mg is therefore the limiting reagent.

Step 5: Calculate the moles of HCl left after the reaction is complete (n_{final})

$$n_{HCl-final} = n_{HCl-initial} - n_{HCl-reacted}$$

$$n_{HCl-final} = (0.0938mol) - (0.0152mol) = 0.0785mol$$

Step 6: Calculate the concentration of HCl(aq) after the reaction is complete (c_{final})

In the absence of contradicting information, assume that the volume of solution does not change.

$$c_{HCl-final} = \frac{0.0785\ mol}{75mL} \times \frac{1000mL}{1L} = 1.0\frac{mol}{L}$$

Step 7: Check your work

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

The concentration of HCl has decreased. This is consistent with some of the HCl being used up in the reaction. Since it was not the limiting reagent, some HCl remains and the concentration is not zero.