Answers to Practice Test Questions 1
Math and Stoichiometry Review

1. $1\ m^3 = (100\ cm)^3 = 1\ 000\ 000\ cm^3 = 10^6\ cm^3$

$$1\ m^3 = 1\ 000\ 000\ cm^3 \times \frac{1\ mL}{1\ cm^3} \times \frac{1\ L}{1000\ mL} = 1000\ L$$

2.

(a) $PV = nRT$ *therefore* $n = \frac{PV}{RT}$

$$n_{Cl_2} = \frac{(75\ kPa)(125\ mL)}{\left(8.314462 \frac{Pa \cdot m^3}{mol \cdot K}\right)(295.15K)} \times \frac{1000\ Pa}{1\ kPa} \times \frac{1\ L}{1000\ mL} \times \frac{1\ m^3}{1000\ L} = 0.0038\ mol$$

You may find it easier to do all of the unit conversions first then plug those values into the formula. If you do that, you still need to write the units to ensure that they all cancel properly.

(b) $m_{Cl_2} = 0.0038\ mol \times \frac{70.9054\ g}{1\ mol} = 0.27\ g$

3. **Step 1: Calculate mass of ethanol (C_2H_5OH; commonly abbreviated as EtOH)**

$d = \frac{m}{V}$ *therefore* $m = V \times d$

$$m_{EtOH} = 1.00\ L\ EtOH \times \frac{1000\ mL}{1\ L} \times 0.789 \frac{g}{mL} = 789\ g\ EtOH$$

Step 2: Calculate moles of ethanol (molar mass of ethanol is 46.069 g/mol)

$$n_{EtOH} = 789\ g\ EtOH \times \frac{1\ mol}{46.069\ g} = 17.1\ mol\ EtOH$$

Step 3*: Calculate number of molecules of ethanol

$$N_{EtOH} = 17.1\ mol\ EtOH \times \frac{6.022141 \times 10^{23}\ molecules}{1\ mol} = 1.03 \times 10^{25}\ molecules\ EtOH$$

Step 4*: Calculate number of atoms of hydrogen

$$N_H = 1.03 \times 10^{25}\ molecules\ EtOH \times \frac{6\ atoms\ H}{1\ molecule\ EtOH} = 6.19 \times 10^{25}\ atoms\ H$$

Step 5: Check your work

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

*The answer is very large (…×10^{25}) – which is to be expected. There *should* be a lot of hydrogen atoms in one liter of ethanol. If the answer had been less than one, you should have recognized that that would be impossible and gone back to find your error.*

* Steps 3 and 4 can be reversed by calculating the moles of H atoms in the sample of EtOH then calculating the number of H atoms in the sample.

4.	$Ca(s)$	+	$2\ H_2O(l)$	→	$Ca(OH)_2(s)$	+	$H_2(g)$
M	40.078 g/mol		18.0152 g/mol		74.093 g/mol		2.0158 g/mol
$m_{initial}$	6.25 g						
$V_{initial}$			2 L				
d			1 g/mL				
$n_{initial}$	0.156 mol		100 mol		0 mol		0 mol
n_{change}	-0.156 mol		-0.312 mol		+0.156 mol		+0.156 mol
n_{final}	0 mol		100 mol		0.156 mol		0.156 mol
m_{final}							0.314 g *or* 314 mg

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.;

values in grey are not necessary for this calculation

Step 3: Calculate moles of Ca and H_2O available to react ($n_{initial}$)

$$n_{Ca-initial} = 6.25\ g\ Ca \times \frac{1mol}{40.078g} = 0.156\ mol\ Ca$$

$$n_{H_2O-initial} = 2\ L\ H_2O \times \frac{1000mL}{1L} \times \frac{1g}{1mL} \times \frac{1mol}{18.0152g} = 1 \times 10^2\ mol\ H_2O$$

Step 4: Identify the limiting reagent

2(0.156 mol) = 0.312 mol water are required to react with 0.156 mol Ca. Since there is much more water than this, the Ca will run out long before the water. Ca is therefore the limiting reagent.

Step 5: Calculate moles of H_2 produced (n_{final})

$$n_{H_2-final} = 0.156\ mol\ Ca \times \frac{1\ mol\ H_2}{1\ mol\ Ca} = 0.156\ mol\ H_2$$

Step 6: Calculate mass of H_2 produced (m_{final})

$$m_{H_2-final} = 0.156\ mol\ H_2 \times \frac{2.0158\ g}{1\ mol} = 0.314\ g\ H_2$$

$$m_{H_2-final} = 0.314\ g\ H_2 \times \frac{1000mg}{1g} = 314\ mg\ H_2$$

Step 7: Check your work

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

The question did not specify units, so 0.314 g or 314mg are both correct (since they are both reasonable choices).

The mass of H_2 produced is about twenty times smaller than that of Ca reacted; however, the molar mass of H_2 is about twenty times smaller than the molar mass of Ca. Since every mole of Ca reacted gives one mole of H_2, the answer seems reasonable.

While there is only 1 sig. fig. in the volume of water (2 L), water is not the limiting reagent so those significant figures are not relevant to the final answer.

5.	$CO_2(aq)$	+	$2\ NaOH(aq)$	→	$Na_2CO_3(aq)$
M	44.010 g/mol		39.9971 g/mol		105.989 g/mol
$c_{initial}$			0.04202 mol/L		
$V_{initial}$	25.00 mL		32.14 mL		
$n_{initial}$	0.0006753 mol		0.001351 mol		0 mol
n_{change}	-0.0006753 mol		-0.001351 mol		+ 0.0006753 mol
n_{final}	0 mol		0 mol		0.0006753 mol
$c_{initial}$	0.02701 mol/L				
$c_{initial}$	1.189 g/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 NaOH reacted ($n_{initial}$)

$c = \frac{n}{V}$ *therefore* $n = V \times c$

$$n_{NaOH-initial} = 32.14\ mL \times \frac{1L}{1000mL} \times \frac{0.04202mol}{1L} = 0.001351\ mol = 1.351 \times 10^{-3}\ mol$$

Step 4: Use mole ratio to calculate moles of CO_2 reacted ($n_{initial}$)

$$n_{CO_2-initial} = 1.351 \times 10^{-3}\ mol\ NaOH \times \frac{1\ mol\ CO_2}{2\ mol\ NaOH} = 6.753 \times 10^{-4}\ mol\ CO_2$$

Step 5*: Calculate concentration of initial CO_2 solution in mol/L ($c_{initial}$)

$$c_{CO_2-initial} = \frac{6.753 \times 10^{-4}\ mol\ CO_2}{25.00\ mL} \times \frac{1000\ mL}{1\ L} = 2.701 \times 10^{-2}\ \frac{mol\ CO_2}{L}$$

Step 6*: Calculate concentration of initial CO_2 solution in g/L ($c_{initial}$)

$$c_{CO_2-initial} = 2.701 \times 10^{-2}\ \frac{mol\ CO_2}{L} \times \frac{44.010\ g}{1\ mol} = 1.189\ \frac{g\ CO_2}{L}$$

Step 7: Check your work

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

* Steps 5 and 6 can be reversed by calculating the mass of CO_2 in the initial solution then dividing by its volume to give the concentration of CO_2.

6.	$Mg(s)$	+	$2\ HCl(aq)$	→	$MgCl_2(aq)$	+	$H_2(g)$
M	24.3050 g/mol		36.4606 g/mol		95.2104 g/mol		2.0158 g/mol
$m_{initial}$	185 mg						
$c_{initial}$			1.25 mol/L				
V			75 mL				
$n_{initial}$	0.00761 mol		0.094 mol		0 mol		0 mol
n_{change}	-0.00761 mol		-0.0152 mol		+0.00761 mol		+0.00761 mol
n_{final}	0 mol		0.079 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

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 available to react ($n_{initial}$)

$$n_{Mg-initial} = 185\ mg \times \frac{1\ g}{1000\ mg} \times \frac{1\ mol}{24.3050\ g} = 0.00761\ mol = 7.61 \times 10^{-3}\ mol$$

$$n_{HCl-initial} = 75\ mL\ \times \frac{1\ L}{1000\ mL} \times \frac{1.25\ mol}{1\ L} = 0.094\ mol = 9.4 \times 10^{-2}\ mol$$

Step 4: Identify the limiting reagent

The question strongly implies that Mg is the limiting reagent, but you should still check.

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 moles of HCl left after reaction is complete (n_{final})

$$n_{HCl-final} = n_{HCl-initial} - n_{HCl-reacted} = 0.094\ mol - 0.0152\ mol = 0.079\ mol$$

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

$$c_{HCl-final} = \frac{0.079\ mol\ HCl}{75\ mL} \times \frac{1000\ mL}{1\ L} = 1.0\ \frac{mol\ HCl}{L}$$

Step 7: Check your work

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

The final concentration of HCl is lower than the initial concentration (but still positive). This is consistent with some of the HCl being consumed by reaction with Mg and the volume of solution not changing.