Diploma of Health Sciences
Diploma of Science
SLE155 Chemistry for the Professional Sciences

This is NOT a practice exam.

The following are a wide range of questions to practise.

You might have seen similar questions before on quizzes because we use questions from the Blackman textbook.

The solutions will be discussed in your week 12 revision classes
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Your final exam will be 2 hours.
It will have 10 questions and be out of a total of 80 marks.
It will be worth 40% of your final mark for Chemistry for the Professional Sciences.

Also check the exam information on Moodle under week 13 for the data sheets which you will be given with your final examination.

Q1 Chemical reactions and stoichiometry

a) For the following balanced equations write a **net ionic equation**.
Make sure that you include **states** in your final **net ionic equation**.

Deduct ½ mark for each mistake up to a maximum of 1 mark [3 marks]

$2\ NaHCO_3(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l) + 2\ CO_2(g)$
$HCO_3^-(aq) + H^+(aq) \rightarrow CO_2(g) + H_2O(l)$
$2\ KOH(aq) + SnCl_2(aq) \rightarrow Sn(OH)_2(s) + 2\ KCl(aq)$
$Sn^{2+}(aq) + 2\ OH^-(aq) \rightarrow Sn(OH)_2(s)$
$2\ Cr(NO_3)_3\ (aq) + 3\ Na_2S\ (aq) \rightarrow Cr_2S_3\ (s) + 6\ NaNO_3\ (aq)$
$2\ Cr^{3+}(aq) + 3\ S^{2-}(aq) \rightarrow Cr_2S_3\ (s)$

b) Propane (C_3H_8) burns in oxygen to produce carbon dioxide and water.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

Suppose that 0.3818 moles of C_3H_8 and 1.718 moles of O_2 are allowed to react and this is the only reaction that occurs. Calculate the mass of carbon dioxide that can be formed from this mixture of reactants.

[5 marks]

0.3818 moles of C_3H_8 would require 1.909 moles of O_2 for complete reaction, therefore O_2 is the limiting reagent **1 mark**

Amount CO_2 produced will be 3/5 amount O_2 used **1 mark**

Amount CO_2 = $\frac{3 \times 1.718}{5}$ mol (= 1.0308 mol) **1 mark**

Molar mass CO_2 = $12.011 + (2 \times 15.999) = 44.009$ g mol^{-1} **1 mark**

(Must use atomic masses given on periodic table, without rounding off)

Using amount substance = $\frac{\text{mass}}{\text{molar mass}}$,

Mass CO_2 = $\frac{3 \times 1.718 \times 44.009}{5}$

= 45.36 g (45.3644772 g) **1 mark**

Deduct ½ mark for incorrect significant figures, ½ mark for incorrect or missing unit

Q1 Chemical reactions and stoichiometry

c) 66.7 mL of 18.0 molar sulfuric acid solution was dissolved in enough water to make 500.0 mL of solution. Calculate the molarity of the diluted mixture.

[2 marks]

Since this is a dilution, we can use $c_1 \times v_1 = c_2 \times v_2$

$c_2 = c_1 \times v_1 / v_2$

$c_2 = \dfrac{66.7 \times 18.0}{500}$

molarity of diluted solution is 2.40 mol L^{-1}

d) What volume in mL of 0.446 M $KMnO_4(aq)$ are required to react with 50.0 mL of 0.200 M $H_2C_2O_4(aq)$ in the presence of excess $H_2SO_4(aq)$? The reaction is:

$$2KMnO_4(aq) + 5H_2C_2O_4(aq) + 3H_2SO_4(aq) \rightarrow 2MnSO_4(aq) + 10CO_2(g) + 8H_2O(l) + K_2SO_4(aq)$$

[3 marks]

Molarity $= \dfrac{\text{amount of solute}}{\text{volume of solution in litres}}$

$\therefore$ amount $H_2C_2O_4$ = 0.200 × 0.0500 = 0.0100 mol

2 mole $KMnO_4(aq)$ requires 5 mole $H_2C_2O_4$ (from chemical equation)

$\therefore$ amount $KMnO_4(aq)$ required is 0.0040 mol

$\therefore$ volume $KMnO_4(aq)$ is $\dfrac{0.0040}{0.446}$ (from volume of solution = $\dfrac{\text{amount}}{\text{molarity}}$)

$\therefore$ volume 0.446 M $KMnO_4(aq)$ required is 0.008969 L

= 8.97 mL to 3 significant figures

Q1 Chemical reactions and stoichiometry

e) The reaction of powdered aluminium and iron(III) oxide :

$$2\ Al + Fe_2O_3 \rightarrow Al_2O_3 + 2\ Fe$$

produces so much heat that the iron that forms is molten. Because of this, the railways use the reaction when laying track to provide molten steel to weld the rails together. Suppose that, in one batch of reactants, 140.3 g of Al was mixed with 278.2 g of Fe_2O_3. Calculate the mass of iron that can be formed from this mixture of reactants.

[4 marks]

n(Al) = 140.3 / 26.982 = 5.200 mol
$M_r(Fe_2O_3)$ = (2 × 55.847) + (3 × 15.999) = 159.691 g mol^{-1} **1 mark**
$n(Fe_2O_3)$ = 278.2 / 159.691 = 1.739 mol
5.200 mol Al would require 2.600 mol Fe_2O_3, therefore Fe_2O_3 is the limiting reagent.
1 mark

1.739 mol Fe_2O_3 will form 3.479 mol Fe. **1 mark**

Using amount substance = $\frac{\text{mass}}{\text{molar mass}}$,

mass Fe = amount × molar mass
= 3.479 × 55.847
= 194.280
= 194.3 g Fe produced (4 significant figures) **1 mark**

f) What volume of 0.140 M $Ca(NO_3)_2(aq)$ solution is needed to react completely with 25.0 mL of 0.185 M $Na_3PO_4(aq)$ solution to give a precipitate of $Ca_3(PO_4)_2$?

$$2\ Na_3PO_4(aq) + 3\ Ca(NO_3)_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\ NaNO_3(aq)$$

[3 marks]

Molarity = $\frac{\text{amount of solute}}{\text{volume of solution in litres}}$
∴ amount Na_3PO_4 = 0.185 × 0.025 = 0.004625 mol **1 mark**
3 mole $Ca(NO_3)_2$ requires 2 mole Na_3PO_4 (from chemical equation)
∴ amount $Ca(NO_3)_2$ required is 0.006938 mol **1 mark**

∴ volume $Ca(NO_3)_2$ is $\frac{0.006938}{0.140}$ (from volume of solution = $\frac{\text{amount}}{\text{molarity}}$)

∴ volume 0.140 M $Ca(NO_3)_2$ required is 0.0496 L
= 49.6 mL to 3 significant figures **1 mark**

Answer can be in mL or L but should be 3 significant figures.

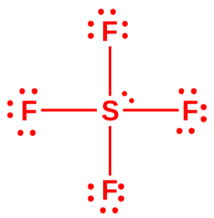
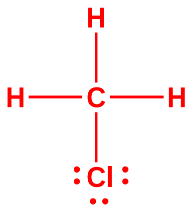
Q2 Chemical bonding and molecular structure

a) Write the Lewis structures for the following molecules:

SF_4
CH_3Cl

State if each molecule would be polar or non-polar and give your reasons.

[2 × 2 = 4 marks]

SF_4	CH_3Cl
 **½ mark bonds** **½ mark lone pairs** **Polar because the molecule is a see-saw shape (lone pair on S) and the polar S–F bonds do not cancel out. 1 mark**	 **½ mark bonds** **½ mark lone pairs** **Polar, because C–Cl bond is polar, and the shape of the molecule is such that the bond dipoles will not cancel out. 1 mark**

b) Both PF_3 and PF_5 are known compounds. NF_3 also exists, but NF_5 does not. Explain why there is no molecule with the formula NF_5.

[2 marks]

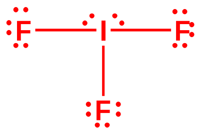
Some kind of statement regarding Lewis structure, such as
The Lewis structures of molecules with formula *X*F_3 have octets around the central atom and formal charges on all atoms of 0, making them stable. Compounds with formula *X*F_5 also have formal charges on all atoms of 0 but have five electron pairs associated with the inner atom. 1 mark

This is possible for phosphorus, a period 3 element that has *d* orbitals available for bonding. It is not possible for nitrogen, a period 2 element that does not have valence *d* orbitals available for bonding. 1 mark

Q2 Chemical bonding and molecular structure c) i) Draw the Lewis structure for IF_3
ii) Name the geometry of the electron pairs (do not draw)
iii) Draw the shape of the molecule indicating polar bonds with δ+ and δ–
iv) Name the shape of the molecule
v) Indicate if the molecule overall is polar or non-polar. Explain.

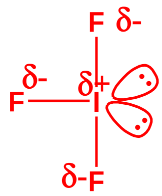
[2 + 1 + 2 + 1 + 2 = 8 marks]

i) Lewis structure drawing



1 mark bonds
1 mark all lone pairs
Indication of shape not necessary

iii) Molecular shape drawing



1 mark polar bonds
Lone pairs not necessary
1 mark attempt to show 3D shape

ii) **Electron pair geometry** name

trigonal bipyramid

iv) Name of **shape of molecule**

T-shaped

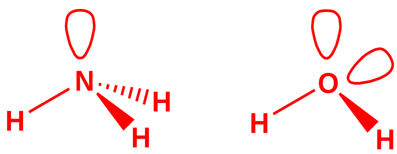
v) Polar or non-polar? Explain

The molecule will be polar **1 mark**

I–F bonds are polar, but the bond dipoles will not cancel out because of the shape of the molecule with lone pairs on I. **1 mark**

d) Draw the shapes of the NH_3 and H_2O molecules, then briefly explain why the bond angles are less than the ideal 109.5°.

[3 marks]



N and O each have four electron clouds around them, but N has one lone pair, and O has two lone pairs. Non-bonding electrons (lone pairs) take up more space than do bonding pairs, pushing the bonds closer together.

I mark each shape (2 marks total)
1 mark for saying lone pairs take up more room.

Q2 Chemical bonding and molecular structure

e) Which is the shorter bond length, C – N or C – C?
Explain the factors that influenced your choice.

[2 marks]

C – N < C – C	**1 mark**
Nitrogen has a smaller radius than carbon	**½ mark**
The C – N bond is polar	**½ mark**

(Should have more than one reason, since word "factors" was used in question)

Q4 Chemical Equilibrium

a) i) Write the equilibrium constant expression, K_c, for each of the following reactions in terms of their molar concentrations.

All or nothing, states not necessary [1 mark]

$2NaHSO_3(s) \rightleftharpoons Na_2SO_3(s) + H_2O(g) + SO_2(g)$	$K_c = [H_2O]_{eq}[SO_2]_{eq}$

ii) Arrange the reactions below in order of their increasing tendency to go towards completion.

All or nothing [1 mark]

$2CH_4(g) \rightleftharpoons C_2H_6(g) + H_2(g)$	$K_c = 9.5 \times 10^{13}$	Eqn 1
$CH_3OH(g) + H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$	$K_c = 3.6 \times 10^{20}$	Eqn 2
$H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$	$K_c = 2.0 \times 10^{9}$	Eqn 3

Reaction 2 goes furthest to completion, followed by reaction 1, then reaction 3

b) Consider the reaction:

$$N_2O(g) + NO_2(g) \rightleftharpoons 3NO(g) \qquad \Delta H^\circ = +155.7 \text{ kJ mol}^{-1}$$

In which direction will this equilibrium be shifted by the following changes?

i) Adding $N_2O(g)$

[1 mark]

Equilibrium will move to the right in the direction of products

(ii) Increasing the temperature of the reaction mixture

[1 mark]

Equilibrium will move to the right in the direction of products

iii) Addition of helium to the reaction mixture at constant volume

[1 mark]

No change

Q4 Chemical Equilibrium

c) At 460 °C, the reaction:

$SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$ has $K_c = 85.0$

A reaction flask at 460 °C contains these gases at the following concentrations:

$[SO_2] = 0.00250 \text{ mol L}^{-1}$ $[NO_2] = 0.00350 \text{ mol L}^{-1}$
$[NO] = 0.0250 \text{ mol L}^{-1}$ $[SO_3] = 0.0400 \text{ mol L}^{-1}$

i) State if the reaction mixture is at equilibrium. Show briefly how you arrived at your conclusion.

[3 marks]

$$K_c = \frac{[NO]_e[SO_3]_e}{[SO_2]_e[NO_2]_e} = 85.0$$

$$Q_c = \frac{[NO][SO_3]}{[SO_2][NO_2]} = \frac{(0.0250)\times(0.0400)}{(0.00250)\times(0.00350)}$$ **1 mark correct substitution**

= 114 **1 mark correct calculation**

1 mark

***$Q_c > K_c$* so reaction is not at equilibrium**

c) ii) If the reaction above is not at equilibrium, predict in which direction a spontaneous change would occur to get to equilibrium. Explain briefly how you arrived at your conclusion.

[2 marks]

Since the value of the reaction quotient Q_c for this system is larger than that of the equilibrium constant K_c, the system shifts to the left to make Q_c smaller and hence reach equilibrium.

2 marks (all or nothing) for both reason and conclusion

Q4 Chemical Equilibrium

d) Water dissociates to a very small amount into oxygen and hydrogen at high temperatures as in the equation below:

$$2\ H_2O(g) \quad\quad 2\ H_2(g) + O_2(g)$$

The value of K_c is 5.4×10^{-5} for the reaction at a temperature of 3290 K.

If a reaction vessel initially contains 0.500 mol L^{-1} H_2O (g), **calculate** the equilibrium concentrations of H_2 and O_2 at 3290 K.

[4 marks]

The balanced chemical equation is: 2 $H_2O(g)$ 2 $H_2(g) + O_2(g)$, so

$$K_c = \frac{[H_2]_e^2[O_2]_e}{[H_2O]_e^2}$$

The concentration table is: **1 mark**

	$[H_2O]$	$[H_2]$	$[O_2]$
Initial	**0.500**	**0**	**0**
Change	$-2x$	$+2x$	$+x$
Equilibrium	$0.500-2x$	$2x$	x

Substituting these equilibrium concentrations into the equilibrium constant expression gives:

$$K_c = \frac{(2x)^2(x)}{(0.500-2x)^2}$$ **1 mark**

$$= 5.4 \times 10^{-5}$$

Since K_c is small, we can assume that $(0.500 - 2x) \approx 0.500$. Thus:

$$\frac{4x^3}{(0.500)^2} = 5.4 \times 10^{-5}$$ **1 mark**

$4x^3 = (0.250) \times 5.4 \times 10^{-5} = 1.35 \times 10^{-5}$

$x^3 = 0.338 \times 10^{-5} = 3.38 \times 10^{-6}$
$x = 1.5 \times 10^{-2}$

Equilibrium concentration of H_2 = $2x = 3.0 \times 10^{-2}$ M **½ mark**
Equilibrium concentration of O_2 = 1.5×10^{-2} M **½ mark**

(or calculation making the assumptions without using the table)
Answers should be 2 significant figures.

Q5 Solutions and solubility

a) i) Write the equation, including states, for the dissolution of calcium fluoride, CaF_2.

All or nothing [1 mark]

$$CaF_2(s) \rightleftharpoons Ca^{2+}(aq) + 2F^{-}(aq)$$

ii) Write the expression for the solubility product, K_{sp}, for calcium fluoride.

All or nothing [1 mark]

$$K_{sp} = [Ca^{2+}][F^{-}]^2$$

b) The solubility product of lead(II) sulfate, $PbSO_4$, at 25 °C is 6.3×10^{-7}. Calculate the molar solubility of lead(II) sulfate at this temperature.
Hint: Assume that S mole of lead(II) sulfate dissociates into lead(II) ions and sulfate ions.
[3 marks]

$PbSO_4(s) \rightleftharpoons Pb^{2+}(aq) + SO_4^{2-}(aq)$

$K_{sp} = [Pb^{2+}][SO_4^{2-}]$

Let $[Pb^{2+}] = S$, therefore $[SO_4^{2-}] = S$ **1 mark**

$K_{sp} = S^2$

$6.3 \times 10^{-7} = S^2$

$S = \sqrt{6.3 \times 10^{-7}}$ **1 mark**

$S = 7.9 \times 10^{-4}$

Molar solubility of lead(II) sulfate is 7.9×10^{-4} mol L^{-1} **1 mark**
Must include unit mol L^{-1}

c) i) Define what is meant by the term "molality"

[1 mark]

$$\text{molality(b)} = \frac{\text{amount of substance (number of moles)}}{\text{mass of solvent expressed in kilogram}}$$

Molality is amount measured in moles of solute per kilogram of solvent.

Q5 Solutions and solubility

c) (ii) Benzene, C_6H_6, and water, H_2O, are immiscible. Explain briefly what this means. Also explain why they are immiscible in terms of the structures of the molecules and the forces of attraction between them.

[2 marks]

'Immiscible' means that benzene and water exist as two separate phases, regardless of the mole ratio of the mixture. **1 mark**

Water molecules are tightly linked to one another by hydrogen bonding. In benzene however, which is a nonpolar organic liquid, there are only weak attractive dispersion forces so water and benzene will not mix. **1 mark**

d) Calculate the mass of sodium chloride, NaCl, (a salt found in seawater) which would have to be dissolved in 500 g of water to give a solution of molality
0.150 mol kg^{-1}.

[2 marks]

Using $\text{molality(b)} = \dfrac{\text{amount of substance (number of moles)}}{\text{mass of solvent expressed in kilogram}}$

Amount of NaCl = molality × mass of solvent
Amount of NaCl = 0.150 × 0.500
= 0.0750 mol **1 mark**

Mass NaCl = 0.0750 × 58.44
= 4.38 g **1 mark**

Must include unit, should be 3 significant figures

e) Explain why the vapour pressure of a solvent above a solution is less than that above the pure solvent at the same temperature.

[2 marks]

There is an equilibrium between the solvent molecules at the surface and those in the vapour phase. **1 mark**

For pure solvent, all the particles at the surface are solvent, so a maximum amount of solvent is able to vaporise.
In a solution, some of the positions, as determined by the mole fractions, are taken up by solute particles, so fewer solvent particles are available to vaporise. **1 mark**

Q5 Solutions and solubility

j) A student prepared a buffer solution made up of NH_3 and NH_4Cl. Determine the pH of the buffer when $[NH_3] = 0.25$ mol L^{-1} and $[NH_4Cl] = 0.45$ mol L^{-1}.
Ammonium ion has a K_a of 5.56×10^{-10} and a pK_a of 9.26 (not usually given)

[3 marks]

$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$$

$$[H_3O^+] = K_a \times \frac{[NH_4^+]}{[NH_3]}$$ **1 mark**

$$[H_3O^+] = 5.56 \times 10^{-10} \times \frac{0.45}{0.25}$$

$$= 1.0 \times 10^{-9}\ M$$ **1 mark**

$$pH = -\log_{10}[H_3O^+]$$

$$= -\log_{10}(1.0 \times 10^{-9})$$

$$= 9.00$$ **1 mark**

OR Henderson's equation could be used if they remember it (but not given in data).

$$pH = pK_a + \log_{10}\frac{[\text{conjugate base}]}{[\text{conjugate acid}]}$$ **1 mark**

$$pH = 9.26 + \log\frac{0.25}{0.45}$$ **1 mark**

$$= 9.00$$ **1 mark**

Q6 Oxidation and reduction

a) i) Assign oxidation numbers to all of the atoms in the following compounds or ions:

½ mark each, must have correct sign (+ or –) [2 marks]

$Cr_2O_7^{2-}$	**Atom**	**Oxidation number**	**Atom**	**Oxidation number**
	Cr	**+6**	**O**	**–2**
MnO_2	**Atom**	**Oxidation number**	**Atom**	**Oxidation number**
	Mn	**+4**	**O**	**–2**

ii) Balance the following half equation for reaction occurring in acidic solution.

$$NO_3^-(aq) \rightarrow NH_4^+(aq)$$

Show your working. Indicate whether the reaction is oxidation or reduction.

[2 marks]

$$NO_3^-(aq) \rightarrow NH_4^+(aq) + 3H_2O(l)$$
$$10H^+(aq) + NO_3^-(aq) \rightarrow NH_4^+(aq) + 3H_2O(l)$$
$$8e^- + 10H^+(aq) + NO_3^-(aq) \rightarrow NH_4^+(aq) + 3H_2O(l)$$

Reduction reaction

1 mark for balancing with H_2O and H^+, ½ mark for electrons, ½ mark for reduction. States do not have to be shown.

iii) Balance the following half equation for reaction occurring in basic solution.

$$H_4IO_6^-(aq) \rightarrow I_2(s)$$

Show your working. Indicate whether the reaction is oxidation or reduction.

[3 marks]

$$2\ H_4IO_6^-(aq) \rightarrow I_2(s)$$
$$2\ H_4IO_6^-(aq) \rightarrow I_2(s) + 12\ H_2O(l)$$
$$16\ H^+(aq) + 2\ H_4IO_6^-(aq) \rightarrow I_2(s) + 12\ H_2O(l)$$
$$14e^- + 16\ H^+(aq) + 2\ H_4IO_6^-(aq) \rightarrow I_2(s) + 12\ H_2O(l)$$

Add 16 OH^- to both sides

$$14e^- + 16\ H^+(aq) + 16\ OH^-(aq) + 2\ H_4IO_6^-(aq) \rightarrow I_2(s) + 12\ H_2O(l) + 16\ OH^-(aq)$$
$$14e^- + 4\ H_2O(l) + 2\ H_4IO_6^-(aq) \rightarrow I_2(s) + 16\ OH^-(aq)$$

Reduction reaction

½ mark for balancing I, 1 mark for balancing with H_2O and H^+, ½ mark for electrons, ½ mark for reduction. ½ mark for OH^- and for cancelling H_2O. States do not have to be shown.

Q6 Oxidation and reduction

$Au^{3+}(aq)\|Au(s)$	$Fe^{3+}(aq)\|Fe^{2+}(aq)$	$Cd^{2+}(aq)\|Cd(s)$	$Fe^{2+}(aq)\|Fe(s)$	$Zn^{2+}(aq)\|Zn(s)$
+1.42 V	+0.77 V	–0.40 V	–0.44 V	–0.76 V

b) i) Use the data in the table above to calculate the standard cell potential for the following reaction at 25 °C.

$$Cd^{2+}(aq) + Fe(s) \rightarrow Cd(s) + Fe^{2+}(aq)$$

[1 mark]

E°cell = E°cathode – E°anode
= –0.40 – (–0.44)
= + 0.04 V (must include + sign)

c) i) For the following reaction, identify the substance being oxidised, and the substance being reduced.
Assign oxidation numbers to the atoms being oxidised and/or reduced.

$$4\ Au + 16\ CN^- + 3\ O_2 + 6\ H_2O \rightarrow 4\ [Au(CN)_4]^- + 12\ OH^-$$

1 mark correct oxid/red, 1 mark all oxidation numbers correct [2 marks]

Substance oxidised	Substance reduced
Au	**O_2**
Oxidation number of atoms being oxidised on each side of the equation **Au = 0 on LHS, +3 on RHS**	Oxidation number of atoms being reduced on each side of the equation **O = 0 on LHS, –2 on RHS**

ii) Balance the following half equation for reaction occurring in acidic solution.

$$BiO_3^-(aq) \rightarrow Bi^{3+}(aq)$$

Show your working. Indicate whether the reaction is oxidation or reduction.

[2 marks]

$BiO_3^-(aq) \rightarrow Bi^{3+}(aq) + 3H_2O(aq)$
$6H^+(aq) + BiO_3^-(aq) \rightarrow Bi^{3+}(aq) + 3H_2O(aq)$
$2e^- + 6H^+(aq) + BiO_3^-(aq) \rightarrow Bi^{3+}(aq) + 3H_2O(aq)$
Reduction reaction
1 mark for balancing with H_2O and H^+, ½ mark for electrons, ½ mark for reduction.
States should be shown.

Q6 Oxidation and reduction

iii) Balance the following half equation for reaction occurring in basic solution.

$$SO_2Cl_2(l) \rightarrow SO_3^{2-}(aq) + Cl^-(aq)$$

Show your working. Indicate whether the reaction is oxidation or reduction.

[3 marks]

$SO_2Cl_2(l) \rightarrow SO_3^{2-}(aq) + 2\,Cl^-(aq)$
$H_2O(aq) + SO_2Cl_2(l) \rightarrow SO_3^{2-}(aq) + 2\,Cl^-(aq)$
$H_2O(aq) + SO_2Cl_2(l) \rightarrow SO_3^{2-}(aq) + 2\,Cl^-(aq) + 2\,H^+(aq)$
$H_2O(aq) + SO_2Cl_2(l) + 2\,e^- \rightarrow SO_3^{2-}(aq) + 2\,Cl^-(aq) + 2\,H^+(aq)$
$H_2O(aq) + SO_2Cl_2(l) + 2\,e^- + 2\,OH^-(aq)$
$\rightarrow SO_3^{2-}(aq) + 2\,Cl^-(aq) + 2\,H^+(aq) + 2\,OH^-(aq)$
$SO_2Cl_2(l) + 2\,e^- + 2\,OH^-(aq) \rightarrow SO_3^{2-}(aq) + 2\,Cl^-(aq) + H_2O(aq)$
Reduction
1 mark for balancing with H_2O and H^+, ½ mark for electrons, ½ mark for reduction.
½ mark for OH^-, ½ marks for cancelling H_2O. States should be shown.

iv) Using half-reactions and showing your working, balance the following equation in basic solution.

$$BH_4^-(aq) + MnO_4^-(aq) \rightarrow H_2BO_3^-(aq) + MnO_2(s)$$

Show your working. Indicate whether each half reaction is oxidation or reduction.

[5 marks]

Oxidation reaction
$BH_4^-(aq) \rightarrow H_2BO_3^-(aq)$
$3\,H_2O(l) + BH_4^-(aq) \rightarrow H_2BO_3^-(aq)$
$3\,H_2O(l) + BH_4^-(aq) \rightarrow H_2BO_3^-(aq) + 8\,H^+(aq)$
$3\,H_2O(l) + BH_4^-(aq) \rightarrow H_2BO_3^-(aq) + 8\,H^+(aq) + 8\,e^-$
Multiply by 3
$9\,H_2O(l) + 3\,BH_4^-(aq) \rightarrow 3\,H_2BO_3^-(aq) + 24\,H^+(aq) + 24\,e^-$
1 mark for balancing with H_2O and H^+, ½ mark for electrons **total 1½ marks**

Reduction reaction
$MnO_4^-(aq) \rightarrow MnO_2(s)$
$MnO_4^-(aq) \rightarrow MnO_2(s) + 2\,H_2O(l)$
$4\,H^+(aq) + MnO_4^-(aq) \rightarrow MnO_2(s) + 2\,H_2O(l)$
$3\,e^- + 4\,H^+(aq) + MnO_4^-(aq) \rightarrow MnO_2(s) + 2\,H_2O(l)$
Multiply by 8
$24\,e^- + 32\,H^+(aq) + 8\,MnO_4^-(aq) \rightarrow 8\,MnO_2(s) + 16\,H_2O(l)$
1 mark for balancing with H_2O and H^+, ½ mark for electrons **total 1½ marks**

Add the two equations and cancel H^+, H_2O and electrons
1 mark for adding the two equations correctly **1 mark**
$8\,H^+(aq) + 3\,BH_4^-(aq) + 8\,MnO_4^-(aq) \rightarrow 3\,H_2BO_3^-(aq) + 8\,MnO_2(s) + 7\,H_2O(l)$
Add 8 $OH^-(aq)$ to both sides
$8\,OH^-(aq) + 8\,H^+(aq) + 3\,BH_4^-(aq) + 8\,MnO_4^-(aq) \rightarrow 3\,H_2BO_3^-(aq) + 8\,MnO_2(s) + 7\,H_2O(l) + 8\,OH^-(aq)$
$H_2O(l) + 3\,BH_4^-(aq) + 8\,MnO_4^-(aq) \rightarrow 3\,H_2BO_3^-(aq) + 8\,MnO_2(s) + 8\,OH^-(aq)$

½ mark for OH^-, ½ marks for cancelling H_2O. States should be shown.
1 mark total