
CHEM1011 Chemistry A

Sample Exam Questions – CHEM1011

NOTES:

1. These sample questions were adapted from past exam papers of previous courses. Originally the questions were worth 10 marks each, however in some cases a sample question is shorter than a typical exam question because unsuitable parts of the old question have been cut out.
2. From 2015 onwards, the structure of the written exam questions have changed to be worth 5 marks each (ie: half as long as they are in this document). Please take this into account during your exam preparation.
3. This list of sample questions is NOT meant to simulate or define the size or scope of your final exam paper. It is intended merely to provide some useful study material. For details on the size and structure of your exam paper, see the specific information about the exam paper published on Moodle. In particular see and read the front page of your exam paper, that is published on Moodle.
4. Some answers to these sample questions, that I will publish, may include extra explanations to help you to understand the answers. Your exam answers do not need to contain more material than is required to answer the questions.
5. A data sheet at the end of these sample questions contains data that might be needed to answer some questions.

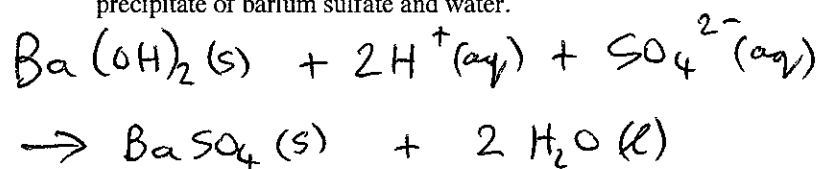
Steve Yannoulatos (Student Support Manager, Chemistry).

<p>For details of the instructions on the front page of your exam paper, see the the document published on Moodle.</p>

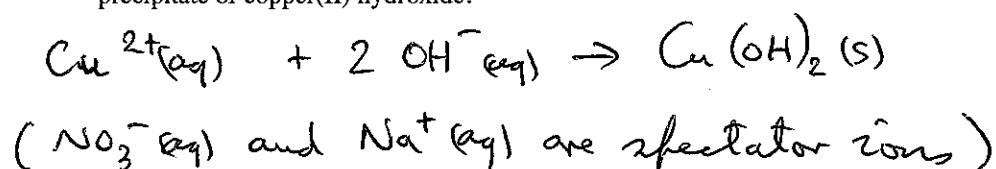
QUESTION 1

Write balanced ionic equations for the following chemical reactions. Include states of matter, but do not include spectator ions.

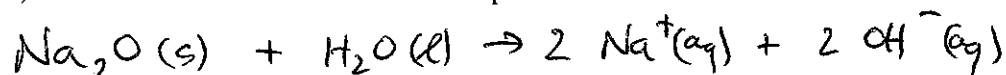
- (i) Solid barium hydroxide reacts with an excess of 0.1 M sulfuric acid to produce a precipitate of barium sulfate and water.



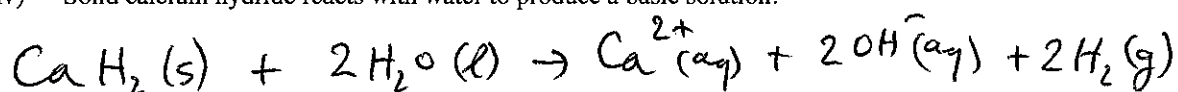
- (ii) A solution of copper(II) nitrate is added to a solution of sodium hydroxide to produce a precipitate of copper(II) hydroxide.



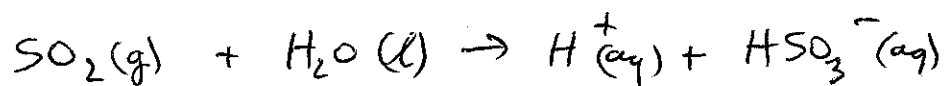
- (iii) Solid sodium oxide reacts with water to produce a basic solution.



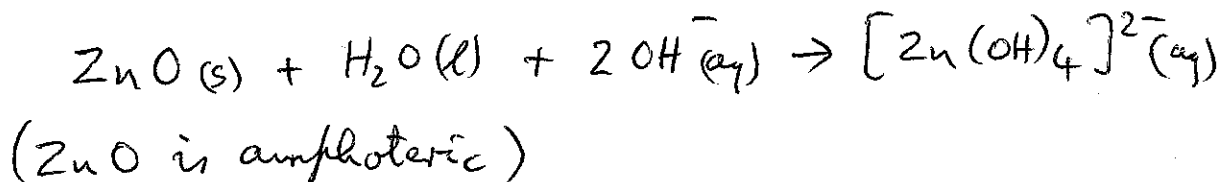
- (iv) Solid calcium hydride reacts with water to produce a basic solution.



- (v) Sulfur dioxide gas reacts with water to produce an acidic solution.



- (vi) Solid zinc oxide reacts with sodium hydroxide solution to give a colourless solution.



QUESTION 2

- (a) Calculate the frequency of electromagnetic radiation which has a wavelength of 532 nm.

$$c = f\lambda \quad \therefore f = \frac{c}{\lambda} = \frac{2.99 \times 10^8}{532 \times 10^{-9}} \text{ s}^{-1}$$

$$= 5.62 \times 10^{14} \text{ s}^{-1}$$

- (b) Calculate the energy of one photon of electromagnetic radiation which has a frequency of 90.0 MHz.

$$E = hf = (6.626 \times 10^{-34})(90.0 \times 10^6) \text{ J}$$

$$= 5.96 \times 10^{-26} \text{ J}$$

- (c) Use the Rydberg equation: $\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$

where $R_H = 1.097 \times 10^7 \text{ m}^{-1}$ to calculate the wavelength of light required to cause a hydrogen atom to undergo a transition from the $n = 1$ electronic state to the $n = 3$ electronic state.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) = (1.097 \times 10^7) \left(\frac{1}{1^2} - \frac{1}{3^2} \right)$$

$$= 9751111 \text{ m}^{-1}$$

$$\therefore \lambda = 1.026 \times 10^{-7} \text{ m} (= 102.6 \text{ nm})$$

- (d) (i) Write down the values of the n and l quantum numbers for these orbitals:

$$3s \quad n = 3 \quad l = 0$$

$$4d \quad n = 4 \quad l = 2$$

- (ii) Write down all the allowed values of the quantum number m_l for the orbitals in a $3p$ sublevel.

$$3p \quad \dots \quad l = 1$$

$$m_l = +1, 0, -1$$

- (e) In the upper atmosphere chloromethane can be decomposed by light according to the reaction:



The energy required for this reaction is 428 kJ mol^{-1} . If each molecule of chloromethane is decomposed by one photon, what is the maximum wavelength that photon can have?

$$\text{MINIMUM ENERGY NEEDED} = 428 \text{ kJ mol}^{-1}$$

$$\therefore \text{ " " " " PER MOLECULE} = \frac{428 \times 10^3}{6.022 \times 10^{23}} \text{ J}$$

$$= 7.1073 \times 10^{-19} \text{ J}$$

$$E = hf = \frac{hc}{\lambda}$$

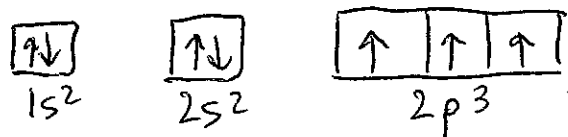
$$\therefore \text{MAX } \lambda = \frac{hc}{E} = \frac{(6.626 \times 10^{-34})(2.99 \times 10^8)}{(7.1073 \times 10^{-19})} \text{ m}$$

$$= 2.79 \times 10^{-7} \text{ m}$$

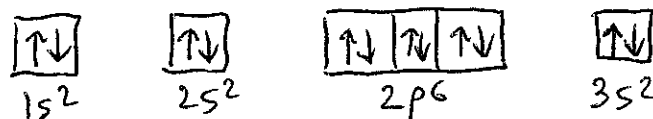
QUESTION 3

- (a) Using 'arrows-in-labeled-boxes' notation, write the complete ground state electronic configurations of the following atoms and ions.

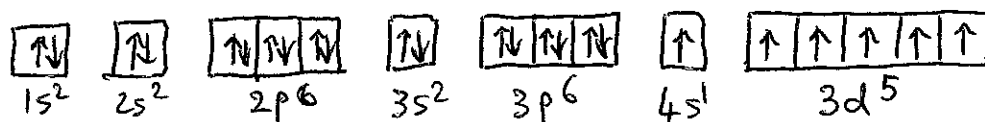
(i) N



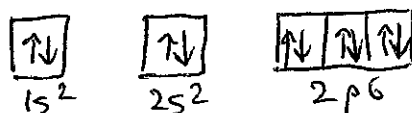
(ii) Mg



(iii) Cr



(iv) F^-



(v) V^+



- (b) Write down all the species from part (a) which are paramagnetic.

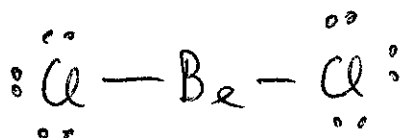
N, Cr, V^+ (all have at least one unpaired electron)

- (c) Write down all the species from part (a) which are isoelectronic with a noble gas.

F^- (ISOELECTRONIC WITH Ne)

- (d) For each of the following molecules draw a Lewis diagram, describe the shape of the molecule, and describe whether the molecule has a permanent dipole moment.

(i) $BeCl_2$



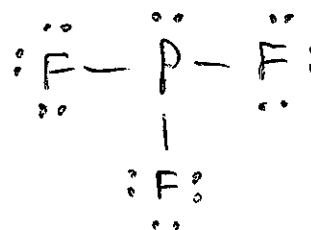
Shape:

LINEAR

Polar?

no

(ii) PF_3



Shape:

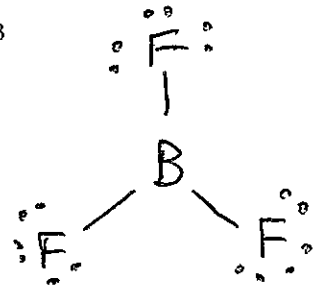
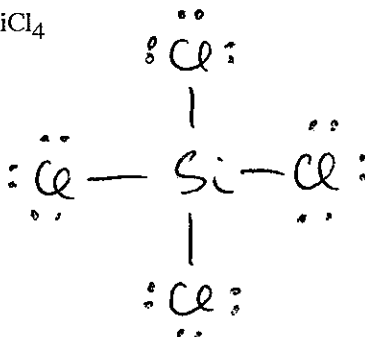

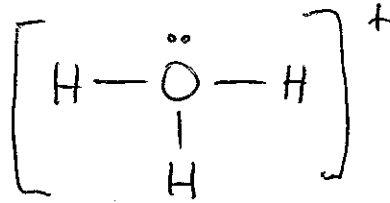
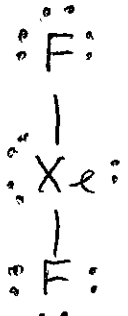
PYRAMIDAL

Polar?

yes

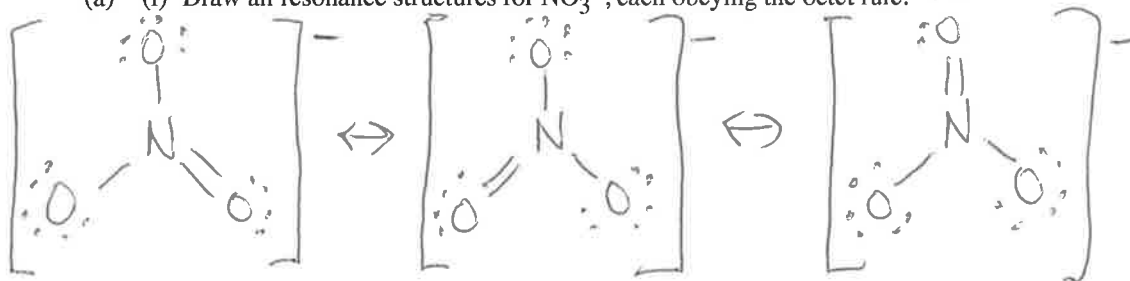
QUESTION 4

For each of the following species draw a Lewis diagram, describe the shape of the species, and describe the hybridisation at the central atom.

<p>(i) BF_3</p>  <p>Shape: TRIGONAL-PLANAR</p> <p>Hybridisation: sp^2</p>	<p>(ii) SiCl_4</p>  <p>Shape: TETRAHEDRAL</p> <p>Hybridisation: sp^3</p>
<p>(iii) CO_2</p>  <p>Shape: LINEAR</p> <p>Hybridisation: sp</p>	<p>(iv) H_3O^+</p>  <p>Shape: PYRAMIDAL</p> <p>Hybridisation: sp^3</p>
<p>(v) XeF_2</p>  <p>Shape: LINEAR</p> <p>Hybridisation: sp^3d</p>	

QUESTION 5

- (a) (i) Draw all resonance structures for
- NO_3^-
- , each obeying the octet rule.



- (ii) Write down the hybridization of the nitrogen atom in
- NO_3^-
- .

$$sp^2$$

- (iii) What type of orbitals on the nitrogen atom and oxygen atoms in
- NO_3^-
- overlap to produce the
- π
- bond in this ion?

The 2p orbitals on N and O overlap sideways to give the π bond.

- (iv) Would you expect the N-O bond in
- NO_3^-
- to be longer or shorter than the N-O bond in hydroxylamine,
- $\text{H}_2\text{N-O-H}$
- ?

Shorter. The resonance hybrid of NO_3^- is $\text{O}=\text{N}(\text{O}^-)-\text{O}$ and it has partial double-bonding which makes the N-O bonds in NO_3^- shorter and stronger than the single N-O bond in $\text{H}_2\text{N-O-H}$.

- (b) The table below shows the normal boiling points of the group 17 hydrides.

Compound	HF	HCl	HBr	HI
Normal boiling point / °C	20	-85	-67	-35

- (i) What is the dominant type of intermolecular force which determines the normal boiling points of the compounds HCl, HBr, and HI?

The b.p. trend follows the trend of the electron count for these three compounds, and this electron count is proportional to dispersion forces.

- (ii) What is the dominant type of intermolecular force which determines the normal boiling point of HF?

HF has hydrogen bonding, in addition to weaker dispersion forces. The stronger hydrogen bonding dominates in determining the b.p. of HF.

- (c) Calculate the pressure of 100.0 mol of xenon in a 15.0 L vessel at 300 K using the van der Waals

equation: $(P + a(\frac{n}{V})^2)(V - nb) = nRT$ where, for xenon, $a = 4.19 \text{ atm L}^2 \text{ mol}^{-2}$ and $b = 0.0511 \text{ L mol}^{-1}$

$$\left(P + 4.19 \left(\frac{100}{15}\right)^2\right)(15 - (100)(0.0511)) = (100)(0.0821)(300)$$

$$(P + 186.2)(9.89) = 2463$$

$$9.89P = 621.26$$

$$P = 62.8 \text{ atm.}$$

QUESTION 6

In the week 3 laboratory experiment the concentration of a sodium hydroxide solution was determined by titration with a standard potassium hydrogenphthalate (KHP) solution.

- (a) Calculate the concentration of KHP standard solution obtained by dissolving 5.10 g of KHP (molar mass $204.22 \text{ g mol}^{-1}$) into 250.0 mL of water, in a volumetric flask.

$$\begin{aligned} \text{Moles KHP} &= (5.10 / 204.22) = 0.02497 \text{ mol.} \\ \therefore [\text{KHP}] &= \frac{1000}{250} \times 0.02497 \\ &= 0.09999 \text{ M.} \end{aligned}$$

- (b) Four 25.00 mL samples of the KHP standard solution were each titrated with the unknown sodium hydroxide solution. The results are given in the table below.

Titration	Initial Reading ($\pm 0.05 \text{ mL}$)	Final Reading ($\pm 0.05 \text{ mL}$)	Titre ($\pm 0.1 \text{ mL}$)	Used in calculation? (Yes/No)
1	0.05	35.55	35.5	NO
2	0.40	35.50	35.1	YES
3	0.45	35.55	35.1	YES
4	1.60	36.70	35.1	YES

- (i) Complete the 'Titre' data in the table.
- (ii) Make use of data in the table to calculate the concentration of the sodium hydroxide solution, assuming that KHP is a monoprotic acid.

$$\begin{aligned} &\text{Use titrations 2, 3, 4 to give average titre of } 35.1 \text{ mL.} \\ &25.00 \text{ mL of } 0.09999 \text{ M KHP contains } \frac{25.00}{1000} \times 0.09999 \text{ mol} \\ &= 2.497 \times 10^{-3} \text{ mol KHP.} \\ \therefore \text{Amount NaOH in } 35.1 \text{ mL titre} &= 2.497 \times 10^{-3} \text{ mol NaOH.} \\ \therefore [\text{NaOH}] &= \left(\frac{1000}{35.1} \right) \times 2.497 \times 10^{-3} \\ &= 0.0712 \text{ M.} \end{aligned}$$

- (iii) Complete the 'Used in calculation?' data in the table, to indicate which titres were used in the calculation of the sodium hydroxide solution. Comment briefly on how you decided which titres to use in the calculation.

Titration data 2, 3, 4 were used in the calculation because they were consistent, whereas titration 1 was not.

$$V_1 = 340 - 20 = 320 \text{ mL} \quad \boxed{\text{GAS}} \rightarrow \boxed{\text{GAS}} \quad V_2 = 1000 - 20 = 980 \text{ mL}$$

QUESTION 7

- (a) A 340 mL cylinder contained xenon gas at 123 kPa and 20.0 mL of liquid 2-butanone at 25 °C.

(Data: the vapour pressure of 2-butanone is 13.3 kPa at 25 °C.)

- (i) The volume of the container is increased to 1.00 L keeping the temperature constant. What is the total pressure in the container?

Consider the Xe alone:- *The liquid butanone will have the same vapour pressure at the same temperature.*

$$P_1 V_1 = P_2 V_2$$

$$(123)(320) = P_2(980)$$

$$P_2 = 40.16 \text{ kPa}$$

$$\therefore P_{\text{butanone}} = 13.3 \text{ kPa}$$

$$\therefore P_{\text{Tot}} = P_{\text{butanone}} + P_{\text{Xe}} = 13.3 + 40.16 = 53.5 \text{ kPa}$$

- (ii) 100 mL of argon gas (measured at 101 kPa and 25 °C) is then added to the 1.00 L container.

What is the total pressure in the container, assuming the volume of liquid is still 20.0 mL?

Consider Ar alone:-

$$P_1 V_1 = P_2 V_2$$

$$(101)(100) = P_2(980)$$

$$P_2 = 10.306 \text{ kPa}$$

$$P_{\text{Tot}} = P_{\text{butanone}} + P_{\text{Xe}} + P_{\text{Ar}}$$

$$= 13.3 + 40.16 + 10.306$$

$$= 63.8 \text{ kPa}$$

- (b) At 25 °C the vapour pressure of pure benzene (C_6H_6 , molar mass = 78.1 g mol⁻¹) is 95.1 Torr and that of pure toluene ($\text{C}_6\text{H}_5\text{CH}_3$, molar mass = 93.1 g mol⁻¹) is 28.4 Torr.

- (i) Calculate the total vapour pressure of a solution containing 40.0 g of benzene and 20.0 g of toluene at 25 °C, assuming ideal behaviour of the solution.

$$n_B = 40.0 / 78.1 = 0.512 \text{ mol}$$

$$n_T = 20.0 / 93.1 = 0.2148 \text{ mol}$$

$$\text{Mole fraction } X_B = (0.512) / (0.512 + 0.2148) = 0.7044$$

$$X_T = 1 - 0.7044 = 0.2956$$

$$P_{\text{Tot}} = X_B P_B^\circ + X_T P_T^\circ$$

$$= (0.7044)(95.1) + (0.2956)(28.4)$$

$$= 66.988 + 8.394$$

$$= 75.4 \text{ Torr}$$

- (ii) Calculate the partial pressure of benzene in the gas phase above the mixture.

$$P_B = X_B P_B^\circ$$

$$= (0.7044)(95.1)$$

$$= 67.0 \text{ Torr}$$

- (iii) A sample of the vapour in equilibrium with the mixture was isolated and cooled until the vapour completely condensed into liquid. Calculate the mole fraction of benzene in this liquid.

$$X_B (\text{vap}) = \frac{P_B}{P_{\text{Tot}}} = \frac{67.0}{75.4} = 0.89$$

QUESTION 8

(a) Complete the following table, writing "none" where appropriate:

(NH_2^-) IS TOO STRONG A BASE TO EXIST IN WATER

Conjugate Acid	Species	Conjugate Base
NH_4^+	NH_3	NONE
NONE	$\text{CH}_3\text{CH}_2\text{COOH}$	$\text{CH}_3\text{CH}_2\text{COO}^-$
H_2CO_3	HCO_3^-	CO_3^{2-}

(b) Supply the missing information in the table below:

	$[\text{H}^+]$ (mol L ⁻¹)	pH	$[\text{OH}^-]$ (mol L ⁻¹)	pOH
(i)	3.4×10^{-4}	3.47	2.95×10^{-11}	10.53
(ii)	6.3×10^{-9}	8.2	1.58×10^{-6}	5.8

(c) Write down the oxidation number of the element indicated in the species shown above it.

Species:	Cu_2O	SO_3^{2-}	P_4	$\text{S}_2\text{O}_3^{2-}$
Element:	Cu	S	P	S
Oxidation number:	+1	+4	0	+2

QUESTION 9

(a) Write equilibrium expression in terms of K_p for each of the following:



$$K_p = p_{\text{O}_2}$$



$$K_p = \frac{(p_{\text{N}_2})(p_{\text{O}_2})}{(p_{\text{NO}})^2}$$

(b) The equilibrium constant K_p for $2 \text{NO(g)} + \text{O}_2\text{(g)} \rightleftharpoons 2 \text{NO}_2\text{(g)}$ at 298K is 1.6×10^{12} (standard state = 1 bar).

Calculate the K_p for $\text{NO}_2\text{(g)} \rightleftharpoons \text{NO(g)} + \frac{1}{2} \text{O}_2\text{(g)}$ at the same temperature.

This equation is the reverse of the first one, with the coefficients halved.

$$\therefore K_{p \text{ NEW}} = \frac{1}{\sqrt{K_{p \text{ OLD}}}} = \frac{1}{\sqrt{1.6 \times 10^{12}}} = 7.9 \times 10^{-7}$$

(c) Calculate whether a net reaction will occur and, if so, in which direction, for the following initial conditions: a 1.0 L flask containing the gases NO, O₂ and NO₂, each with a partial pressure of 0.010 bar at 298 K.

$$Q = \frac{(p_{\text{NO}_2})^2}{(p_{\text{NO}})^2 (p_{\text{O}_2})} = \frac{(0.010)^2}{(0.010)^2 (0.010)}$$

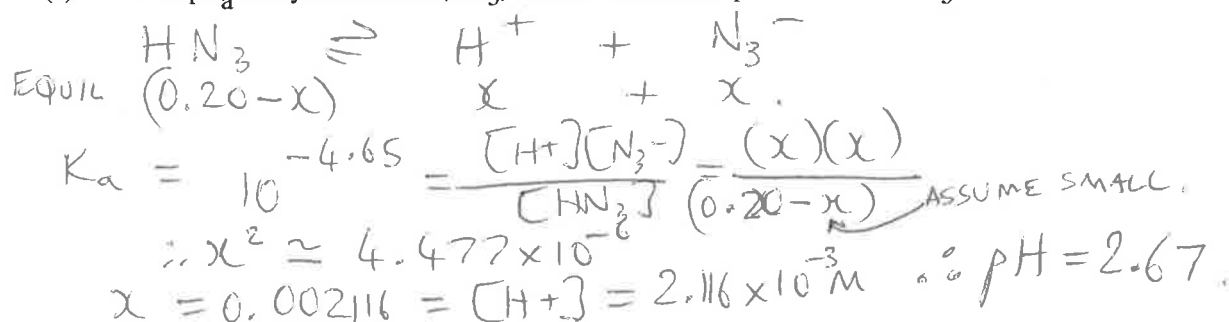
$$= 100.$$

$$\therefore Q < K_p$$

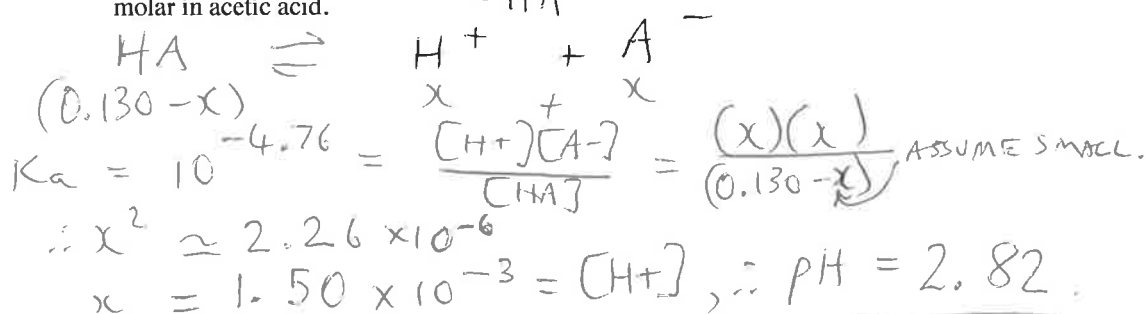
\therefore A net reaction will occur in which p_{NO} and p_{O_2} will fall, whilst p_{NO_2} rises.

QUESTION 10

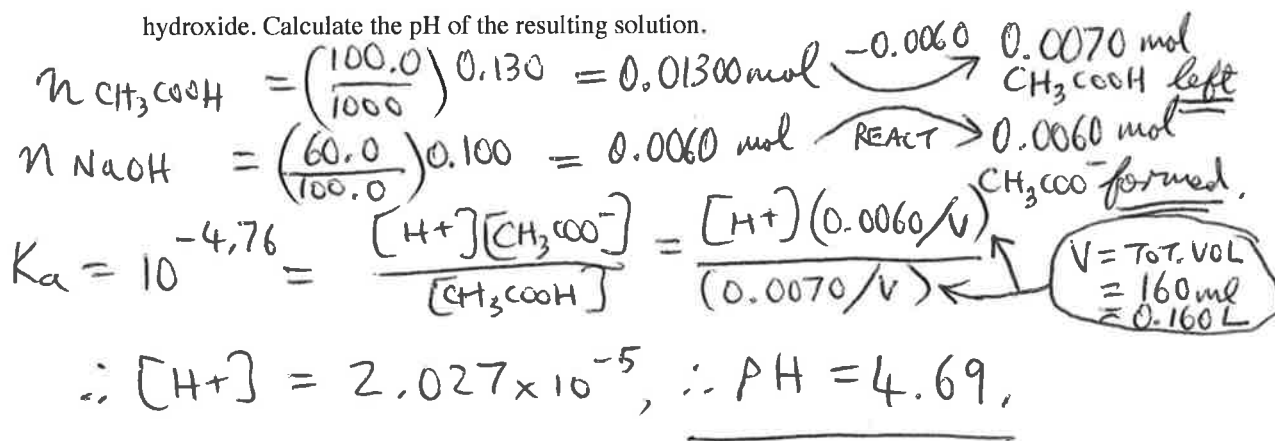
- (a) The pK_a for hydrazoic acid (HN_3) is 4.65. What is the pH of a 0.20 M HN_3 solution?



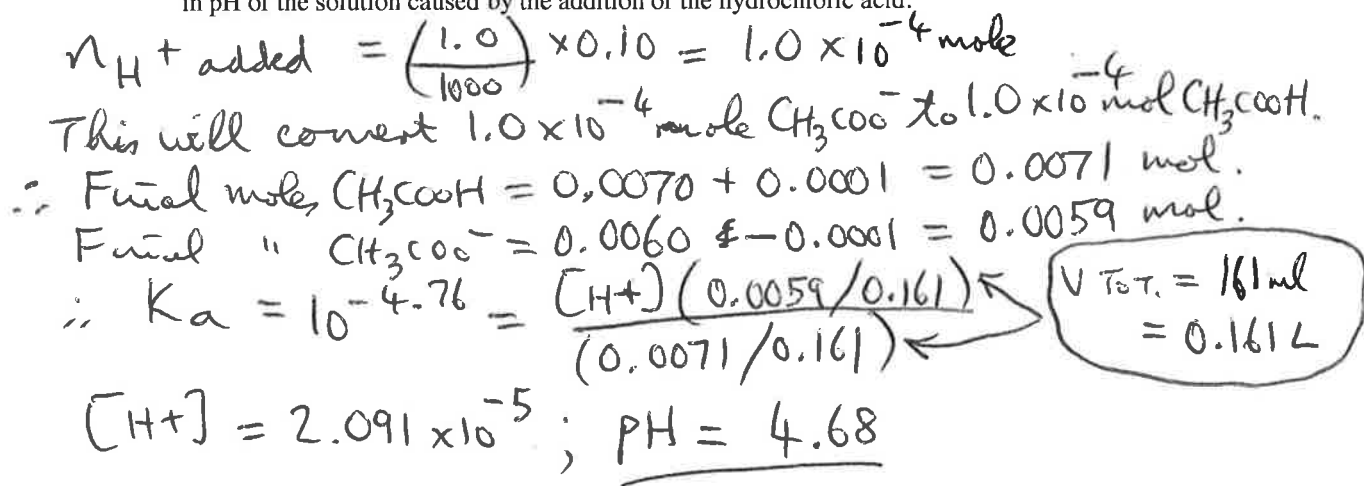
- (b) (i) The pK_a for acetic acid (CH_3COOH) is 4.76. Calculate the pH of a solution that is 0.130 molar in acetic acid.



- (ii) 100.0 mL of 0.130 molar acetic acid is mixed with 60.0 mL of 0.100 molar sodium hydroxide. Calculate the pH of the resulting solution.



- (iii) 1.0 mL of 0.10 mol L^{-1} HCl(aq) is added to the solution from part (ii). Calculate the change in pH of the solution caused by the addition of the hydrochloric acid.



QUESTION 11

- (a) The pK_a for acetic acid (CH_3COOH) is 4.76. Calculate the pH of a solution that is 0.130 molar in acetic acid and 0.240 molar in sodium acetate.

A mixture of CH_3COOH and $\text{Na}^+\text{CH}_3\text{COO}^-$ will be a buffer solution, as it is a mixture of a weak acid and its conjugate base.

Use the Henderson-Hasselbalch equation:

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log\left(\frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}\right) \\ &= 4.76 + \log\left(\frac{0.240}{0.130}\right) \\ &= \underline{5.03} \end{aligned}$$

- (b) Briefly describe TWO methods that could be employed to prepare the buffer in part (a) in the laboratory.

To make the buffer described in (a)...

1. Dissolve 0.240 mole sodium acetate and 0.130 mole of acetic acid in water, and make up to a final volume of 1.00 L.
2. Take 500 mL of 0.740 M sodium acetate (containing 0.370 mole), add to it 100 mL of 1.30 M HCl (containing 0.130 mole of HCl) and make up the mixture to a final volume of 1.00 L.

[Comment: The 0.130 mole of HCl converts 0.130 mole of the total 0.370 mole NaCH_3CO_2 into CH_3COOH , leaving $0.370 - 0.130 = 0.240$ mole of NaCH_3CO_2 behind. The 0.130 mole of CH_3COOH and the 0.240 mole NaCH_3CO_2 are in 1 L final volume.]

- (c) The following questions refer to the titration curve shown at the bottom of this page.
- (i) From the curve estimate the pH at the equivalence point.
estimated by inspection as pH = 8.8
 - (ii) Estimate the pK_a of the substance being titrated.

estimated as 5.0 by finding the pH at the half-equivalence point.

(iii) What class of titration does this curve typify? Circle your answer from the list below.

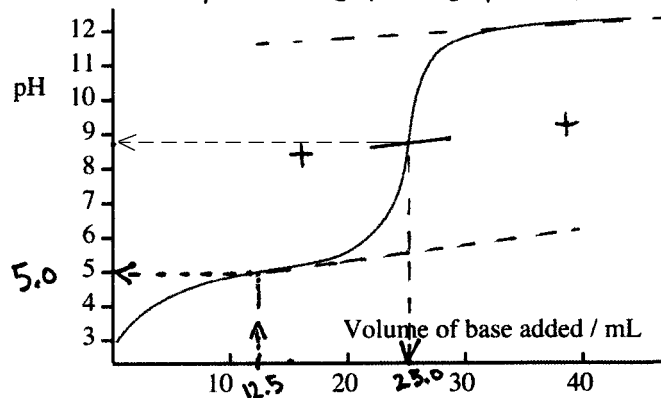
strong acid/strong base

strong acid/weak base

weak acid/strong base

(iv) Given the pK_a values for several indicators in the table on the right below, choose the BEST indicator for this titration. Justify your answer.

PHENOLPHTHALEIN, BECAUSE IT WILL CHANGE COLOUR CLOSEST TO THE EQUIVALENCE POINT (~pH 8.8).



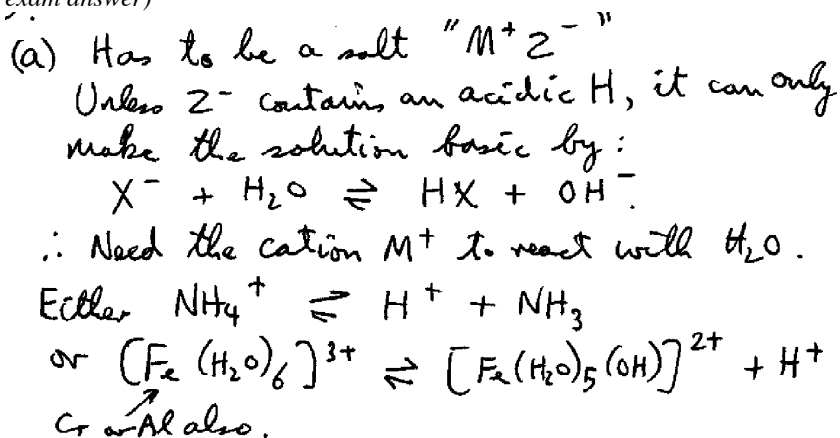
	pK_a at 25 °C
methyl orange	3.7
methyl red	5.1
bromothymol blue	7.0
phenolphthalein	9.6

QUESTION 12

- (a) Write the formula for a salt which gives an acidic solution in water, and write an ionic equation to show the acidic behaviour.

salt is NH_4Cl or FeCl_3 ... (could be others).....

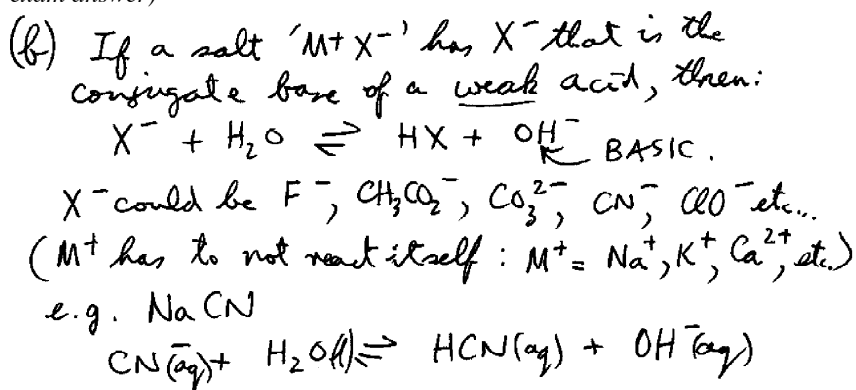
ionic equation : (below is the equation – and the detailed reasoning that is not required in the exam answer)



- (b) Write the formula for a salt which gives a basic solution in water, and write an ionic equation to show the basic behaviour.

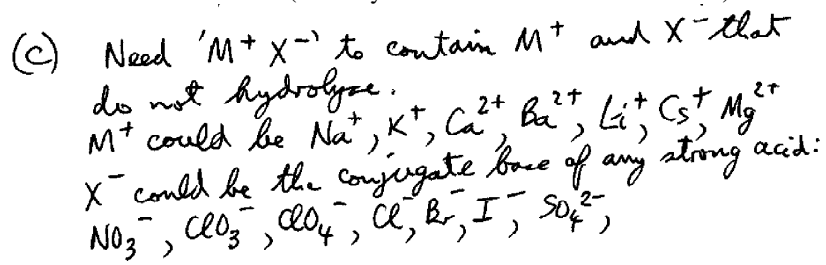
salt is NaCN (for example).....

ionic equation : (below is the equation – and the detailed reasoning that is not required in the exam answer)



- (c) Write the formula of a salt which gives a neutral solution in water.

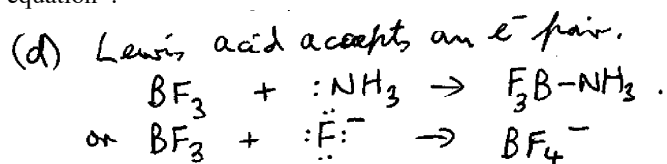
salt is NaCl (or many other combinations – see below).....



- (d) Write the formula for a Lewis acid, and write an equation to show the Lewis acidic behaviour.

Lewis acid BF_3

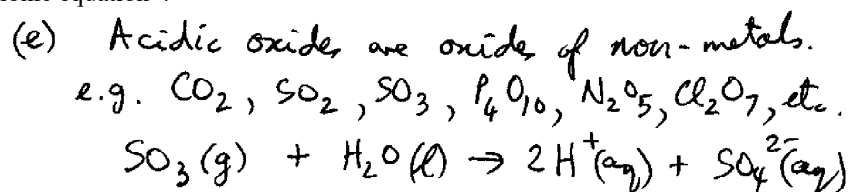
equation :



- (e) Write the formula for an oxide which gives an ionic acidic solution to show the acidic behaviour.

oxide is SO_3

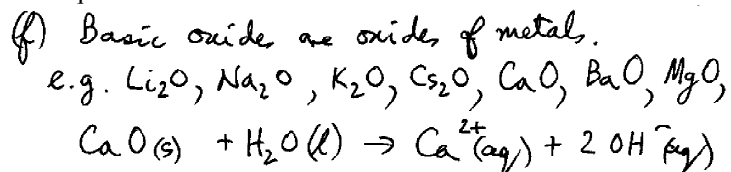
ionic equation :



- (f) Write the formula for an oxide which gives a basic solution in water, and write an ionic equation to show the basic behaviour.

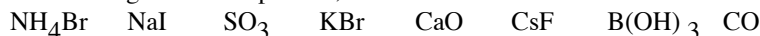
oxide is CaO

ionic equation :



QUESTION 13

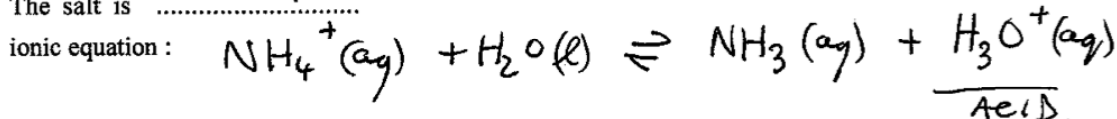
From the following list of compounds;



- (a) Choose the formula for a salt that gives an acidic solution in water, and write an ionic equation to show the

acidic behaviour.

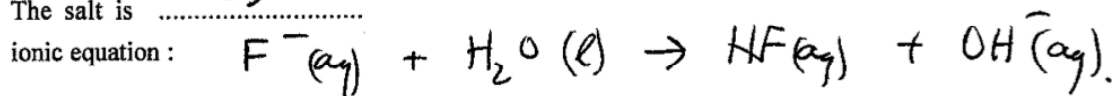
The salt is NH_4Br



- (b) Choose the formula for a salt that gives a basic solution in water, and write an ionic equation to show the basic

behaviour.

The salt is CsF

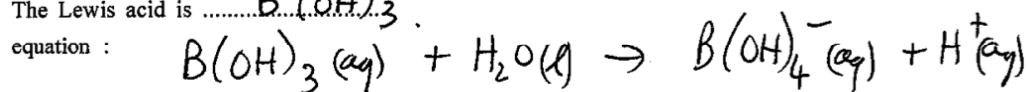


- (c) Choose the formula of a salt that gives a neutral solution in water.

The salt is KBr , NaI

- (d) Choose the formula for a Lewis acid, and write an equation to show the reaction of the Lewis acid with water.

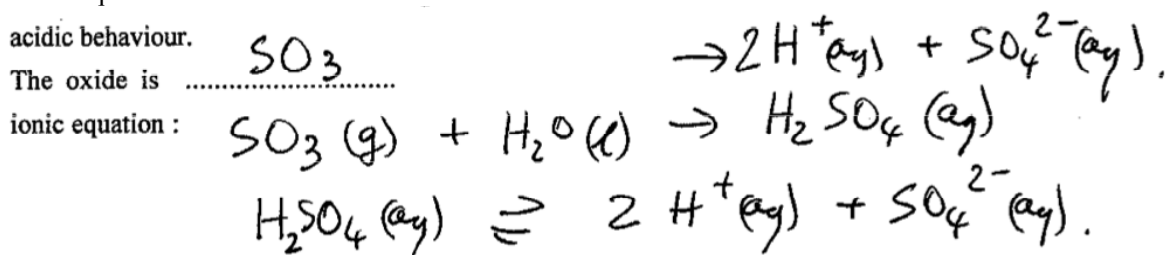
The Lewis acid is B(OH)_3



- (e) Choose the formula for an oxide that gives an acidic solution in water, and write an ionic equation to show

acidic behaviour.

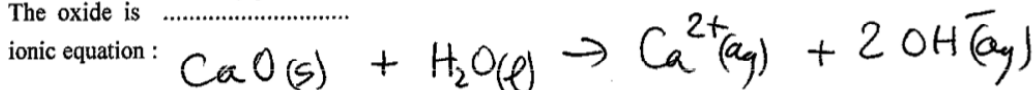
The oxide is SO_3



- (f) Write the formula for an oxide that gives a basic solution in water, and write an ionic equation to show the basic

behaviour.

The oxide is CaO



QUESTION 14

- (a) When 2.50 g of ammonium chloride is dissolved in 100.0 g of water the temperature falls by 1.70 K. Calculate the enthalpy of solution of ammonium chloride, assuming a specific heat capacity of $4.2 \text{ J K}^{-1} \text{ g}^{-1}$ and a density of 1.00 g mL^{-1} for all solutions.

$$\begin{aligned} \Delta H &= m C_p \Delta T \\ &= (102.5)(4.2)(1.70) \text{ J} \\ &= 731.85 \text{ J} \end{aligned}$$

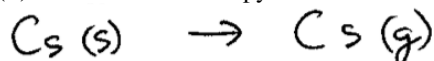
$1 \text{ mole } \text{NH}_4\text{Cl} = 53.492 \text{ g mol}^{-1}$
 $\therefore 2.50 \text{ g} = 2.50/53.492 \text{ mole}$
 $= 0.04674 \text{ mole}$
 $\therefore \text{Molar enthalpy} = \left(\frac{1}{0.04674}\right)(731.85 \text{ J})$
 $= 15.7 \text{ kJ mol}^{-1}$

- (b) Write suitable chemical equations including the states of matter for all reactants and products, to represent each of the following processes:

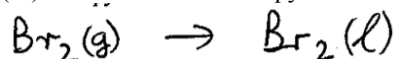
- (i) the standard enthalpy of formation of solid calcium fluoride.



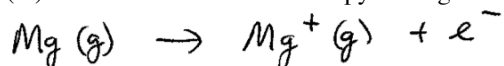
- (ii) the molar enthalpy of sublimation of caesium.



- (iii) the standard enthalpy of condensation of bromine.



- (iv) the first ionization enthalpy of magnesium.

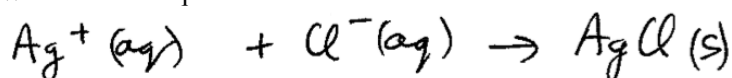


- (v) the first electron affinity of oxygen.



- (c) A calorimeter was used to measure the heat of reaction for mixing 100 mL of $0.100 \text{ M Ag}^+(\text{aq})$ with 100 mL of $0.100 \text{ M Cl}^-(\text{aq})$. The observed temperature change was $+0.770 \text{ K}$ for the 200 mL of solution (specific heat = $4.35 \text{ J K}^{-1} \text{ g}^{-1}$; density of solution = 1.00 g mL^{-1}) and calorimeter. The heat capacity of a calorimeter was 7.31 J K^{-1} .

- (i) Write the ionic equation for the reaction.



- (ii) How much heat was absorbed by the calorimeter?

$$\Delta H_{\text{cal}} = C_{\text{cal}} \Delta T = (7.31)(0.770) \text{ J} = 5.63 \text{ J}$$

- (iii) How much heat was absorbed by the solution?

$$\begin{aligned} \rho \text{ of } 200 \text{ mL solution} &= 1 \text{ g mL}^{-1} \\ \therefore \text{soln weighs } 200 \text{ g} \end{aligned}$$

$$\begin{aligned} \Delta H &= m C_p \Delta T \\ &= (200 \text{ g})(4.35)(0.770) \text{ J} = 670 \text{ J} \end{aligned}$$

QUESTION 15

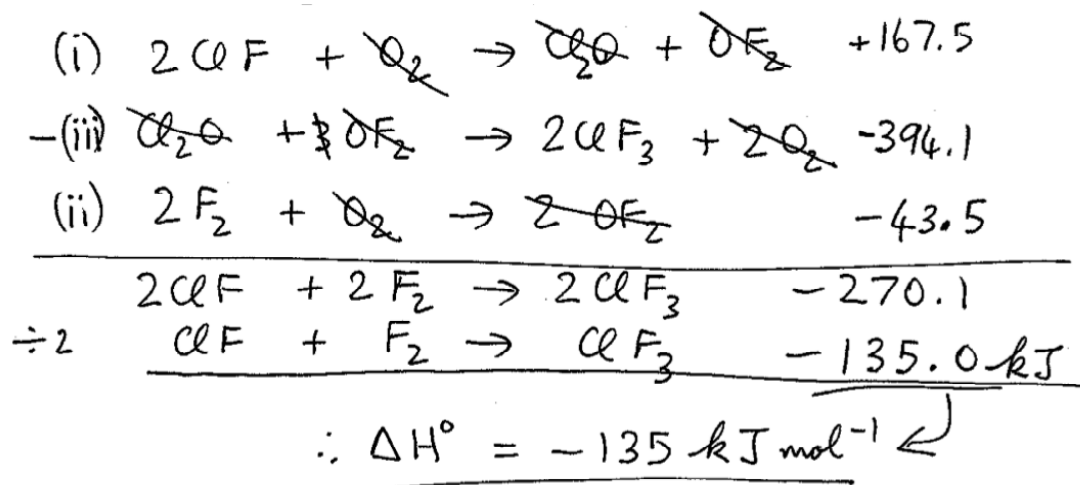
- (a) State Hess's Law.

or
(enthalpy)

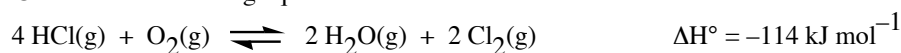
The heat change for a multi-step process equals the sum of the heat changes for the component steps of the process.

- (b) (i) $2 \text{ClF(g)} + \text{O}_2\text{(g)} \rightarrow \text{Cl}_2\text{O(g)} + \text{OF}_2\text{(g)}$ $\Delta H^\circ = 167.5 \text{ kJ}$
 (ii) $2 \text{F}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{OF}_2\text{(g)}$ $\Delta H^\circ = -43.5 \text{ kJ}$
 (iii) $2 \text{ClF}_3\text{(l)} + 2 \text{O}_2\text{(g)} \rightarrow \text{Cl}_2\text{O(g)} + 3 \text{OF}_2\text{(g)}$ $\Delta H^\circ = 394.1 \text{ kJ}$

Using appropriate data from the thermochemical equations above, calculate the standard enthalpy change for the following reaction:



- (c) Consider the following equilibrium :

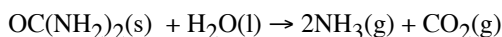


A mixture of HCl, O₂, H₂O, and Cl₂ is brought to equilibrium in a container at 400 °C. By placing a *tick* in the appropriate box indicate the effect on the equilibrium amount of Cl₂(g) in the container.

	Increase	Decrease	No Change
O ₂ is added at constant volume.	✓		
HCl is removed at constant volume.		✓	
Volume of the reaction container is doubled.		✓	
A catalyst is added.			✓
Temperature is increased to 500 °C.		✓	

QUESTION 16

- (a) Urea ($\text{OC}(\text{NH}_2)_2$) can react with water to produce carbon dioxide and ammonia:



- (i) Calculate ΔH° for the above reaction at 25 °C.

Thermodynamic data (at 25 °C, standard state = 1 bar)		
	$\Delta H_f^\circ / \text{kJ mol}^{-1}$	$S^\circ / \text{J K}^{-1} \text{mol}^{-1}$
$\text{NH}_3(\text{g})$	-46	193
$\text{CO}_2(\text{g})$	-394	214
$\text{OC}(\text{NH}_2)_2(\text{s})$	-334	105
$\text{H}_2\text{O}(\text{l})$	-286	70

$$\begin{aligned}\Delta H^\circ &= \sum n \Delta H_f^\circ \text{ PROD} - \sum n \Delta H_f^\circ \text{ REACT} \\ &= [2(-46) + (-394)] - [(-334) + (-286)] \\ &= +134 \text{ kJ mol}^{-1}\end{aligned}$$

- (ii) Calculate ΔS° for the above reaction at 25 °C.

$$\begin{aligned}\Delta S^\circ &= \sum n S_f^\circ \text{ PROD} - \sum n S_f^\circ \text{ REACT} \\ &= [2(193) + (214)] - [(105) + (70)] \\ &= +425 \text{ J K}^{-1} \text{mol}^{-1}\end{aligned}$$

- (iii) Calculate ΔG° for the above reaction at 25 °C.

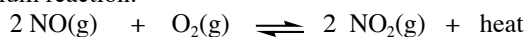
$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T \Delta S^\circ \\ &= (134 \times 10^3) - (298)(425) \text{ J mol}^{-1} \\ &= 7.35 \times 10^3 \text{ J mol}^{-1} \\ &= 7.35 \text{ kJ mol}^{-1}\end{aligned}$$

- (iv) Based on your answer to (iii), state whether the reaction above is spontaneous or not spontaneous under standard state conditions at 298 K. Justify your answer.

$\Delta G^\circ > 0$ \therefore The reaction is not spontaneous in the direction written.

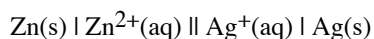
QUESTION 17

- (a) Complete the following table, by writing "increase", "decrease" or "no change" to indicate the effect each of the listed changes will have when applied to the mixture of gases in the following equilibrium reaction:

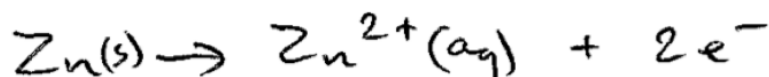


Change	Effect on amount of NO ₂ at equilibrium	Effect on K _p
Removing some NO from container	DECREASE	NO CHANGE
Adding a catalyst	NO CHANGE	NO CHANGE
Adding O ₂ to container	INCREASE	NO CHANGE
Increasing temperature	DECREASE	SMALLER
Compressing the container to a smaller volume (at constant temperature)	INCREASE	NO CHANGE

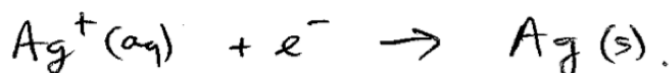
- (b) Consider the electrochemical cell represented by the following cell diagram:



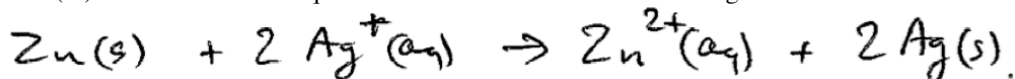
- (i) Write a balanced half-equation for the reaction occurring at the anode.



- (ii) Write a balanced half-equation for the reaction occurring at the cathode.



- (iii) Write a balanced equation for the overall reaction occurring in this cell.



- (iv) Using data from the adjacent table calculate the standard emf for this cell.

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{CATH.}} - E^\circ_{\text{ANODE}} \\ &= 0.80 - (-0.76) \\ &= 1.56 \text{ V} \end{aligned}$$

Standard reduction potentials, E° / V (at 25 °C, standard state = 1 mol L ⁻¹)	
Zn ²⁺ Zn	-0.76
Ag ⁺ Ag	0.80

- (v) Use the cell emf to calculate ΔG° for the cell reaction.

NOTE:

$$n = \text{THE NUMBER OF MOLES OF ELECTRONS EXCHANGED IN THE CELL EQUATION.}$$

$$\Delta G^\circ = -n F E^\circ = -(2)(96500)(1.56) = -301080 \text{ J mol}^{-1} = -301.1 \text{ kJ mol}^{-1}.$$

- (b) Calculate the current required (in amps) to deposit 0.500 g of chromium metal from a solution containing Cr³⁺(aq) in a period of 1.00 hour.

$$\begin{aligned} Q &= i \times t = n F \\ i (1)(60)(60) &= (3) \left(\frac{0.500}{52.0} \right) (96500) \\ i &= 0.773 \text{ amp.} \end{aligned}$$

CHEM1011**DATA SHEET**

$$0\text{ }^{\circ}\text{C} = 273\text{ K}$$

$$1\text{ atm} = 1.031\text{ bar} = 760\text{ mmHg} = 101.3\text{ kPa} = 760\text{ Torr}$$

$$\text{Gas Constant } R = 8.314\text{ J mol}^{-1}\text{ K}^{-1} = 0.08206\text{ L atm mol}^{-1}\text{ K}^{-1}$$

$$\text{Avogadro Number } N_{\text{A}} = 6.022 \times 10^{23}\text{ mol}^{-1}$$

$$\text{Speed of Light } c = 2.99 \times 10^8\text{ m s}^{-1}$$

$$\text{Planck Constant } h = 6.626 \times 10^{-34}\text{ J s}$$

$$\text{Planck equation : } E = h\nu = hc/\lambda$$

$$\text{Faraday Constant } F = 96,500\text{ C mol}^{-1}$$

$$\text{Nernst Equation } E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln Q$$

$$\text{Faraday Equation } Q = i \times t = \text{amount electrons (mole)} \times F$$

$$\text{Heat capacity of water, } C_{\text{p}} = 75\text{ J K}^{-1}\text{ mol}^{-1}$$

$$\text{Enthalpy of fusion of water, } \Delta H_{\text{m}} = 6\text{ kJ mol}^{-1}$$