

SEMESTER 1 EXAMINATION 2011

QUESTION/ANSWER BOOKLET

CHEMISTRY Stage 3

NAME: _____

CLASS: _____

TIME ALLOWED FOR THIS PAPER

Reading time before commencing work: Ten minutes

Working time for paper: Two and a half hours

MATERIAL REQUIRED/RECOMMENDED FOR THIS PAPER

TO BE PROVIDED BY THE SUPERVISOR

This Question/Answer Booklet

Separate Multiple Choice Answer Sheet

Chemistry Data Sheet

TO BE PROVIDED BY THE CANDIDATE

Standard Items: Pens, pencils, eraser, correction fluid, ruler, highlighters

Special Items: Non-programmable calculators satisfying the conditions set by the Curriculum Council for this course

No other items may be taken into the examination room.

It is your responsibility to ensure that you do not have any unauthorised notes or other items of a non-personal nature in the examination room. If you have any unauthorised material with you, hand it to the supervisor BEFORE reading any further.

STRUCTURE OF PAPER

Section	Number of questions available	No of Questions to be answered	Suggested working time (minutes)	Marks available	Percentage of exam
Part One: Multiple-choice	25	25	40	25	20
Part Two: Short answer	8	8	45	50	40
Part Three: Calculator	4	4	45	40	32
Part Four: Extended answer	2	1	20	10	8
					100

INSTRUCTIONS TO CANDIDATES

1. The rules for the conduct of Tuart College Examinations are detailed in the College Diary. Sitting this examination implies that you agree to and abide by these rules.
2. Answer the questions according to the following instructions:

Section One: Answer all questions on the separate Multiple-choice Answer Sheet provided. For each question shade the box to indicate your answer. Use only a blue or black pen to shade the boxes. If you make a mistake, place a cross through that square, do not erase or use correction fluid, and shade your new answer. Marks will not be deducted for incorrect answers. No marks will be given if more than one answer is completed for any question.

Sections Two and Three: Write answers in this Question/Answer booklet

3. When calculating numerical answers, show your working or reasoning clearly unless instructed otherwise. For some questions full marks will not be awarded for an answer without supporting working. The working must show concise logical steps.
4. Numerical answers are to be given to the appropriate number of significant figures.
5. For full marks, chemical equations should refer only to those species consumed in the reaction and new species produced. These species may be ions [for example $\text{Ag}^+(\text{aq})$], molecules [for example $\text{NH}_3(\text{g})$, $\text{NH}_3(\text{aq})$, $\text{CH}_3\text{COOH}(\text{l})$, $\text{CH}_3\text{COOH}(\text{aq})$] or solids [for example $\text{BaSO}_4(\text{s})$, $\text{Cu}(\text{s})$, $\text{Na}_2\text{CO}_3(\text{s})$].

Part One: Multiple Choice (25 marks)

This section has **25** questions. Answer **all** questions on the separate Multiple-choice Answer Sheet provided. For each question shade the box to indicate your answer. Use only a blue or black pen to shade the boxes. If you make a mistake, place a cross through that square, do not erase or use correction fluid, and shade your new answer. Marks will not be deducted for incorrect answers. No marks will be given if more than one answer is completed for any question.

Suggested working time for this section is 40 minutes

-
1. Which one of the following statements best describes the halogens?
 - (a) They are monatomic molecules that form X^{2-} ions
 - ☒ (b) They are diatomic molecules that form X^- ions
 - (c) They are monatomic molecules that form X^+ ions
 - (d) They are monatomic molecules that form X_2^+ ions

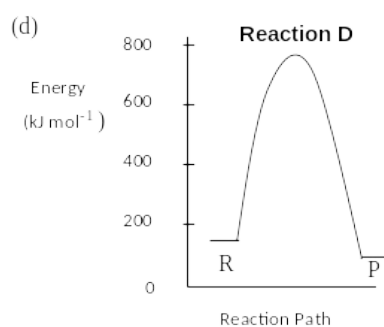
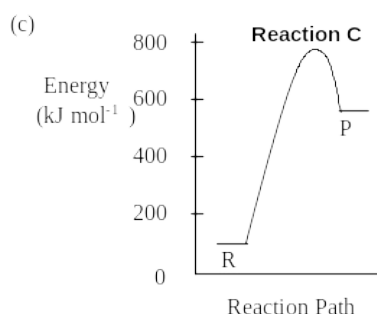
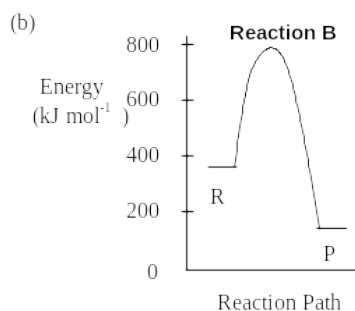
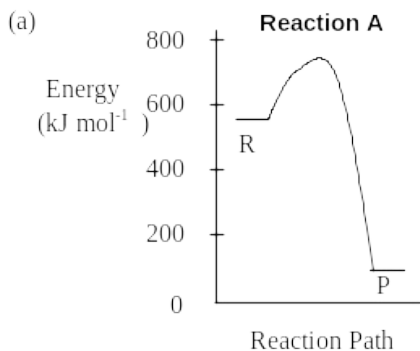
2. Which one of the following **best** explains why metals conduct electricity?
- (a) All metals are strongly electronegative
 - (b) In a metal crystal lattice, the metal atoms are not bonded strongly to each other
 - (c) Metal atoms have two electrons in their valence shells and these are easily removed
 - ☒ (d) The valence electrons of a metal do not belong to a particular atom but to the crystal or liquid as a whole.
3. Which one of the following best explains the polarity of carbon dioxide?
- ☒ (a) The CO_2 molecule is non-polar because, although the carbon to oxygen bond is polar, the molecule is linear.
 - (b) The CO_2 molecule is non-polar because O is more electronegative than C.
 - (c) The CO_2 molecule is non-polar because the molecule is bent or V-shaped.
 - (d) The CO_2 molecule is non-polar because the molecule is bent or V-shaped and O is more electronegative than C.
4. Which one of the following pairs of substances forms a buffer in aqueous solution?
- ☒ (a) $\text{H}_2\text{CO}_3(\text{aq})$ and $\text{HCO}_3^-(\text{aq})$
 - (b) $\text{H}_3\text{O}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$
 - (c) $\text{HCl}(\text{aq})$ and $\text{Cl}^-(\text{aq})$
 - (d) $\text{NaOH}(\text{aq})$ and $\text{HNO}_3(\text{aq})$
5. The inert gas element with the lowest boiling point is:
- ☒ (a) Helium
 - (b) Neon
 - (c) Argon
 - (d) Chlorine
6. The set that contains a molecule, an ionic solid and a network solid (not necessarily in that order) is:
- (a) NaCl , MgO and Al_2O_3
 - ☒ (b) NaCl , SiO_2 and CO_2
 - (c) HCl , H_2O and SiO_2
 - (d) NaCl , MgO and SiC
7. Which one of the following pairs of reagents would give **no** observable evidence of a chemical reaction when mixed?
- (a) Lead (II) nitrate solution and potassium iodide solution
 - (b) Potassium and water
 - (c) Sodium hydroxide and warm ammonium chloride solution
 - ☒ (d) Barium hydroxide solution and potassium chloride solution
8. Identify the strongest base in the following equilibrium. The concentration of products is much higher than the concentration of reactants.
- $$\text{HTe}^- + \text{NH}_3 \rightleftharpoons \text{H}_2\text{Te} + \text{NH}_2^-$$
- ☒ (a) HTe^-
 - (b) NH_3
 - (c) H_2Te
 - (d) NH_2^-
9. Which one of the following statements is true concerning the titration of a 0.100 mol L^{-1} hydrochloric acid solution with 20.00 mL of an approximately 0.1 mol L^{-1} sodium carbonate solution?
- ☒ (a) The equivalence point would be somewhat acidic
 - (b) You would expect to use exactly 20.00 mL of acid
 - (c) You would expect to use approximately 20 mL of acid
 - (d) A suitable indicator would be phenolphthalein
10. A 1.00 L solution contains 1.5 mol of $\text{Ca}(\text{NO}_3)_2$ and 2.0 mol of NaNO_3 . The concentration of nitrate ions in this solution is:
- ☒ (a) 5.0 mol L^{-1}
 - (b) 3.5 mol L^{-1}
 - (c) less than 3.5 mol L^{-1}
 - (d) between 3.5 mol L^{-1} and 5.0 mol L^{-1}
11. Which one of the following statements does **not** apply to 1.00 mol of hydrogen gas molecules?

1.00 mol of hydrogen molecules,

- a) contains 6.022×10^{23} molecules
- b) occupies 22.41 L at S.T.P.
- c) contains 1.2044×10^{24} atoms
- ☒ has a mass of 1.008 g

The following information applies to questions 12, 13, and 14:

Each diagram below shows the energy change for four different reactions:



12. An endothermic reaction is:

- (a) A
- (b) B
- ☒ C
- (d) D

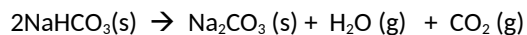
13. The reaction with enthalpy change of about -50 kJ mol^{-1} is:

- (a) A
- (b) B
- (c) C
- ☒ D

14. The reaction with the highest activation energy is:

- (a) A
- (b) B
- ☒ C
- (d) D

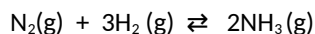
15. Sodium hydrogencarbonate decomposes on heating as shown below:



If 0.40 mole of sodium hydrogencarbonate is heated, the total number of moles of gaseous products will be

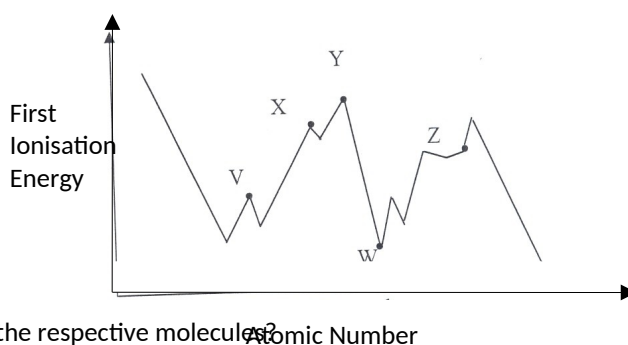
- (a) 0.20 mol
- ☒ 0.40 mol
- (c) 0.60 mol
- (d) 0.80 mol

16. Which one of the following statements about the following reversible reaction is true?



- a) ☒ K = $\frac{1}{27}$
- (b) ☐ K is constant under all conditions
- (c) ☐ A catalyst increases the yield of ammonia by increasing ΔH
- (d) ☐ Increasing the pressure increases K
17. Barium hydroxide dissociates completely in water. The pH of a $0.0005 \text{ mol L}^{-1}$ solution of barium hydroxide in water is:
- (a) ☐ 3.0
- (b) ☐ 3.3
- (c) ☐ 10.7
- ☒ 11.0
18. Which one of the following aqueous solutions would have a pH greater than 7.00 at 25°C ?
- ☒ 2.00 mol L^{-1} sodium ethanoate solution
- (b) ☐ 2.00 mol L^{-1} sodium chloride solution
- (c) ☐ 2.00 mol L^{-1} ammonium chloride solution
- (d) ☐ 2.00 mol L^{-1} hydrogen chloride solution
19. In which one of the following reactions is the **boldly highlighted** substance acting as an acid?
- (a) ☐ $\text{C}_6\text{H}_5\text{COO}^- (\text{aq}) + \text{H}_2\text{SO}_3 (\text{aq}) \rightarrow \text{HSO}_3^- (\text{aq}) + \text{C}_6\text{H}_5\text{COOH} (\text{s})$
- (b) ☐ $2\text{Na} (\text{s}) + 2\text{H}_2\text{O} (\ell) \rightarrow 2\text{Na}^+ (\text{aq}) + 2\text{OH}^- (\text{aq}) + \text{H}_2 (\text{g})$
- (c) ☐ $\text{HCO}_3^- (\text{aq}) + \text{H}_2\text{O} (\ell) \rightarrow \text{H}_2\text{CO}_3 (\text{aq}) + \text{OH}^- (\text{aq})$
- ☒ $\text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{NH}_3 (\text{g}) + \text{H}_2\text{O} (\ell)$
20. Which one of the following elements has an oxide which reacts with water to form an acidic solution?
- (a) ☐ Iron
- ☒ Phosphorus
- (c) ☐ Magnesium
- (d) ☐ Aluminium

21. Consider the elements labelled V to Z on the diagram
- Which pair of these is most likely to form a covalent compound?



- (a) ☐ X and Y
- (b) ☐ X and V
- ☒ X and Z
- (d) ☐ V and Z

... shapes of the respective molecules

P = pyramidal

L = linear

B = bent or V-shaped

- | | H_2CO | NH_3 | BF_3 | H_2S |
|-------------------------------------|-----------------------|---------------|---------------|----------------------|
| (a) <input type="checkbox"/> | T | T | P | B |
| <input checked="" type="checkbox"/> | T | P | T | B |
| (c) <input type="checkbox"/> | P | P | T | L |
| (d) <input type="checkbox"/> | P | P | P | L |

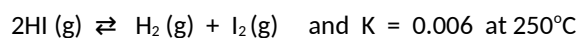
* does not obey the octet rule

23. "Saline solution", which is used in hospitals, has a sodium chloride concentration of 0.154 mol L^{-1}
- The mass of sodium chloride required to prepare 0.500 L of this solution is:
- (a) ☐ 14.9 g
- ☒ 4.50 g
- (c) ☐ 4.50 mg
- (d) ☐ 1.49 mg

24. Which one of the following formulae for the fluorides of the elements in the second row of the periodic table is **incorrect**?

- (a) LiF
- (b) CF₄
- ☒ (c) NF₅
- (d) OF₂

25. Hydrogen iodide decomposes according to the reaction:



If equimolar amounts of H₂, I₂ and HI were mixed at this temperature,

- ☒ (a) the amount of iodine would be reduced
- (b) the concentration of HI in the mixture would decrease
- (c) the value of K would increase to 1.0
- (d) the number of molecules of gas would increase

END OF MULTICHOICE SECTION

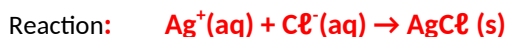
Part Two: Short Answer (50 marks)

This section has 8 questions. Answer all questions. Write your answers in the spaces provided

1. Write equations for any reactions that occur in the following procedures. If no reaction occurs, write "no reaction". Reactions must relate to the actual species reacting. In each case describe in full what you would observe, including any
- colours
 - odours
 - precipitates (give the colour)
 - gases evolved (give the colour or describe as colourless)

If a reaction occurs but is not visible, you should state this.

- (a) Dilute hydrochloric acid is added to silver nitrate solution



Observation: *white precipitate is formed from the mixing of two colourless solutions*

(4 marks)

- (b) A dilute solution of barium chloride is added to a dilute solution of potassium hydroxide

Reaction: *No Reaction*

Observation: *No visible observation*

(4 marks)

2. Use the Bronsted-Lowry Theory to write equations to illustrate the following reactions in water:

- (a) The hydrogencarbonate ion acting as a base:



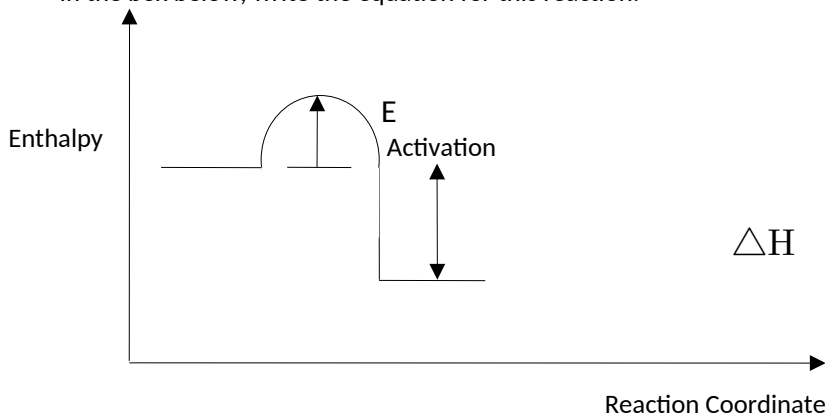
- (b) The hydrogensulfate ion acting as an acid:



(4 marks)

3. In the Thermite process, a finely divided mixture of aluminium powder and iron(III) oxide can react to produce white-hot molten iron. The mixture initially needs to be ignited with burning magnesium.

- (a) On the axes below, draw a potential energy diagram for the reaction between aluminium and iron(III) oxide
- (b) Label the vertical axis as well as the activation energy and ΔH term
- (c) In the box below, write the equation for this reaction:



4. Explain the following data:

(5 marks)

Boiling Points (°C)

Group 14 hydrides

CH ₄	-162
SiH ₄	-111
GeH ₄	-88
SnH ₄	-52
PbH ₄	-13

Group 17 hydrides

HF	20
HCl	-85
HBr	-67
HI	-35

For **group 14 hydrides**, boiling points increase as you go down the group because molecules get larger and therefore dispersion forces increase.

For **group 17 hydrides**, the boiling point of HF is unusually high due to the presence of intermolecular hydrogen bonding. As we go from HCl to HI, the dispersion forces increase (larger molecules) and therefore the boiling points increase.

(4 marks)

5. Complete the following table:

Species	Electron Dot diagram	Shape (Drawing)	Name of shape
SO ₂	<pre> : O : S :: O </pre>		Bent or V-shaped
H ₂ S	<pre> .. H : S : H .. </pre>		Bent or V-shaped
CH ₂ Cl ₂	<pre> H H : C : Cl : : Cl : .. </pre>		Tetrahedral

(9 marks)

6. **Name** a substance whose properties match those described below:
- (a) A highly polar molecular substance that is very soluble in water. Individual molecules are described as having a pyramidal shape. **ammonia**
 - (b) Combines with another element by sharing one pair of valence electrons. When combined with hydrogen it forms molecules that exhibit hydrogen bonding and are acidic. **fluorine**
 - (c) Has a very first high ionisation energy, a very low boiling point and does not form bonds with the two elements immediately before or after it in the periodic table. **helium, neon, argon**
 - (d) In its pure form it is very malleable and ductile and reacts vigorously with water. Salts containing this ion are always soluble. **potassium, sodium**

(4 marks)

7. (a) Some compounds behave as buffers, that is, they have a buffering capacity.

(i) Explain qualitatively the concept of buffering capacity, and state one factor on which buffering capacity depends.

(2 marks)

Buffering capacity is the ability of a buffer to resist changes in pH.

Buffering capacity depends on the relative amounts of a weak acid (and its conjugate base) or the relative amounts of a weak base (and its conjugate acid)

(ii) Explain how buffers respond to the addition of H^+ and OH^- ions.

Using the example of CH_3COOH/CH_3COO^- ,

(2 marks)



Adding H^+ ions: Since there is a large amount of CH_3COO^- (mainly from the added salt), the added H^+ ions are consumed by the significant amount of CH_3COO^- ions, until a new equilibrium is established, producing virtually the same pH as before.

Adding OH^- ions: The added hydroxide ions react with the H^+ ions, and the equilibrium position shifts to the right (to replace the H^+ ions), producing virtually the same pH as before.

(iii) How would the buffering capacity of a $0.01 \text{ mol L}^{-1} \text{ NH}_3 / 0.01 \text{ mol L}^{-1} \text{ NH}_4\text{NO}_3$ solution differ, if at all, from a $0.01 \text{ mol L}^{-1} \text{ NH}_3 / 0.01 \text{ mol L}^{-1} \text{ NH}_4\text{Cl}$ solution?

In other words, would the buffering capacity of the $0.01 \text{ mol L}^{-1} \text{ NH}_3 / 0.01 \text{ mol L}^{-1} \text{ NH}_4\text{NO}_3$ solution be greater than, less than or the same as the $0.01 \text{ mol L}^{-1} \text{ NH}_3 / 0.01 \text{ mol L}^{-1} \text{ NH}_4\text{Cl}$ solution? Explain your answer.

Circle the correct answer

Greater than

Less than

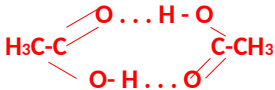
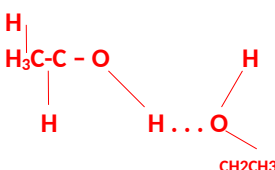
Same as

Explanation: **The important factor is the ratio of $[NH_4^+]$ from the salt to the $[NH_3]$, which is the same in both cases**

(3 marks)

8. Complete the following table:

(9 marks)

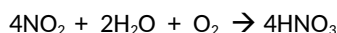
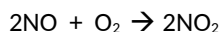
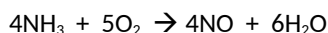
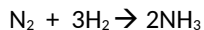
Pairs of substances	Substance with the higher boiling point	most significant intermolecular force	Reason why this force is greater for the substance with the higher boiling point
CH_3OH and CH_3NH_2	CH_3OH	Hydrogen bonding	Oxygen is more electronegative than Nitrogen (H bonding for CH_3OH is therefore stronger)
$\text{C}_{15}\text{H}_{32}$ and C_5H_{12}	$\text{C}_{15}\text{H}_{32}$	Dispersion forces	Dispersion forces increase as molecular size increases
CH_3COOH and $\text{CH}_3\text{CH}_2\text{OH}$	CH_3COOH	Hydrogen bonding	<p>The C of the CH_3COOH has two opportunities to link via hydrogen bonding to another - COOH unit whereas the CH_2OH can only link once via H bonding.</p>  

END OF PART TWO

Part 3 (40 marks)

Answer all questions in part 3. The calculations are to be set out in detail in this Question/Answer Booklet. Marks will be allocated for correct equations and clear setting out, even if you cannot complete the problem. When questions are divided into sections, clearly distinguish each section using (a), (b) and so on. Express your final numerical answers to three (3) significant figures where appropriate, and provide units where applicable. Information which may be necessary for solving the problems is located on the separate Chemistry Data Sheet. Show clear reasoning: if you don't, you will lose marks.

1. The reactions involved in the manufacture of nitric acid can be represented as follows:



- (a) Calculate the mass of nitrogen required to manufacture 1.00 tonne (1.00×10^3 kg) of nitric acid (4 marks)

$$\begin{aligned} n(\text{N}_2) &= \frac{1}{2} \times n(\text{HNO}_3) \\ &= \frac{1}{2} \times \frac{(1000 \times 1000)}{[1.008 + 14.01 + 3 \times 16]} \end{aligned}$$

$$n(\text{N}_2) = 7.93... \times 10^3$$

Hence

$$\begin{aligned} m(\text{N}_2) &= (7.93... \times 10^3) [2(14.01)] \\ &= 2.22 \times 10^5 \text{ g} \end{aligned}$$

- b) Calculate the volume of nitrogen gas at 130 kPa and 25.0 °C used to manufacture 1.00 tonne of nitric acid. (3 marks)

$$\begin{aligned} V &= \frac{nRT}{P} \\ &= \frac{(7.93... \times 10^3)(8.315)(25.0 + 273.1)}{130} \\ &= 1.51 \times 10^5 \text{ L} \end{aligned}$$

2. 11.2 g of anhydrous sodium hydroxide is dissolved in 5.80×10^2 mL of a 0.560 mol L^{-1} hydrochloric acid solution.

- (a) Calculate the pH of the final solution, assuming the volume does not change on adding the solid to the acid.

$$n(\text{OH}^-) = \frac{11.2}{[22.99 + 16.0 + 1.008]} = 2.80... \times 10^{-1}$$

$$\text{Also, } n(\text{H}^+) = cV = 0.560 \times 0.580 = 3.24... \times 10^{-1}$$

REACTION is $\text{OH}^- + \text{H}^+ \rightarrow \text{H}_2\text{O}$

$$\text{Therefore moles of excess } \text{H}^+ = 3.24... \times 10^{-1} - 2.80... \times 10^{-1} = 4.47... \times 10^{-2}$$

$$\text{Hence } c(\text{H}^+) = \frac{n(\text{H}^+)_{\text{excess}}}{V_{\text{total}}} = \frac{4.47... \times 10^{-2}}{0.580} = 7.72... \times 10^{-2} \text{ mol L}^{-1}$$

$$\text{Therefore pH} = -\log(7.72... \times 10^{-2}) = 1.11$$

(5 marks)

- (b) A student was then given a pH meter to measure the pH of the final solution above and discovered that the solution was not neutral. To make the final solution neutral the student needed to use either more solid KOH or some 0.102 mol L⁻¹ nitric acid. Determine which of the reagents (solid KOH or 0.102 mol L⁻¹ nitric acid) is required to neutralise the solution. If it is the solid, calculate the mass required (in g). If it is the nitric acid, calculate the volume required (in mL).

(6 marks)

$$n(\text{H}^+) \text{ excess} = 4.47... \times 10^{-2}$$

To neutralise this solution you need 4.47... $\times 10^{-2}$ moles of OH⁻ (from KOH)

$$\text{Therefore } n(\text{KOH}) = 4.47... \times 10^{-2}$$

$$\begin{aligned} m(\text{KOH}) &= 4.47... \times 10^{-2} [39.10 + 16.0 + 1.008] \\ &= 2.51 \end{aligned}$$

3. Salinity in soil is a serious problem in many areas of Western Australia. In an effort to monitor changes in the sodium chloride content of soil a technician from Agriculture WA collected a soil sample for analysis. The technician carried out the following analysis.

A 500.0 g sample of soil was mixed with enough water to dissolve all the sodium chloride in the sample. The mixture was filtered, then made up to 1.00 L. 100.0 mL samples of this solution were then titrated with a standardised 4.998×10^{-3} mol L⁻¹ silver nitrate solution using potassium chromate as an indicator of when all the chloride ions had been precipitated.

After several trials the average volume of silver nitrate solution used was found to be 23.65 mL. Assume that the only source of chloride ions was sodium chloride.

- (a) Calculate the concentration of sodium chloride in the soil measured in parts per million (ppm).

(7 marks)



$$n(\text{Cl}^-) = n(\text{Ag}^+) = (4.998 \times 10^{-3})(0.02365) = 1.18... \times 10^{-4}$$

$$\text{therefore } n(\text{NaCl}) = 1.18... \times 10^{-4}$$

$$m(\text{NaCl}) = (1.18... \times 10^{-4})[22.99 + 35.35] = 6.90... \times 10^{-3} \text{ g}$$

For a 100 mL sample

$$\begin{aligned} \text{For a 1000 mL sample, } m(\text{NaCl}) &= 6.90... \times 10^{-3} \times 10 = 6.90... \times 10^{-2} \text{ g} \\ &= 6.90... \times 10^{-2} (1000) \text{ mg} \\ &= 6.90... \times 10^1 \text{ mg} \end{aligned}$$

$$\begin{aligned} \text{Concentration (in ppm)} &= \frac{6.90... \times 10^1 \text{ mg}}{0.500 \text{ kg}} \\ &= 138 \text{ ppm} \end{aligned}$$

- (b) What is the concentration (in mol L⁻¹), of sodium ions in the 1.00 L of sodium chloride solution?

(4 marks)

$$n(\text{NaCl}) = n(\text{Cl}^-) = 1.18... \times 10^{-4} \text{ for a 100 mL sample}$$

$$\text{For a 1000 mL sample, } n(\text{NaCl}) = n(\text{Cl}^-) = 1.18... \times 10^{-4} \times (10) = 1.18... \times 10^{-3}$$

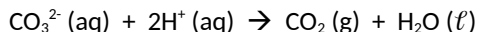
$$c(\text{Na}^+) = \frac{1.18... \times 10^{-3}}{1.00} = 1.18 \times 10^{-3} \text{ mol L}^{-1}$$

OR the 100 mL sample has the same concentration as the 1.00 L sample, and

$$c(\text{Na}^+) = 1.18... \times 10^{-4} / (0.100) = 1.18 \times 10^{-3}$$

4. One brand of concrete-cleaning liquid consists essentially of a mixture of water, hydrochloric acid and a small amount of detergent which has no acid-base properties. The concentration of hydrochloric acid in the liquid may be determined by titrating it with a base. A chemist analysed the liquid using anhydrous sodium carbonate as a base, according to the following method:

10.6 g of anhydrous sodium carbonate was made up to 250.0 mL in a volumetric flask. After thorough mixing, 25.0 mL of this solution was pipetted into a dry conical flask. The concrete-cleaning liquid was added slowly from a burette until the end point of the reaction was reached. Methyl orange indicator was used to indicate the completion of the reaction,



The volumes of cleaning liquid titrated from the burette were:

21.89 mL (rough titration) , 22.03 mL, 25.09 mL, 21.99 mL, 22.00 mL

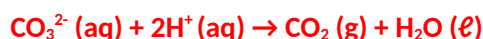
- (a) Use the data provided to calculate the concentration of hydrochloric acid in the cleaning liquid. Express your answer in mol L^{-1} . (7 marks)

$$\begin{aligned} \text{Mean volume used in titration} &= \frac{(22.00 + 21.99)}{2} \text{ or } \frac{(22.00 + 21.99 + 22.03)}{3} \\ &= 22.00 \text{ mL} \end{aligned}$$

To calculate the concentration of Na_2CO_3 :

$$c = \frac{n}{v} = \frac{\frac{10}{250} [2(22.99) + 12.01 + 3(16)]}{0.250} = 0.400 \text{ mol L}^{-1}$$

From the equation,



$$n(\text{H}^+) = 2 \times n(\text{CO}_3^{2-})$$

$$c(\text{H}^+)(0.02200) = 2 \times (0.400)(0.02500)$$

$$c(\text{H}^+) = 9.09 \times 10^{-1} \text{ mol L}^{-1}$$

- (b) Would the value of the concentration determined in (a) above be increased, decreased or unchanged if she had

- (i) Washed the conical flask with water, but not dried it prior to pipetting in the sodium carbonate solution?

(2 marks)

ANSWER: (Write " increased" , "decreased" or "unchanged")

UNCHANGED

REASON FOR ANSWER:

The critical factor is the number of moles of sodium carbonate in the conical flask. This is the same even if there is some residual water in the conical flask.

- (ii) Unknowingly used hydrated sodium carbonate ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) instead of the anhydrous form of sodium carbonate?

(2 marks)

ANSWER: (Write " increased" , "decreased" or "unchanged")

DECREASED

REASON FOR ANSWER:

Concentration of Na_2CO_3 would be less; therefore H^+ concentration would be less (refer back to calculation in part (a))

END OF PART 3

Part 4 (10 marks)

Answer ONE of the following extended answer questions. Where applicable use equations, diagrams and illustrative examples of the chemistry you are describing. Marks are awarded for the relevant chemical content of your answer, but you will lose marks if what you write is presented in point form. As a guide, a minimum of two pages is required.

1. "An understanding of the three dimensional structure of a covalent molecule enables its polarity and intermolecular forces to be predicted".

Expand on this statement by discussing the following topics.

- (i) electron pair repulsion theory
- (ii) shapes of molecules
- (iii) molecular polarity
- (iv) intermolecular forces.

You may use water, methane, carbon dioxide, ammonia and any other appropriate molecules as examples.

OR

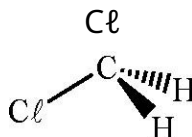
Write on BOTH the following fundamental aspects of chemistry:

- (a) Consider the elements of the Periodic Table in the row from sodium to argon. For each of these elements discuss the relationship that exists between the electronic configuration, valency and the type of bonding exhibited by the element.
- (b) Describe the nature of the bonding present in an ionic solid, in a metal and in a network solid by reference to one specific example of each type. Relate this to the physical properties of melting point and electrical conductivity exhibited by each of the solids.

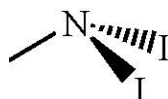
ESSAY 1

Electron Pair repulsion Theory: Basic ideas are that (a) electron pairs arrange themselves as far apart as possible in order to minimise repulsion and (b) lone pairs repel bonding pairs more than bonding pairs repel other bonding pairs because of the fact that lone pairs are always very close to atoms.

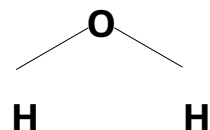
For example methane (CCl_2H_2) has 4 bonding pairs which arrange themselves so as to produce a tetrahedral shape



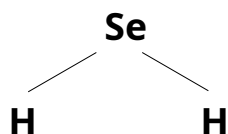
Similarly, NI_3 molecules arrange themselves to form a pyramidal shape




Also water molecules arrange themselves to form a bent shape:

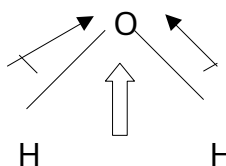


Finally a linear shape is possible with molecules like H_2Se

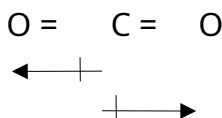


The polarity of a molecule is the nett sum of the individual bond dipoles (dipoles arise from two atoms having different electronegativities or different electron attracting powers)

Hence in the example of water above, there is a nett dipole () as follows:



Contrast this to the situation in CO_2 where there are individual dipoles which cancel out to give zero nett dipole, that is the molecule is non-polar

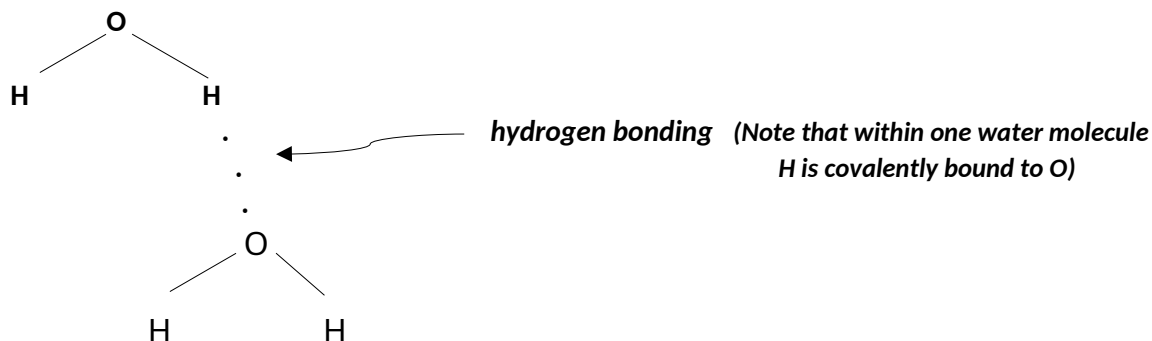


Intermolecular forces can be either hydrogen bonds, dipole- dipole forces or dispersion forces.

Dispersion forces exist between all molecules.

Hydrogen bonds are the interactions between F, O or N (on one molecule) to H (on another molecule). Note that the F, O or N must be covalently bonded to H in the single molecule and also lone pair(s) must be present on either F, O or N.

Example of Intermolecular Hydrogen bonding



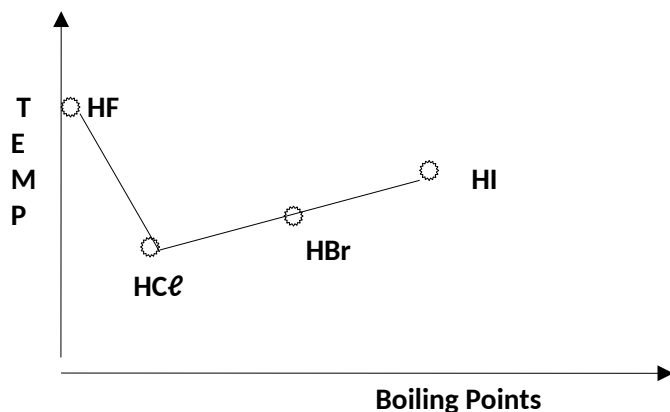
Dipole- dipole interactions are of similar type but not involving F, O , N, For example:



dipole- dipole interaction (not as strong as hydrogen bonding on account of the electronegativities of elements other than F, O or N are less than the values for F, O or N).

Dispersion forces arise from temporary induced dipoles and are significant for larger molecules (leading to higher melting and boiling points) and also are the only intermolecular forces between totally symmetric or non polar molecules)

If we consider the trends for the boiling points of the hydrogen halides (HF, HCl, HBr, HI), we obtain a qualitative graph like:



The abnormally high boiling point of HF is attributed to hydrogen bonding (dispersion forces while present are not significant). The dipole -dipole forces for HCl, HBr and HI decrease since electronegativities decrease as you go down group 17.

However the boiling points increase due to the fact that the sizes of the molecules increase and therefore dispersion forces increase markedly. This latter trend in increasing dispersion forces outweighs the trend for a decrease in dipole -dipole forces.

ESSAY NUMBER 2

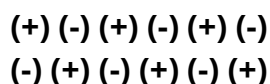
Na is in group 1 and has 1 outer electron. Mg is in group 2 and has 2 outer electrons. The trend continues right across the period through to Ar which has 8 outer electrons in its outside shell.

Since elements achieve stability by attempting to have a full outer shell, the metallic elements Na, Mg, Al will lose one, two and three electrons respectively therefore resulting in a full outer shell.

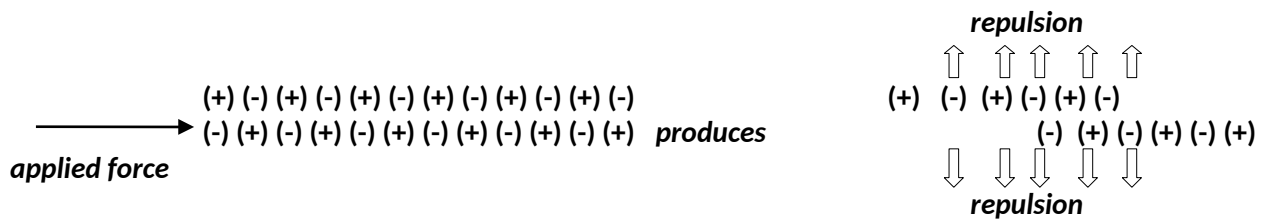
Hence Na has a valency of (+1), Mg has a valency of (+2) and Al has a valency of (+3). In a similar manner, the elements P, S and Cl will gain 3, 2 and 1 electron respectively to attain a full outer shell configuration. Hence these elements can have valencies of (-3), (-2) and (-1) respectively.

On the other hand silicon can achieve stability by sharing its outer 4 electrons – hence it does not have a valency.

In ionic solids such as NaCl, the sodium atom originally donates its single outer electron to the chlorine atom which needs one electron to attain a full outer shell configuration. The result is that in sodium chloride, sodium ions (+) and chloride ions (-) are formed. These ions arrange themselves in a alternating (+) (-) array that is held together by strong electrostatic attraction. The array can be represented two dimensionally as follows:

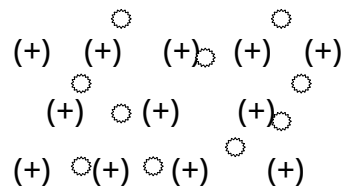


In three dimensions, the alternating (+) (-) ions create a substance which has a high melting point (on account of the strong electrostatic forces of attraction) and one which is hard and brittle. Attempts to deform the three dimensional structure with a applied force result in adjacent ions repelling each other as shown below:

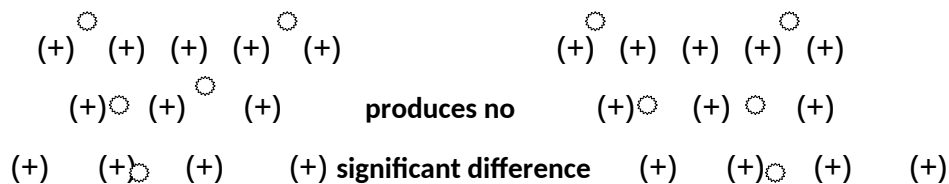


In the solid state an ionic solid consists of alternately charged ions held in fixed positions. Therefore electrical conductivity is not possible. However if the solid is dissolved or is heated up to melt it, the ions can now move relatively freely and therefore ionic solids conduct electricity in the molten and aqueous state.

For metallic substances, a typical structure consists of positively charged nuclei with electrons moving around the intervening spaces as follows:



The electrons are relatively delocalised since they are outer shell valence electrons that are not very strongly held to any particular nucleus. The melting and boiling points of most metals however are still reasonably high on account of the electrostatic attraction of all electrons to all positive nuclei. The fact that the electrons are basically delocalised accounts for the electrical conductivity of metals. Metals are also malleable/ductile because of the fact that the layers of positive ions that are surrounded by a "sea" of electrons can be easily moved, i.e.,



On the other hand silicon shares its electrons with (say oxygen) to form a tetrahedral structure (SiO_2) which resembles that of diamond.

The structure consists of one giant lattice of atoms of Si and O which are covalently bound to each other in three dimensions. This 3-DIM set up creates a structure that is incredibly hard to break up and therefore the melting point is extremely high.

In addition, SiO_2 has not got any free electrons (like metals do) so it is therefore a non-conductor.