

**Western Australian Certificate of Education**

**Semester 1 Examination, 2017**

**Question/Answer Booklet**

**NAME**

**ATAR CHEMISTRY**

**Unit 3**

**Time allowed for this paper**

Reading time before commencing work: **10 minutes**

Working time for paper: **3 hours**

**Materials required/recommended for this paper**

***To be provided by the supervisor***

This Question/Answer Booklet

Multiple-choice Answer Sheet

Chemistry Data Sheet

***To be provided by the candidate***

Standard items: pens, pencils, eraser, correction fluid/tape, ruler, highlighters

Special items: non-programmable calculators satisfying the conditions set out by the

Curriculum Council for this course

**OFFICE USE ONLY**

|  |  |  |  |
| --- | --- | --- | --- |
|  | **MARK** | **TOTAL** | **%** |
| **Section 1** |  | **50** |  |
| **Section 2** |  | **70** |  |
| **Section 3** |  | **80** |  |
| **TOTAL** |  | **200** |  |

**Important note to candidates**

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 THE PAPER**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Section** | **No. of Questions Set** | **No. of Questions to be Attempted** | **Marks Allocated** | **Recommended Time (approx) in Minutes** |
| Section1:  Multiple-choice | 25 | 25 | 50 (25%) | 50 |
| Section 2:  Short answers | 10 | 10 | 70 (35%) | 60 |
| Section 3:  Extended answer | 6 | 6 | 80 (40%) | 70 |

**Total Marks for Paper = 200 (100%)**

**INSTRUCTIONS TO CANDIDATES**

**Reading Time**: The examiners recommend that candidates spend the reading time mainly reading the Instructions to Candidates and Sections One, Two and Three.

***Section One - Multiple Choice***

Answer on the separate Multiple Choice Answer Sheet.

If you consider that two or more of the alternative responses are correct, choose the one that you think is best. If you think you know an answer, mark it even if you are not certain you are correct. 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***

Use a ballpoint or ink pen. Answer in the spaces provided after each question, and should you need further space there are blank pages at the end of the booklet. Please make sure you highlight in the relevant section and question if you have continued to use this additional space to direct the marker of the paper.

**Section One: Multiple-choice 25% (50 Marks)**

This section has **25** questions. Answer **all** questions on the separate Multiple-choice Answer Sheet provided. For each question place a cross in the box to indicate your answer. Use only a blue or black pen to cross the boxes. If you make a mistake, circle the incorrect answer and place a cross in a new box.

**Suggested working time: 50 minutes.**

1. When a candle burns, there are many different chemical processes occurring. Firstly the solid candle wax (C46H92O2) is melted by the heat of the flame. This liquid wax is then drawn up the wick, where the heat of the flame vaporises it. The wax vapour then burns in air to produce the heat and light seen. The equation below represents the physical changes taking place in the candle wax **before** combustion occurs.

**A B**

C46H92O2(s) ⇌ C46H92O2(l) ⇌ C46H92O2(g)

Classify the processes labelled A and B as endothermic or exothermic.

**A B**

1. endothermic endothermic
2. endothermic exothermic
3. exothermic exothermic
4. exothermic endothermic

2. In a chemical reaction at constant temperature, the addition of a catalyst:

(a) increases the concentration of the products at equilibrium.

(b) increases the energy of the molecules so more can successfully collide.

(c) lowers the amount of energy released in the overall reaction.

(d) decreases the time required for equilibrium to be reached.

3. The conjugate base of the acid HPO32– is:

(a) H2PO3–

(b) PO32–

(c) H3PO3

(d) PO33–

**Questions 4 and 5 refer to the information below.**

Consider the following information for a 1.00 mol L–1 solution of Arsenous acid, (H3AsO4):

H3AsO4 (aq) ⇌ H+ (aq) + H2AsO4–(aq)

Ka (at 25°C) = [H+] [H2AsO4–] = 6.6 x 10–10 [H3AsO4]

4. At equilibrium at 25°C, which of the following species will be present in the greatest concentration?

1. H+ (aq)
2. H2AsO4–(aq)
3. H3AsO4 (aq)
4. OH–(aq)

5. Which of the following statements best describe the value of the equilibrium constant (K) for Arsenous acid at 25o C?

1. Arsenous acid is a strong acid existing essentially as molecules.
2. Arsenous acid is a weak acid existing essentially as molecules.
3. Arsenous acid is a weak acid existing essentially as ionic species.
4. Arsenous acid is strong acid existing essentially as ionic species.

1. The pH of a solution was measured with a pH meter during a titration, and was observed to decrease from 4.0 to 2.0. Which of the following statements about the hydrogen ion concentration in the solution is correct?
2. It doubled.
3. It decreased by half.
4. It increased by a factor of 100.
5. It decreased by a factor of 100.

7. The following statements refer to the chemical reaction between magnesium carbonate granules, (MgCO3) and a dilute hydrochloric acid solution, (HCl). Which one of the following statements about this reaction is **FALSE**?

(a) The rate of the reaction decreases with increasing time.

(b) The rate of reaction increases with increasing initial temperature.

(c) The rate of reaction increases with increasing initial concentration of HCl (aq).

(d) The initial rate of reaction is independent of the state of sub-division of MgCO3 (s).

8. Which one of the following statements about the following reversible reaction is **TRUE**?

2SO2(g) + O2 (g) ⇌ 2SO3 (g)

1. K = [SO2]2 [O2]

[SO3]2

(b) K is constant under all reaction conditions.

(c) Sulfur trioxide is being formed when the reaction is at equilibrium.

(d) A catalyst increases the yield of sulfur trioxide by increasing ∆H.

9. In which of the following reactions at equilibrium and at constant temperature is there a shift to the “left” if the pressure of the closed system is increased?

(a) 2NO2 (g) ⇌ N­2O4 (g)

(b) N2 (g) + 3H2 (g) ⇌ 2NH3 (g)

(c) H2O (g) + C (s) ⇌ H2 (g) + CO (g)

(d) H2 (g) + F2 (g) ⇌ 2HF (g)

1. Bromophenol blue is an acid-base indicator that has a colour change from yellow to blue between pH 3.0 and 4.6. A potassium hydroxide solution (in a conical flask), containing a few drops of bromophenol blue indicator, is titrated with an acetic (ethanoic) acid solution (from a burette).

Which one of the following statements about this titration is **TRUE**?

(a) The end point and the equivalence point occur at the same time.

(b) The end point occurs after the equivalence point.

(c) The end point occurs before the equivalence point.

(d) The indicator will be yellow at the equivalence point of the titration.

11. How many moles of electrons are required when the following half-equation is balanced using the smallest possible coefficients?

I2 (s) + H2O (l) ⇌ IO3– (aq) + H+ (aq) + e–

1. 2
2. 5
3. 10
4. 12

12. Consider the statements about the following reaction:

2H2O2 (l)  2H2O (l) + O2 (g)

I H2O2 is reduced.

II H2O2 is oxidised.

III H2O2 acts as a reducing agent.

IV This is not a redox reaction.

Which of the above statements is / are **TRUE**?

(a) IV only

(b) II and III only

(c) I only

(d) I, II and III only

13. Which choice correctly describes the properties of aqueous solutions of the following salts?

|  |  |  |  |
| --- | --- | --- | --- |
|  | Sodium ethanoate  (NaCH3COO) | Potassium nitrate  (KNO3) | Ammonium chloride  (NH4Cl) |
| (a) | neutral | acidic | basic |
| (b) | basic | neutral | acidic |
| (c) | acidic | neutral | basic |
| (d) | basic | acidic | neutral |

**Questions 14 and 15 refer to the information below.**

The Haber process is used in the production of ammonia. It involves the reaction of nitrogen and hydrogen gases in the presence of an iron/iron oxide catalyst. This process is carried out at 350-550 °C and 15-35 MPa. The reaction can be represented by the equation below.

N2(g) + 3 H2(g) ⇌ 2 NH3(g)

14. Which statement is **not** correct regarding the action of a catalyst?

1. A catalyst increases the rate of reaction.
2. A catalyst increases the average kinetic energy of the reactant particles.
3. A catalyst allows a greater proportion of particles to react.
4. A catalyst provides an alternate reaction pathway.

15. The iron/iron oxide catalyst is added to this system **before** it establishes equilibrium. What is the resulting effect?

1. Both forward and reverse reactions would be favoured equally.
2. The forward reaction rate would be increased more than the reverse reaction rate.
3. The system would establish equilibrium faster.
4. The yield of NH3 would be increased.

16. Consider the buffer solution represented by the chemical reaction below:

H2PO4– (aq) + H2O (l) ⇌ HPO42– (aq) + H3O+ (aq)

Which of the following would be **true** after the addition of a small volume of 2.0 mol L-1 sodium hydroxide solution to the buffer solution?

1. The forward reaction rate would be unaffected.
2. The concentration of H2PO4¯ (aq) present in the system would increase.
3. The pH of the system would decrease.
4. The equilibrium would shift to the right.

17. A student had five different 0.2 mol L-1 solutions on her lab bench. They were;

* + - nitric acid, HNO3(aq)
    - zinc chloride, ZnCl2(aq)
    - lithium hydrogencarbonate, LiHCO3(aq)
    - potassium hydroxide, KOH(aq)
    - ammonium chloride, NH4Cl(aq)

Rank these solutions in order of **increasing** pH (i.e. lowest to highest).

1. HNO3 < NH4Cl < ZnCl2 < LiHCO3 < KOH
2. KOH < NH4Cl < ZnCl2 < LiHCO3 < HNO3
3. HNO3 < LiHCO3 < NH4Cl < ZnCl2 < KOH
4. KOH < ZnCl2 < LiHCO3 < NH4Cl < HNO3

18. Hydrogen can be produced by the steam reforming of methane as in the following reaction:

CH4 (g) + H2O (g) ⇌ CO (g) + 3H2 (g) ∆H > 0

Which one of the following will increase the equilibrium yield of hydrogen?

1. Increasing the total pressure of the reaction system.
2. Decreasing the partial pressure of the water vapour.
3. Removing the carbon monoxide from the system as it is produced.
4. Decreasing the temperature of the system.

**Questions 19, 20 and 21 refer to the following information.**

A student was asked to determine the concentration of a solution of hydrofluoric acid that had a concentration of approximately 0.400 mol L–1. He pipetted 20.0 mL of a 0.500 mol L–1 solution of sodium hydroxide into a conical flask, and titrated the hydrofluoric acid against the standardised sodium hydroxide solution, using phenolphthalein as the indicator.

19. What is the pH of the sodium hydroxide solution at the start of the titration?

(a) 13.7

(b) 7.00

(c) 14.0

(d) 12.7

20. If the hydrofluoric acid was added until it was slightly in excess, which of the following pH graphs would show the variation of pH during the titration?

1. pH (c) pH

7

14

Volume of acid added

7

14

Volume of acid added

1. pH (d) pH

7

14

Volume of acid added

7

14

Volume of acid added

21. What approximate volume of hydrofluoric acid would the student expect to have added at the end point of the titration?

(a) 20 mL

(b) 25 mL

(c) 30 mL

(d) 35 mL

**Questions 22 and 23 refer to the following information.**

Consider the following equilibrium which is formed between iron(III) ions and citrate ions, when 50 mL of iron(III) nitrate, Fe(NO3)3(aq), is mixed with 50 mL of sodium citrate, Na3C6H5O7(aq).

Fe3+(aq) + C6H5O73-(aq) ⇌ FeC6H5O7(aq) K = 6.3 x 1011 at 25 °C

*very pale brown colourless yellow*

22. If this system was allowed to establish equilibrium at 25 °C, which of the following would be the **best prediction** of the appearance of the system?

1. The solution would appear yellow.
2. The solution would appear brown-yellow.
3. The solution would appear very pale brown.
4. The colour of the solution cannot be predicted.

23. If a few drops of citric acid (C6H8O7) were added to this equilibrium system, which of the following gives the expected result?

**Favoured direction New concentration of Fe3+(aq)**

1. forward increased
2. forward decreased
3. reverse increased
4. reverse decreased

24. The equation for the complete combustion of butanol is shown below.

C4H10O(l) + 6 O2(g) → 4 CO2(g) + 5 H2O(l)

How does the oxidation number of the element carbon (C) change during this reaction?

1. +2 to +4
2. -2 to +2
3. -2 to +4
4. +2 to 0

25. Consider the diagram below, which shows the various colours of the indicator ‘cresol red’ under different pH conditions.

0 1 2 3 4 5 6 7 8 9 10 11 12 13

pH

|  |  |  |  |
| --- | --- | --- | --- |
| *cresol red* | red | yellow | red |

Which of the following pairs of solutions would **most easily** be distinguished by adding a few drops of cresol red indicator to each?

1. 0.5 mol L-1 NaH2PO4 and 0.5 mol L-1 NaCl
2. 2.0 mol L-1 H2SO4 and 2.0 mol L-1 NaOH
3. 0.25 mol L-1 NH3 and water
4. 0.7 mol L-1 H2CO3 and 0.7 mol L-1 NH4NO3

### END OF SECTION ONE

**Section Two: Short answer 35% (70 Marks)**

This section has **11** questions. Answer **ALL** questions. Write your answers in the question/answer book provided.

**Suggested working time: 60 minutes.**

**Question 26 (6 marks)**

Arsenic acid (H3AsO4) can be produced by reacting solid arsenic trioxide (As2O3) with nitric acid. This produces arsenic acid, as well as the gaseous by-product dinitrogen trioxide.

Write the oxidation and reduction half-equations and the overall redox equation for this reaction, assuming acidic conditions.

|  |  |
| --- | --- |
| Oxidation half-equation |  |
| Reduction half-equation |  |
| Overall redox equation |  |

**Question 27 (4 marks)**

Write observations for any reactions that occur in the following procedures. 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 no change is observed, you should state this.

(Note: No chemical equations necessary).

(a) Some hydrochloric acid solution is mixed with solid sodium carbonate.

(2 marks)

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(b) Some solid copper (II) hydroxide is mixed with a dilute nitric acid solution. (2 marks)

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**Question 28 (6 marks)**

The Brønsted – Lowry theory can be used to account for the acidic and basic properties of a much wider array of substances whose properties cannot be easily explained using earlier theories.

Complete the following table by stating the pH, and give a supporting balanced chemical equation to explain the pH for each of the substances listed.

(6 marks)

|  |  |  |
| --- | --- | --- |
| **Substance** | **pH (acidic, basic or neutral)** | **Equation** |
| Mg(CH3COO)2 (aq) |  |  |
| NH4Cl (aq) |  |  |
| NaHSO4 (aq) |  |  |

**Question 29 (4 Marks)**

The following chemical equation represents an unbalanced redox reaction.

MnO4– (aq) + C2O42– (aq) Mn2+ (aq) + CO2 (g)

In the appropriate spaces below, write the two separate half-equations and the overall balanced redox equation. (4 marks)

Oxidation:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Reduction:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Overall Redox:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Question 30 (6 marks)**

Bromine water, which is a dilute aqueous solution of bromine in water, is slightly acidic because of its reaction with water, represented by the following equation:

Br2 (aq) + H2O (l) ⇌ HBrO (aq) + H+ (aq) + Br –(aq)

Orange Colourless

In aqueous solution, bromine, Br2 (aq) is orange. Hypobromous acid, HBrO (aq), and bromide ions, Br – (aq) are both colourless.

State and explain the colour changes that would be observed, if the following changes are made to the system at equilibrium.

(a) Addition of NaOH (aq). (3 marks)

Colour: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Explanation: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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(b) Addition of excess HCl (aq). (3 marks)

Colour: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Explanation: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

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**Question 31 (5 marks)**

Calculate the pH of the resultant solution, if 25.0 mL of 2.00 mol L–1 sodium hydroxide and 52.0 mL of 1.00 mol L–1 hydrochloric acid are mixed together. (5 marks)

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**Question 32 (9 marks)**

The manufacture of ammonia on an industrial scale is carried out using the Haber process, which relies on the reversible reaction of nitrogen and hydrogen in the presence of an iron catalyst, as shown in the following equation:

N2(g) + 3 H2(g) 2 NH3(g) ΔH = -92 kJ mol–1

The conditions for the reaction in industry must be chosen carefully, taking into consideration not only the yield, but also the rate of the reaction. Commonly, a temperature of around 500°C is used, and the reaction operated at a pressure of around 20,000 kPa. Since ammonia has a much higher boiling point than the other gases, it can easily be removed from the equilibrium mixture by condensation.

(a) In the space provided below, draw a fully labelled enthalpy level diagram for the Haber process, showing **∆H**, **EA**, **catalysed** and **uncatalysed** reaction pathways, and **axes with correct units** stated.

(5 marks)

A sealed vessel containing an equilibrium mixture of nitrogen, hydrogen and ammonia was subjected to the following changes in conditions:

* At a time, t1, the temperature of the vessel was increased
* At a time, eqm1, the system had returned to equilibrium
* At a time, t2, all ammonia was removed from the system
* At a time, eqm2, the system had again returned to equilibrium

1. Complete the following graph, to show what happens to the concentrations of nitrogen and ammonia as the above changes are made.

(4 marks)

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Concentration (molL-1) |  | [N2(g)] |  |  |  |  |
|  |  |  |  |  |  |
|  | [NH3(g)] |  |  |  |  |
|  |  |  |  |  |  |
|  |  | t1 | eqm1 | t2 | eqm2 |  |

**Question 33 (10 marks)**

Aluminium salts are acidic due to the presence of the hexaaqualuminate ion, [Al(H2O)6]3+ which is formed when a soluble aluminium salt is dissolved in water. This ion undergoes hydrolysis as follows:

[Al(H2O)6]3+ (aq) + H2O (l) ⇌ [Al(OH)(H2O)5]2+ (aq) + H3O+ (aq)

1. Write the equilibrium constant (K) expression for this reaction. (1 mark)

|  |
| --- |
|  |

(b) A solution of aluminium nitrate has a pH of 5.6.

1. Using the above equilibrium reaction, explain how the pH of the solution would change, if more crystals of hydrated aluminium nitrate were dissolved into the solution. (3 marks)

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1. When a small volume of dilute sodium hydroxide was added to a sample of the original solution, the pH initially increased from 5.6 to 6.0, and then decreased back to 5.8. Explain these observations. (3 marks)

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(c) It was found that when the aluminium nitrate solution was warmed, the pH of the solution decreased. From this information, deduce whether the forward reaction in the above equilibrium is endothermic or exothermic. Explain your reasoning. (3 marks)

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**Question 34 (6 marks)**

Tellurium (Te) is a rare, silver metalloid that can be used in solar panels and as a semiconducting material. It can be produced by reacting the mineral tellurite (TeO2) with hypophosphorous acid (H3PO2). This produces tellurium metal and phosphorous acid (H3PO3).

Write the oxidation and reduction half-equations and the overall redox equation for this reaction, assuming acidic conditions.

|  |  |
| --- | --- |
| Oxidation half-equation |  |
| Reduction half-equation |  |
| Overall redox equation |  |

**Question 35 (8 marks)**

Phosphoric acid is a weak, **triprotic** acid. In an experiment, a solution of approximately 0.2 mol L–1 phosphoric acid (H3PO4) is titrated with a standard solution of 0.200 mol L–1 sodium hydroxide in order to determine the accurate concentration of the acid. 30.00 mL of the sodium hydroxide solution was pipetted into a conical flask, and the phosphoric acid added from the burette.

1. Write a balanced molecular equation, including state symbols, for the reaction occurring.

(2 marks)

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(b) On the axis below, sketch a graph showing how the pH would be expected to change during the titration, until an excess of the acid was added.

(3 marks)

14

**pH**

7

0

10 20 30

**Volume of H3PO4 Added (mL)**

(c) On the graph above, label the equivalence point for this reaction. (1 mark)

(d) What should the pipette be rinsed with, immediately prior to use? (1 mark)

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(e) From the list below, circle the correct indicator, that would be suitable for use in this particular titration. (1 mark)

**Methyl orange Phenolphthalein Bromothymol blue**

(pH 3.1 – 4.4) (pH 8.3 – 10.0) (pH 6.0 – 7.6)

**Question 36** **(6 marks)**

Use the Standard Reduction Potentials from your Data Booklet to answer the following questions. In each case, write all relevant half-equations with their respective Eo values. (If the reaction is likely to occur, write an overall balanced redox equation with the resultant cell voltage). Then you must state clearly if the reaction is likely or unlikely to occur as described.

1. A piece of aluminium metal is placed in a 1.00 mol L–1 nickel nitrate solution.

(3 marks)

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1. Silver metal is added to a 1.00 mol L–1 sulfuric acid solution.

(3 marks)

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**END OF SECTION TWO**

**Section Three: Extended answer 40% (80 Marks)**

This section contains **five (5)** questions. You must answer **ALL** questions. Write your answers in the question/answer booklet.

Any calculations are to be set out ***in detail*** in the question/answer booklet provided. You may be penalized significantly for failure to show appropriate working, even if you obtain the correct numerical answer. Marks will be allocated for correct equations and clear setting out of a partial answer, even if you cannot complete the problem.

This part carries **80 marks**.

Numerical answers **MUST** be corrected to **THREE (3) SIGNIFICANT FIGURES**.

**Suggested working time: 70 minutes.**

**Question 37 (16 marks)**

Rising carbon dioxide levels in the atmosphere are believed to play an important role in the life of organisms known as calcifiers, a group that includes many forms of coral and crustaceans. These organisms use a precipitation reaction between calcium ions and carbonate ions present in sea-water to form shells and skeletons.

Measurements have detected a fall of around 0.1 in the pH of the oceans since the beginning of the industrial revolution at the end of the 18th century. Scientists believe this acidification can be attributed to an increase in the partial pressure of carbon dioxide in the atmosphere over the same period.

1. Use appropriate chemical equations, to explain why a rise in the partial pressure of carbon dioxide in the atmosphere has caused a decrease in the pH of the oceans. (3 marks)

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A student wished to investigate the composition of prawn shells. In order to do this, the student carried out a series of reactions to convert all the carbonate in the shells, (present as CaCO3), to a soluble form, (i.e. CO32-).

The steps that the student carried out were as follows:

* The shells of 10 prawns were ground to a fine powder using a mortar and pestle.
* 2.17 g of the powder was placed in a beaker, where it was chemically treated to convert all the carbonate into a soluble form.
* The resulting mixture was then filtered to remove any insoluble substances and the filtrate transferred to a 250 mL volumetric flask and made up to the mark with distilled water.
* 20 mL aliquots of the solution in the volumetric flask were titrated against a standard solution of nitric acid with a concentration of 0.0502 mol L–1.
* The student took all burette readings from the **top of the meniscus**.
* The average titre of nitric acid used was 35.05 mL.

1. Write a balanced ionic equation for the titration reaction. (2 marks)

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1. Calculate the number of moles of nitric acid titrated from the burette. (1 mark)

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1. Calculate the number of moles of carbonate in the 20.0 mL aliquots. (2 marks)

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1. Calculate the number of moles of carbonate in the original 2.17 g of powdered prawn shells, and thus calculate the percentage by mass of calcium carbonate in the sample of prawn shells. (You may assume that the moles of CaCO3 are equal to the moles of Na2CO3).

(5 marks)

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1. State and explain what effect the student’s decision to read the burette from the top of the meniscus would have had on the calculated percentage by mass. (3 marks)

|  |  |  |  |
| --- | --- | --- | --- |
| **Effect on calculated percentage (circle one)** | Artificially high | No effect | Artificially low |

Explanation

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**Question 38 (13 marks)**

Silicon dioxide (SiO2) is one of the most abundant substances in the Earth’s crust. It is most commonly found in quartz and many types of sand and because of this, is often found in metal ores that are mined from the Earth’s crust. One of the methods that can be used to remove silicon dioxide from an ore is illustrated in the equation below.

SiO2(s) + H2(g) ⇌ SiO(g) + H2O(g) ΔH = +534 kJ mol-1

This chemical equation represents the equilibrium that forms between silicon dioxide and silicon monoxide. The activation energy for this reaction is 565 kJ mol-1.

(a) Identify the oxidising agent (oxidant) and reducing agent (reductant) in this reaction. Use oxidation numbers to support your answer. (2 marks)

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Quite a low pressure, around atmospheric (approx. 100kPa), is used for this process. However a high temperature of 1550 °C is maintained in the reaction chamber.

(b) Explain, in terms of the collision theory, the effect of each of the following on reaction rate. (5 marks)

|  |  |
| --- | --- |
| Use of high temperature |  |
| Use of low pressure |  |

(c) State the effect (increased, decreased or no change) of each of the following on yield. (2 marks)

|  |  |
| --- | --- |
| Use of high temperature |  |
| Use of low pressure |  |

Since the silicon monoxide is produced in gaseous form, SiO(g), it is easily removed from the system. This in turn helps to favour the forward reaction.

(d) Continue the concentration graph below, by sketching the effect of removing a batch of SiO(g) from the reaction chamber. (4 marks)

Concentration (mol L-1)

Removal of SiO

Equilibrium

re-established

H2

SiO

**Question 39 (22 marks)**

Propanoic acid, CH3CH2COOH, is a weak monoprotic acid that is produced by bacteria in the skin. In an experiment to determine the concentration of an aqueous solution of propanoic acid, a student titrated 25.0 mL aliquots of the solution with a previously standardised 0.976 mol L–1 solution of sodium hydroxide in a conical flask, using a pH meter to monitor the change in pH.

The student’s results are shown in the table below.

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **Volume of NaOH (mL)** | 20.75 | 20.80 | 20.85 | 20.90 | 20.95 | 21.00 | 21.05 | 21.10 | 21.15 |
| **pH of solution** | 4.7 | 5.3 | 5.2 | 5.6 | 7.9 | 12.7 | 13.0 | 13.2 | 13.3 |

1. Explain why a failure to restandardise the sodium hydroxide solution would have led to a systematic error, and what effect it would have on the calculated value for the concentration of the acid. (3 marks)

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1. Plot the results from the experiment on the graph paper provided below, and use your graph to estimate the pH at the equivalence point. Include clearly labelled axes and an appropriate scale. (5 marks)

Estimated pH at equivalence point: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (1 mark)

1. Use an appropriate equation, to describe and explain the pH at the equivalence point of this titration.

(3 marks)

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1. Use an appropriate chemical equation, to describe and explain why the reaction mixture in the flask was able to act as a buffer before less than 20 mL of sodium hydroxide was added. (4 marks)

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After repeating the experiment a number of times, the student found the concentration of the propanoic acid solution was 0.815 mol L–1.

1. Using the data provided, calculate the pH of the mixture in the flask if 30.0 mL of sodium hydroxide is added to a 25.0 mL aliquot of propanoic acid. (6 marks)

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**Question 40 (14 marks)**

When soils containing iron pyrite (FeS2) are exposed to air, the following reaction occurs.

2 FeS2(s) + 7 O2(g) + 2 H2O(l) → 2 Fe2+(aq) + 4 SO42–(aq) + 4 H+(aq)

These types of soils are called acid sulfate soils. The pH of groundwater in these soils will decrease. If this groundwater discharges into lakes and rivers it will also cause their pH to decrease.

1. Explain how this reaction causes the pH of groundwater to decrease.

(2 marks)

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A titration was carried out on a sample of lake water, suspected of being contaminated with acid soils, to determine its pH.

A student took a standardised solution of 0.005 molL–1 NaOH from a volumetric flask and filled the burette.

The student then titrated the NaOH solution against 50.0 mL samples of the lake water and obtained the following results.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Trial 1 | Trial 2 | Trial 3 | Trial 4 |
| Final burette reading (mL) | 4.25 | 8.05 | 12.00 | 16.05 |
| Initial burette reading (mL) | 0.00 | 4.10 | 8.10 | 12.05 |
| Volume of NaOH used (mL) |  |  |  |  |

(b) Determine the average volume of NaOH used. (2 marks)

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(c) Calculate the average number of moles of NaOH used to neutralise the acid. (1 mark)

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(d) Assuming that the lake water is the only source of H+ ions and that complete ionisation of the acid in the lake water has occurred, determine the pH of the lake water. (3 marks)

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(e) Complete the following table (6 marks)

|  |  |  |
| --- | --- | --- |
| Equipment | What is it used for in this experiment? | What should it be rinsed with before use? |
| Burette |  |  |
| Pipette |  |  |
| Conical flask |  |  |

**Question 41 (15 marks)**

(a) When 25.00 mL of 0.1000 mol L-1 silver nitrate solution is mixed with 35.50 mL of 0.0550 mol L-1 sodium carbonate solution a yellow silver salt precipitate is formed according to the following ionic equation:

2Ag+(aq) + CO32- (aq) → Ag2CO3 (s)

Determine the mass of the precipitate formed making sure to justify your choice of limiting reactant.

(5 marks)

(b) Barium hydroxide, Ba(OH)2(s), also known as ‘baryta’ is often found in hydrated form and appears as white crystals. It is corrosive, toxic and moderately soluble in water. It can be used in titrations, for the manufacture of organic substances, as well as in the synthesis of other barium-containing compounds.

A barium hydroxide solution was made by dissolving 1.31 g of solid anhydrous Ba(OH)2 crystals into 795 mL of water.

Calculate the pH of this solution. (5 marks)

Hydrocyanic acid, HCN(aq), is an extremely poisonous acid with a

Ka value of **6.17 x 10-10.**

It is made by dissolving liquid or gaseous hydrogen cyanide in water. Although small amounts of hydrogen cyanide can be extracted from the stones of some fruits such as cherries, apricots and apples, it is generally manufactured on an industrial scale.

(c) What information does the value of Ka give us about hydrocyanic acid, HCN(aq)? Explain your answer. (2 marks)

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(d) Write out the ionization of HCN in water. On this Bronsted-Lowry equation, label the conjugate acid-base pairs. (3 marks)

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### END OF SECTION THREE

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**Spare Page**

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**Spare Page**