

**CHEMISTRY**

**UNIT 3**

**2017**

**MARKING GUIDE**

**Section One: Multiple-choice (50 marks)**

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| 1 | a □ b □ c □ d ■ |  | 11 | a □ b □ c ■ d □ |  | 21 | a □ b □ c ■ d □ |
| 2 | a □ b ■ c □ d □ |  | 12 | a □ b □ c □ d ■ |  | 22 | a ■ b □ c □ d □ |
| 3 | a □ b □ c □ d ■ |  | 13 | a □ b ■ c □ d □ |  | 23 | a □ b □ c ■ d □ |
| 4 | a □ b □ c ■ d □ |  | 14 | a □ b □ c □ d ■ |  | 24 | a □ b □ c ■ d □ |
| 5 | a □ b ■ c □ d □ |  | 15 | a □ b □ c □ d ■ |  | 25 | a □ b □ c □ d ■ |

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| 6 | a □ b □ c ■ d □ |  | 16 | a □ b □ c ■ d □ |  |  |  |
| 7 | a □ b □ c □ d ■ |  | 17 | a □ b □ c □ d ■ |  |  |  |
| 8 | a □ b □ c ■ d □ |  | 18 | a □ b □ c ■ d □ |  |  | (2 marks per question) |
| 9 | a □ b □ c ■ d □ |  | 19 | a ■ b □ c □ d □ |  |  |  |
| 10 | a □ b ■ c □ d □ |  | 20 | a □ b □ c ■ d □ |  |  |  |

**Section Two: Short answer (70 marks)**

**Question 26 (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)

**A white solid dissolves in a colourless solution, producing a colourless and odourless gas.(2)**

***(\*Must have two observations for both marks).***

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

**A blue solid dissolves in a colourless solution to produce a blue solution. (2)**

***(\*Must have two observations for both marks).***

**Question 27 (6 Marks)**

The uptake of carbon dioxide from the atmosphere by the oceans is leading to gradual acidification of the oceans (i.e. the oceans are becoming more acidic). When carbon dioxide dissolves, it reacts with water to form carbonic acid, which in turn forms hydrogen carbonate and then carbonate ions.

1. Write balanced chemical equations showing carbon dioxide reacting with water to form carbonic acid, and then the two successive ionisation reactions that carbonic acid undergo in water. (3 marks)
2. **CO2 (g) + H2O (l)** ⇌ **H2CO3 (aq) (1)**
3. **H2CO3 (aq) + H2O (l)** ⇌ **HCO3- (aq) + H3O+ (aq) (1)**
4. **HCO3- (aq) + H2O (l)** ⇌ **CO32- (aq) + H3O+ (aq) (1)**

One of the most significant consequences of ocean acidification is the effect that it has on shellfish and other marine life that produce calcium carbonate and relies on it as a major component of the exoskeleton or other supporting structure. If the water is sufficiently acidic, the carbonate structures may not form completely. Ocean acidification is thought to lead to a reduction in the availability of carbonate ions. Further reaction of the dissolved carbon dioxide occurs as shown below.

CO2 (g) + CO32– (aq) + H2O (l) ⇌ 2 HCO3– (aq)

(b) Identify a conjugate acid-base pair in this reaction, and explain why it is classified as a Brønsted – Lowry acid-base reaction.

(3 marks)

**Conjugate A/B pair = CO32- / HCO3- (1) \*Also accept HCO3- / CO32-**

**This equation is classified as a Brønsted – Lowry acid-base reaction because in the forward reaction, H2O donates a proton, thus acting as a B-L acid, (1)**

**while CO32- accepts a proton, thus acting as a B-L base. (1)**

**Question 28 (6 Marks)**

The Bronsted-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) | **Basic (1)** | **CH3COO- + H2O** ⇌ **CH3COOH + OH- (1)** |
| NH4Cl (aq) | **Acidic (1)** | **NH4+ + H2O** ⇌ **NH3 + H3O+ (1)** |
| NaHSO4 (aq) | **Acidic (1)** | **HSO4- + H2O** ⇌ **SO42- + H3O+ (1)** |

**\* Also accept “greater than 7” or “less than 7” respectively, for each salt.**

**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: (**C2O42- 2 CO2 + 2e- ) x 5 (1)**

Reduction: (**MnO4- + 8 H+ + 5 e- Mn2+ + 4 H2O) x 2 (1)**

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

Overall Redox: **5 C2O42- + 2 MnO4- + 16 H+ 10 CO2 + 2 Mn2+ + 8 H2O (2)**

**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)

In aqueous solution, bromine, Br2 (aq) is brown. 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: **Brown colour fades, or solution turns less brown. (1)**

Explanation:  **Addition of OH- causes a decrease in the [H+] as the combination of the two ions produce water (H2O). (1) This will result in the rate of collision of reactants being greater than that of products, shifting the equilibrium to the right, favouring the forward reaction rate. Thus the [Br2] decreases causing the brown colour to fade. (1)**

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

Colour:  **Brown colour becomes more intense, or solution becomes more brown. (1)**

Explanation:  **Addition of HCl causes an increase in the [H+] on product side, leading to a higher rate of collision of products than the reactants. (1) This will shift the equilibrium to the left, favouring the reverse reaction, leading to an increase in the [Br2], and the solution becomes more brown. (1)**

**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)

**NaOH + HCl NaCl + H2O**

**n(NaOH) = cV = 2.00 x 0.025 = 0.05 mol (1)**

**n(HCl) = cV = 1.00 x 0.052 = 0.052 mol (1)**

**n(HCl)excess = (0.052 - 0.05) = 0.002 mol (1)**

**[HCl] = n(H+) = 0.002 = 0.025974 mol L-1 (1)**

**VTot 0.077**

**pH solution = -log [H+] = -log (0.025974) = 1.59 (1)**

**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)

**Axes (1)**

**Exo. shape (1)**

**Enthalpy (H) EA & H (1)**

**(kJ mol-1) Fe Catalyst Reactants &**

**3 H2 + N2 EA products (1)**

**EA Catalyst (1)**

**H = -92 kJ**

**2 NH3**

**Progress of Reaction**

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)

**Award (2) marks for showing the correct shape and orientation for the N2 and (2) marks for the correct shape and orientation for the NH3 lines.**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| Concentration/mol L-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)

|  |
| --- |
| **K = [(Al(OH)(H2O)5)2+] [H3O+] (1)**  **[(Al(H2O)6)3+]** |

(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)

**The addition of a soluble Al – salt will lead to an increase in [(Al(H2O)6)3+]. (1)**

**Thus the rate of collision of the reactants will increase, leading to an increase in the forward reaction rate. (1)**

**Consequently leading to a higher [H3O+] and a lowering in the pH. (1)**

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)

**Initially the addition of excess OH- will cause an increase in pH to 6.0. (1)**

**As the neutralisation of OH- and H+ takes place, the rate of collision of reactants will be higher than that of the products, thus the rate of the F’wd reaction is favoured. (1)**

**This will lead to an increase in [H3O+] and thus decrease the pH to 5.8. (1)**

(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)

**As the pH has decreased due to an increase in the [H+], caused by an increase in temp; (1) clearly the F’wd reaction has been favoured by this imposed change, (ie. higher temp). (1)**

**In order for the reaction to respond in this way, (ie. shifting the equilibrium to the right),**

**the F’wd reaction must be ENDOTHERMIC. (1)**

**Question 34 (8 Marks)**

Ethanoic acid is a weak, **monoprotic** acid. In an experiment, a solution of approximately 0.2 mol L–1 ethanoic acid (CH3COOH) 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 ethanoic acid added from the burette.

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

(2 marks)

**CH3COOH (aq) + NaOH (aq) NaCH3COO (aq) + H2O (l) (2)**

***\*Deduct 1 x mark if missing or incorrect state symbols.***

(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

**High pH start (1)**

**Equivalence** **above pH 7 (1)**

**X**

pH

7

**Final pH NOT too low (1)**

**Equivalence to coincide with**

**30 mL CH3COOH added. (1)**

0

30 60 90

Volume of CH3COOH 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)

**The NaOH solution. (1)**

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

**(1)**

**Methyl orange Phenolphthalein Bromothymol blue**

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

**Question 35 (6 Marks)**

Below is a representation of an electrochemical cell, which involves the reaction of hydrogen and chlorine: **e-**

**(1)**

Flow of Hydrogen (H2) gas

Flow of Chlorine (Cl2) Gas

Platinum electrodes

1.0 molL–1 HCl

Salt bridge

1. Give the half equation for the reactions occurring at the anode and at the cathode and then write an overall balanced redox equation for the reaction occurring in the cell. (3 marks)

|  |
| --- |
| Cathode half-equation:  **Cl2 + 2 e- 2 Cl - (1) Eo = + 1.36 V** |
| Anode half-equation:  **H2 2 H+ + 2 e- (1) Eo = 0.00 V** |
| Overall equation:  **Cl2 + H2 2 Cl - + 2 H+ (1)** |

(b) Using the standard reduction potential values from the data sheet, calculate the maximum theoretical voltage (e.m.f.) that could be produced by this cell.

(1 mark)

**E.m.f. = (+1.36) + (0.00) = +1.36 V (1)**

(c) Show the direction of the flow of electrons in the external circuit by means of an **arrow**

“( )” in the diagram above.

**\*See on Diagram above.** (1 mark)

(d) Suggest a reason why platinum, (Pt), is used for the electrodes. (1 mark)

**Platinum is INERT so it will not take part in the reaction. (1)**

**\*Can also accept, “will allow for electron transfer”.**

**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)

**2 x (Al Al 3+ + 3 e-) Eo = +1.68 V**

**3 x (Ni2+ + 2 e- Ni) Eo = - 0.24 V (1)**

**2 Al + 3 Ni2+ 2 Al 3+ + 3 Ni EMF = + 1.44V (1)**

**Positive EMF, thus reaction WILL occur. (1)**

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

(3 marks)

**2 (Ag Ag+ + e-) Eo = - 0.80 V**

**2 H+ + 2e-  H2 Eo = 0.00 V (1)**

**2 Ag + 2 H+ 2 Ag+ + H2 EMF = -0.80 V (1)**

**Negative EMF, thus reaction will NOT occur. (1)**

**\*Note: Overall redox equation NOT necessary, as reaction will not occur.**

**End of Section Two**

Turn to next page

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

This section contains **five (5)** questions. You must answer **all** questions. Write your answers in the spaces provided below.

Where questions require an explanation and/or description, marks are awarded for the relevant chemical content and also for coherence and clarity of expression. Lists or dot points are unlikely to gain full marks.

Final answers to calculations should be expressed to the appropriate number of significant figures.

Spare pages are included at the end of this booklet. They can be used for planning your responses and/or as additional space if required to continue an answer.

* Planning: If you use the spare pages for planning, indicate this clearly at the top of the page.
* Continuing an answer: If you need to use the space to continue an answer, indicate in the original answer space where the answer is continued, i.e. give the page number. Fill in the number of the question(s) that you are continuing to answer at the top of the page.

Suggested working time: 70 minutes.

**Question 37 (16 marks)**

A group of students were given the task of determining the percentage of calcium carbonate in a limestone rock sample that they found on the beach. The sample was known to be made up largely of calcium carbonate, but it also contained traces of sodium carbonate.

The main source of calcium carbonate in limestone is from the remains of shells and the exoskeleton of many other marine organisms, originating from calcification reactions between dissolved calcium and carbonate ions in the water.

1. Write a balanced chemical equation (including states), to represent the information given in the second paragraph above. (2 marks)

***Ca2+(aq) + CO3-2(aq) CaCO3(s) (2)***

The students employed the following procedure to carry out an analysis to determine the percentage by mass of calcium carbonate in the limestone sample.

* A sample of limestone was dried in an oven overnight and then accurately weighed the next morning. The students recorded a constant mass of 20.25 g.
* The sample was then crushed into a fine powder and added to 100 mL of distilled water where it was stirred to dissolve all the sodium carbonate in the mixture.
* The mixture was then filtered and the filtrate was quantitatively transferred to a 250 mL volumetric flask and the solution made up to the mark with distilled water.
* After it was well mixed, 20.0 mL aliquots of the sodium carbonate solution were titrated against a standardised hydrochloric acid solution with a concentration of 0.0908 molL-1.
* The average titre of hydrochloric acid used was 19.25 mL.

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

***Na2CO3 + 2 HCl 2 NaCl + CO2 + H2O (2)***

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

***n(HCl ) = cV = 0.0908 x 0.01925 = 0.001748 mol (1)***

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

***n (Na2CO3) 20mL = ½ n (HCl ) = 0.000874 mol***

1. ***(1)***
2. Calculate the number of moles of sodium carbonate in the 250 mL volumetric flask and thus the mass of sodium carbonate in the original 20.25 g sample of limestone. (4 marks)

***n (Na2CO3) 250mL = 250/20 x 0.000874 = 0.0109245 mol (2)***

***m (Na2CO3) = nM = 0.0109244 x 105.99 = 1.1579 g (2)***

1. From the results obtained in part (e) above, calculate the mass of calcium carbonate in the limestone sample and thus its percentage purity. (3 marks)

***Thus mass CaCO3 = Mass sample – Mass Na2CO3 = 20.25 - 1.158***

***= 19.092 g (1)***

***Therefore % CaCO3 = (19.092 / 20.25) x 100 = 94.3 %***

***(1) (1)***

1. If the students consistently read the burette volumes from the top of the meniscus, this would have led to a **systematic** error. Explain what effect, if any, this would have had on the calculated percentage of calcium carbonate in the limestone. (2 marks)

***No effect on percentage calculated. (1)***

***As both initial and final readings were consistently read from top of meniscus, the actual titrated volumes (obtained by subtraction), were not affected and the results would have been the same.***

***(1)***

**Question 38 (14 marks)**

The electroplating of various metals plays an extremely important role in industry. These reactions can be carried out on a small scale in the laboratory using standard laboratory equipment.

A typical spoon can be chrome electroplated utilising a chromium electrode and an acidified aqueous chromium nitrate solution. Using a labelled diagram, explain the process involved in electroplating the spoon.

Your answer should pay particular attention to the following areas:

1. How the cell can be constructed. (A diagram with clear labels for the anode, cathode, electrolyte, direction of flow of electrons and ions). (6 marks)
2. Describe the processes occurring at each electrode. (Including half-equations). (4 marks)

1. Observations made at each electrode. (2 marks)
2. The role of the electrolyte. (1 mark)
3. An example for the industrial importance or application of the process. (1 marks)

**+ -**

1. **e-**

**Cr Anode Spoon Cathode**

**Correct polarity (1)**

**Correct electrodes (2)**

**Anions Electron directn. (1)**

**Ions directn. (1)**

**Cations Electrolyte (1)**

**Cr(NO3)3 (aq)**

1. **At the cathode, the negative terminal of the cell provides electron for the reduction of chromium ions to chromium metal being deposited on the spoon. (1)**

**Cr3+ (aq) + 3 e- Cr (s) (1)**

**At the anode, the positive terminal of the cell ensures the oxidation of the chromium electrode to produce chromium ions in solution. (1)**

**Cr (s) Cr3+ (aq) + 3 e- (1)**

1. **At the cathode, the mass of the spoon will increase. (1)**

**At the anode, the mass of the electrode will decrease. (1)**

1. **The electrolyte allows the transfer of ions; ie. (Cr3+) cations towards the cathode and (NO3-) nitrate ions towards the anode, in order to balance the charges during the normal operation of the cell. (1)**
2. **Industrial importance of electrolysis includes the coating of a cheap metal with a more noble metal, ie. Jewellery, etc.**

**Or for the coating of iron and other reactive metals with a corrosion resistant metal such as chromium, etc.**

**\*Accept any “one” realistic application! (1)**

**Question 39 (22 marks)**

Ammonia has been used as a source of fertiliser since the early part of the twentieth century. Its manufacture is widely attributed to Fritz Haber, a German chemist, in a process still used today and appropriately called the Haber process. In this process, nitrogen gas and hydrogen gas are combined under suitable conditions to produce ammonia and also liberating heat energy. In addition to its use as a fertiliser, ammonia has also found many other uses in a variety of different products, including applications as a general cleaner and as a weak base.

1. Write a balanced chemical equation for the manufacture of ammonia as described by the Haber process. (2 marks)

***3 H2 (g) + N2 (g) ⇌ 2 NH3 (g) + Heat (2)***

1. Is this reaction exothermic or endothermic? ***Exothermic (1)***

(1 mark)

1. As stated above, ammonia is a weak base. Use relevant chemical theory and an appropriate chemical equation, which demonstrates how ammonia reacts with water, to support this statement. (4 marks)

***NH3 + H2O ⇌ NH4+ + OH –  (2)***

***Ammonia accepts a proton from the water molecule (1)***

***In doing this it is acting as a Bronstead Lowry base (1)***

1. In another application, ammonium salts like ammonium chloride, NH4Cl, are also used as pH buffers. Explain what a buffer is and use a relevant chemical equation which demonstrates how ammonium chloride can act as a buffer when a small volume of dilute sodium hydroxide, NaOH, solution is added. (4 marks)

***A buffer is a substance that is capable of resisting changes in pH when a small volume of dilute acid or base is added.***

***NH4+ + H2O ⇌ NH3 + H3O+ (1)***

***When a small volume of OH – is added, this will neutralise the hydronium ions in the reaction above as water molecules are formed, lowering the [H3O+]. (1)***

***This will shift the equilibrium to the right as collisions amongst the reactants are higher, thus favouring the forward reaction. (1)***

***As the hydronium ion concentration is partially replaced, the pH will not increase as a result of the imposed change. (1)***

An inquisitive chemistry student wanted to monitor the change in pH of a typical ammonia cleaner, as he carried out a titration using a pH meter and other general volumetric apparatus. A 250 mL bottle of liquid ammonia was purchased from a local supermarket and the following procedure was carried by the student:

STEP 1. 25.0 mL of the cleaner was pipetted from its container into a 500 mL volumetric flask.

STEP 2. The solution was made up to the mark with distilled water.

STEP 3. A 20.0mL aliquot of this solution was then titrated against a standardised solution of 0.0490 molL-1 HCl from a burette.

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

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **Volume of HCl (mL)** | 15.60 | 15.70 | 15.80 | 15.90 | 16.00 | 16.10 | 16.20 | 16.30 | 16.40 |
| **pH of solution** | 8.5 | 8.3 | 8.1 | 8.0 | 3.5 | 3.3 | 3.2 | 3.0 | 2.9 |

1. Explain why the ammonia solution was diluted? (2 marks)

***The ammonia solution was diluted as a “dilute" solution of hydrochloric acid was used. (1)***

***If this was not done, an excessively large volume of the acid would have been required to cause any appreciable change in pH. (1)***

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)

**pH**

**13**

**11**

**Equivalence (1)**

**9**

**Axes (1)**

**7 Scale (1)**

**Points (1)**

**5 Smooth curve (1)**

**3**

**1**

**0**

**15.5 15.6 15.7 15.8 15.9 16.0 16.1 16.2 16.3 16.4 Volume HCl (mL)**

Estimated pH at equivalence point: ***6.0 (1) accept reasonable estimate***

(1 mark)

1. Use an appropriate equation, to describe and explain the pH at the equivalence point of this titration. (3 marks)

***Salt formed from the reaction is Ammonium chloride, NH4Cl (1)***

***When this salt hydrolyses in water, the ammonium ion will react as follows:***

***NH4+ + H2O ⇌ NH3 + H3O+ (1)***

***As hydronium ions are formed, the pH at the equivalence point is acidic. (1)***

**Question 40 (14 marks)**

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

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)

**As the reaction proceeds, H+ are produced, thus increasing [H+], (1)**

**and DECREASING pH. (1)**

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

A student placed a standardised solution of 0.005 molL–1 NaOH in 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) | **4.25** | **3.95** | **3.90** | **4.00** |

**Calculated titres in Table (1)**

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

**Av Titre = 3.95 + 3.90 + 4.00 = 3.95 mL (1)**

**3**

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

**n = cV = 0.0050 x 0.00395 = 1.975 x 10-5 mol (3 x SF) (1)**

(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)

**n(H+) = n(NaOH) = 1.975 x 10-5 mol (1)**

**[H+] = n/V = 1.975 x 10-5 / 0.050 = 3.95 x 10-4 molL-1 (1)**

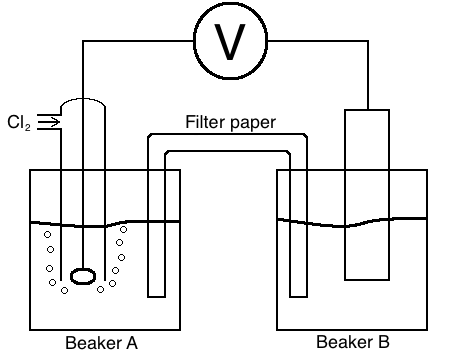
**pH = -log[H+] = -log (3.95 x 10-4) = 3.40 (3 x SF) (1)**

(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 | **To deliver accurate volume of NaOH. (1)** | **The NaOH solution. (1)** |
| Pipette | **To measure 50.0 mL of lake water. (1)** | **The lake water. (1)** |
| Conical flask | **Where the titration reaction takes place. (1)** | **Distilled water. (1)** |

**Question 41 (14 marks)**

The cell, Cu(s) / Cu2+(aq) and C2(g) / C–(aq) with a platinum electrode, was set up as shown in the diagram below. **Beaker A** contained a 1.00 mol L-1 aqueous solution of ammonium chloride, and the filter paper shown in the diagram was soaked in an aqueous solution of potassium nitrate before being placed in the two beakers.



**Cathode (+)**

**(1)**

**Anode (-)**

**(1)**

Pt Electrode

1. Give the name or formula of a suitable electrolyte for use in **Beaker B**. (1 mark)

**Suitable electrolyte = Copper (II) nitrate or Cu(NO3)2 solution (1)**

1. Label the **anode** and **cathode** in the diagram above, including their respective **polarities**. (2 marks)
2. Give **two** reasons why potassium nitrate was a suitable material for soaking the filter paper.

(2 marks)

1. **KNO3 is a “strong electrolyte”, thus a high concentration of ions available for transfer between cells to balance the charge. (1)**
2. **Neither ion, (K+) nor (NO3-), will form a precipitate with other ions. (1)**
3. Calculate the maximum theoretical EMF you could measure for the cell. (2 marks)

**EMF = (+1.36) + (-0.34) = + 1.02 V (2)**

1. Give **one** reason why the measured cell potential might differ from the value calculated in

part (d) above. (1 mark)

**Concentrations may not be 1.0 mol L-1, or Cl2 (g) may not be at STP,**

**or reaction not carried out at 250 C. \*Accept any one valid reason. (1)**

1. Describe the changes that would be observed in **Beaker B** during the operation of the cell?

(2 marks)

**Blue colour of solution would intensify. (1)**

**Mass of salmon pink electrode would decrease. (1) Do NOT accept “dissolve”.**

1. Using relevant chemical theory and a chemical equation, state and explain how the voltmeter reading would change if a few drops of silver nitrate solution were placed in **Beaker A**. (4 marks)

**The introduction of Ag+ ions in Beaker A would cause the following reaction to occur:**

**Ag+ (aq) + Cl - (aq) AgCl (s) (1)**

**The silver ions (Ag+) would remove chloride ions from solution, thus favouring the forward reaction and more chlorine (Cl2) to dissolve in order to re-establish equilibrium. (1)**

**This would cause an INCREASE in the voltmeter reading, (cell EMF), (1)**

**as more electrons would be required for the reduction of chlorine. (1)**

**\*Can also accept other valid explanations; (i.e. more electrons would be required for f’wd or reduction reaction).**

**End of question**