**Multiple Choice ( 19 Marks )**

1. Which option correctly identifies the conjugate acid-base pairs in the following reaction?

CH3COOH (aq) + H2O (**l**) ↔ H3O+ (aq) + CH3COO ─ (aq)

1. CH3COOH and H2O
2. CH3COOH / CH3COO─  and H3O+ / H2O
3. CH3COOH / H2O and H3O+ / CH3COO ─
4. CH3COOH and CH3COO ─
5. Pure water undergoes self-ionisation according to the following equation:

2 H2O (**l**) ↔ H3O+ (aq) + OH - (aq)

The equilibrium constant (Kw) for this reaction is 2.92 x 10-15 at a temperature of 283K.

What is the pH of the water at this temperature?

1. 6.73
2. 7.00
3. 7.27
4. 14.80
5. Consider the following reaction : HCN(aq) + NH3(aq) ⇄ CN―(aq) + NH4+(aq)

Which of the species in this equilibrium mixture are ***acting as bases***?

* 1. HCN(aq) and NH4+(aq)
  2. NH3(aq) and CN―(aq)
  3. HCN(aq) and CN―(aq)
  4. NH3(aq) and NH4+(aq)

1. The pH of an aqueous solution registers 11.0 on a pH meter. Which of the following solutions could be its identity?
   1. 0.0010 molL-1 KOH
   2. 0.0100 molL-1 NaOH
   3. 0.0010 molL-1 HCl
   4. 0.0050 molL-1 Ca(OH)2
2. Consider the table below regarding the acidity constants (measured at 25 oC) of three monoprotic organic acids.

|  |  |
| --- | --- |
| Propanoic acid (CH3CH2COOH) | Ka = 1.3 x 10-5 |
| Ethanoic acid (CH3COOH) | Ka = 1.8 x 10-5 |
| Methanoic acid (HCOOH) | Ka = 1.8 x 10-4 |

From this data, we could conclude that

1. as the acid’s number of carbon atoms increases, the strength of the acid increases.
2. increasing the temperature will increase the strength of the acids.
3. solutions of methanoic acid will always have a lower pH than solutions of propanoic acid.
4. a 0.1 mol L–1 solution of propanoic acid will have a higher pH than a 0.1 mol L–1 solution of methanoic acid.
5. The autoionization of water can be represented by the equation below.

H2O(l) + H2O(l) + heat ⇌ H3O+(aq) + OH-(aq)

Distilled water at a temperature of 15 °C would have

1. a concentration of hydronium ions greater than 1.0 x 10-7 mol L-1.
2. a concentration of hydroxide ions greater than 1.0 x 10-7 mol L-1.
3. a Kw value greater than 1.0 x 10-14.
4. a pH greater than 7.
5. Which statement best describes the equivalence point in a titration between a strong acid and a strong base?
6. The point at which equal moles of H+ ions and OH- ions have been added together.
7. The point at which equal moles of acid and base have been added together.
8. The point at which the first sign of a colour change occurs.
9. The point at which the rate of the forward reaction equals the rate of the reverse reaction.
10. Four beakers (A, B, C and D) were placed on a laboratory bench, each containing distilled water and a pH meter was used to measure their pH. A small sample of a different salt was then dissolved into each beaker, according to the table below.

|  |  |  |  |
| --- | --- | --- | --- |
| **Beaker A** | **Beaker B** | **Beaker C** | **Beaker D** |
| + NH4NO3(s) | + MgF2(s) | + CH3COOK(s) | + Na3PO4(s) |

The pH of the solution in beaker

1. A would be above 7.
2. B would be below 7.
3. C would be below 7.
4. D would be above 7.
5. Which one of the following underlines species is acting as an acid?
   1. CH3CH2CH2CH2NH2 + CH3COOH ⇌ CH3CH2CH2CH2NH3+ + CH3COO-
   2. HSO3- + NH3 ⇌ SO32- + NH4+
   3. NH4+ + CH3COO- ⇌ NH3 + CH3COOH
   4. [Fe(H2O)6]3+ + H2O ⇌ [Fe(OH)(H2O)5]2+ + H3O+
6. Which of the following classifications is correct?

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | **KCl** | **CH3COOK** | **NH4Cl** | **KHSO4** |
| a) | Neutral | Basic | Acidic | Acidic |
| b) | Neutral | Basic | Acidic | Basic |
| c) | Acidic | Acidic | Basic | Basic |
| d) | Neutral | Acidic | Basic | Acidic |

Questions 11 and 12 refer to the table below.

|  |  |  |
| --- | --- | --- |
| Name of indicator | pH range | Colour (low pH – high pH) |
| 1. Methyl red | 4.4 – 6.2 | Red- yellow |
| 1. Bromothymol blue | 6.0 – 7.6 | Yellow – blue |
| 1. Phenolphthalein | 8.3 – 10.0 | Colourless – pink |
| 1. Methyl violet | 0.0 – 2.0 | Yellow – violet |

1. A chemist uses a 0.1034 molL-1 sodium hydroxide solution to standardise a nitric acid solution. Which of the following indicators would be suitable?
   1. 2 only.
   2. 2, 3 and 4 only.
   3. 1, 2 and 3 only.
   4. All of 1, 2, 3 and 4.
2. If methyl red is used in a titration between ethanoic acid (added from burette) and a standard solution of sodium hydroxide (in a conical flask with indicator) then:
   1. The end point of the titration would occur after the equivalence point.
   2. The end point would occur at the equivalence point of the titration.
   3. No colour change would occur.
   4. The end point of the titration would occur before the equivalence point has been reached.
3. What would be the **most** likely pH of a 0.10 mol L-1 solution of sulfuric acid?
4. Less than 0.5
5. Between 0.5 and 1
6. Exactly 1
7. Approximately 1.5
8. Which of the following does **not** contribute to the problems faced by calcifying species as a direct result of ocean acidification?
   1. A decrease in ocean CO32-(aq) concentration.
   2. A decrease in ocean Ca2+(aq) concentration.
   3. An increase in ocean H3O+(aq) concentration.
   4. A decrease in the presence of CaCO3(s).
9. Which row of the table describes what happens when a solution of a weak acid is diluted?

(Assume constant temperature)

|  |  |  |
| --- | --- | --- |
|  | **Ka** | **Extent of Acid Ionisation** |
|  | Decreases | Increases |
|  | Decreases | Decreases |
| (c) | Remains the same | Increases |
| (d) | Remains the same | Decreases |

1. Diagram, rectangle

   Description automatically generatedWhich of the following **correctly** identifies the labels represented by X, Y and Z?

|  |  |  |  |
| --- | --- | --- | --- |
|  | **X** | **Y** | **Z** |
| (a) | Concentration | Volume of acid added | Equivalence point |
| (b) | pH | Volume of base added | Indicator end point |
| (c) | Volume of acid added | Equivalence point | pH |
| (d) | pH | Volume of base added | Equivalence point |

Use the following information to answer Questions 17 and 18.

*A chemist is titrating a volume of an unknown monoprotic acid against 50mL of 0.30M NaOH, using methyl red as an indicator. The chemist observes the first permanent colour change at 23.65 mL.*

1. A valid conclusion that can be drawn from this information is that:
   1. the concentration of the unknown compound is 0.14 M.
   2. the concentration of the unknown compound is 0.5M.
   3. the concentration of the unknown compound is 0.28M.
   4. the concentration of the unknown compound is 1.0M.
2. If the titration is repeated several times, averaging the results will reduce the
   1. accuracy of the results.
   2. reliability of the results.
   3. effect of random errors.
   4. effect of systematic errors
3. A student used the following method to titrate an acetic acid solution of unknown concentration with a standardised solution of dilute sodium hydroxide:

• Rinse burette with deionised water.

• Fill burette with sodium hydroxide solution.

• Rinse pipette and conical flask with acetic acid solution.

• Pipette 25.00 mL of acetic acid solution into conical flask.

• Add appropriate indicator to the conical flask.

• Titrate to endpoint and record volume of sodium hydroxide solution used.

Compared to the actual concentration of the acetic acid, the calculated concentration will be

1. Lower
2. Higher
3. the same
4. different, but higher or lower cannot be predicted.

**END OF MULTIPLE CHOICE QUESTIONS**

**Short Answer ( 32 Marks )**

1. Malic acid (H2C4H4O5) is a weak, diprotic acid. The equation for the first stage of dissociation is shown below.

H2C4H4O5(aq) + H2O(l) ⇌ HC4H4O5-(aq) + H3O+(aq)

* 1. Write the equation for the second stage of ionisation of malic acid. (2 marks)

HC4H4O5-(aq) + H2O(l) ⇌ C4H4O5-2(aq) + H3O+(aq)

1 mark for reactants and products

1 mark for states

* 1. Malic acid can completely react with a solution of potassium hydroxide to form a salt. Will the salt formed be acidic, neutral, or basic? Justify your answer using chemical equations.

(4 marks)

|  |  |
| --- | --- |
| **Description** | **Marks** |
| Salt formed is potassium malate (K2C4H4O5) à 2K+ + C4H4O52– | 1 |
| The K+ ions are neutral and do not undergo hydrolysis | 1 |
| The C4H4O52– ion undergoes hydrolysis  C4H4O52– + H2O(ℓ) ⇌ HC4H4O51– + OH1– (aq) | 1 |
| Salt formed is basic due to the increased concentration of OH1– ions from the equation above. (i.e. hydrolysis reactions cause [OH1–] > [H+] ; | 1 |
| **Total** | **4** |

* 1. Calculate the pH of the potassium hydroxide solution used in part b, given that 0.130g was dissolved into 25.0mL Assume temperature of 25°C.

1. n(OH) = n(KOH) = 0.130 / 56.108 = 0.00231696014
2. [OH-] = n/0.025 = 0.09267840593

pOH = 1.03

1. pH = 12.966 = 13.0

(3 marks)

1. Propanoic acid, CH­3CH2COOH, is a weak monoprotic acid. An acidic buffer solution is prepared by reacting 40.0mL of 1.00molL-1 propanoic acid with 20.0mL of 1.00molL-1 potassium hydroxide solution.
   1. Write an equation for the buffer system that was produced.

CH­3CH2COOH + H2O(l) ⇌ CH3CH2COO-(aq) + H3O+(aq)

* + 1. marks)

1 mark for reactants and products

1 mark for states

* 1. Explain why the two quantities specifically were used to make this buffer solution.

(3 marks)

|  |  |
| --- | --- |
| **Description** | **Marks** |
| This quantity of KOH solution converts half of the ethanoic acid into the salt (ethanoate ion) | 1 |
| The resulting mixture has equal moles of both ethanoic acid and ethanoate ions | 1 |
| High and equal concentrations of weak acid and conjugate base give rise to buffers with the highest buffering capacity. | 1 |
| **Total** | **3** |

* 1. Using the chemical equation from parts a, and the collision theory, explain how this buffer solution operates when a few drops of sodium hydroxide (NaOH) are added.

(3 marks)

|  |  |
| --- | --- |
| **Description** | **Marks** |
| Adding a few drops of concentrated sodium hydroxide solution increases [OH–] and this instantly reduces [H3O+] | 1 |
| This causes the frequency of **successful** collisions between H3O + and CH3COO– ions to decrease reducing rate of the reverse reaction | 1 |
| For a period of time the forward reaction rate is faster than the reverse and so the [H3O +] comes back up, resisting pH change | 1 |
| **Total** | **3** |

If solution to part a is incorrect, give marks for discussion that is relevant and correct for the equation they’ve written

1. Oxalic acid is a diprotic acid, of which crystals are commonly used as a primary standard in acid-base volumetric analysis.
   1. List **one** characteristics you expect oxalic acid crystals to have to justify this classification as a suitable primary standard.

(1 mark)

High purity / high molar mass / not very hydroscopic / not deliquescent / does not react with substances in the atmosphere / highly soluble / predictive reactivity

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Final reading (mL)** | 20.65 | 19.60 | 20.75 | 21.85 |
| **Initial reading (mL)** | 4.45 | 3.90 | 4.95 | 5.95 |
| **Titration volume (mL)** | 16.2 | 15.7 | 15.8 | 15.9 |

* 1. Write the equation of the reaction taking place during the titration.

(1 mark)

H2C2O4 + 2NaOH à 2H2O + Na2C2O4

* 1. Complete the table and determine the average titre value for the oxalic acid solution.

(1 mark)

15.8 mL

* 1. Calculate the mass of oxalic acid in the 2.50g sample. (4 marks)

Moles NaOH = 0.025mL x 0.100M = 0.0025 mol (1)

Moles H2C2O4 =  0.00250/2  =  0.00125 mol (1)

Moles H2C2O4 in 250mL =   250/15.8 = 0.0198 mol (1)

Mass H2C2O4  =  0.0198 x 90.04  =  1.78g (1)

* 1. Use your result to determine the value of ‘n’ in H2C2O4.**n**H2O. (3 marks)

In 2.50 g sample, mass of water = 2.50 – 1.782  =  0.717g       (1)

Moles water = 0.717/18 = 0.0398 mol                                         (1)

Mole ratio:        0.0398 mol H2O / 0.0198 mol H2C2O4    =  2      (1)

1. Acids and bases exist as conjugate acid–base pairs. Below is a table showing the Ka value for a number of acids and the Kb for the corresponding conjugate bases.

|  |  |  |  |
| --- | --- | --- | --- |
| **Some Conjugate Acid–Base Pairs at 25 °C** | | | |
| **Acid** | **Ka** | **Base** | **Kb** |
| HF | 6.8 x 10–4 | F– | 1.5 x 10–11 |
| CH3COOH | 1.8 x 10–5 | CH3COO– | 5.6 x 10–10 |
| H2CO3 | 4.3 x 10–7 | HCO3– | 2.3 x 10–8 |
| NH4+ | 5.6 x 10–10 | NH3 | 1.8 x 10–5 |
| HCO3– | 5.6 x 10–11 | CO32– | 1.8 x 10–4 |

1. “The stronger the acids, the stronger the conjugate base.” State whether this statement is true or false, justifying your answer using the data provided.

(2 marks)

|  |  |
| --- | --- |
| **Description** | **Marks** |
| False | 1 |
| Refer to data above to show that stronger acids with large Ka values produce weaker bases with small Kb values | 1 |
| **Total** | **2** |

1. Use the dihydrogen phosphate ion to demonstrate conjugate acid-base pairing with a labelled equation.

(2 marks)

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

A CB

1- equation 1- conjugate pair labelled

**END OF TEST**