**A-LEVEL CHEMISTRY**

**PAPER 1**

**PRACTICE PAPER 4**

Answer all questions

Max 105 marks

2 hours

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|  | Name …………………………………………………………….. |  |
|  | Mark ……../105 ……....% Grade ……… |  |

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| **2.** |  |
|  | (Total 5 marks) |
| **3.** | (Total 4 marks) |

**4.**       Hess’s Law is used to calculate the enthalpy change in reactions for which it is difficult to determine a value experimentally.

(a)     State the meaning of the term *enthalpy change*.

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**(1)**

(b)     State Hess’s Law.

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**(1)**

(c)     Consider the following table of data and the scheme of reactions.

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| Reaction | Enthalpy change / kJ mol–1 |
|             HCl(g)  →        H+(aq)  +          Cl–(aq) | –75 |
|             H(g)     +          Cl(g)    →        HCl(g) | –432 |
|     H(g)   +     Cl(g)    →    H+(g)    + Cl–(g) | +963 |

 

Use the data in the table, the scheme of reactions and Hess’s Law to calculate a value for ∆*H*r

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 **(3)**

**(Total 5 marks)**

**5.** (a)     Write an equation for the process that has an enthalpy change equal to the electron affinity of chlorine.

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(b)     In terms of electrostatic forces, suggest why the electron affinity of fluorine has a negative value.

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(c)     (i)      Complete the Born–Haber cycle for silver fluoride by adding the missing species on the dotted lines.

                   

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(ii)     Use the cycle in part (i) and the data in the table to calculate a value, in kJ mol–1, for the bond enthalpy of the fluorine–fluorine bond.

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| --- | --- | --- |
|   | **Enthalpy change** | **Value / kJ mol–1** |
|   | Enthalpy of atomisation for silver | +298 |
|   | First ionisation energy for silver | +732 |
|   | Electron affinity for fluorine | –348 |
|   | Experimental enthalpy of lattice dissociation for silver fluoride | +955 |
|   | Enthalpy of formation for silver fluoride | –203 |

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(d)     A theoretical value for enthalpy of lattice dissociation can be calculated using a perfect ionic model.

The theoretical enthalpy of lattice dissociation for silver fluoride is +870 kJ mol–1.

(i)      Explain why the theoretical enthalpy of lattice dissociation for silver fluoride is different from the experimental value that can be calculated using a Born–Haber cycle.

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(ii)     The theoretical enthalpy of lattice dissociation for silver chloride is +770 kJ mol–1.

Explain why this value is less than the value for silver fluoride.

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**(Total 12 marks)**

**6.** In this question, give all pH values to 2 decimal places.

(a)     (i)      Write expressions for the ionic product of water, *K*w, and for pH.

KW = ….................................................................................................

pH = ....................................................................................................

(ii)     At 318 K, the value of *K*w is 4.02 × 10–14 mol2 dm–6 and hence the pH of pure water is 6.70
State why pure water is not acidic at 318 K.

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(iii)     Calculate the number of moles of sodium hydroxide in 2.00 cm3 of 0.500 mol dm–3 aqueous sodium hydroxide.

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(iv)    Use the value of *K*w given above and your answer to part (a)(iii) to calculate the pH of the solution formed when 2.00 cm3 of 0.500 mol dm–3 aqueous sodium hydroxide are added to 998 cm3 of pure water at 318 K.

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**(6)**

(b)     At 298 K, the acid dissociation constant, *K*a, for propanoic acid, CH3CH2COOH, has the value 1.35 × 10–5mol dm–3.

(i)      Write an expression for *K*a for propanoic acid.

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(ii)     Calculate the pH of 0.125 mol dm–3 aqueous propanoic acid at 298 K.

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(c)     Sodium hydroxide reacts with propanoic acid as shown in the following equation.

NaOH + CH3CH2COOH → CH3CH2COONa + H2O

A buffer solution is formed when sodium hydroxide is added to an excess of aqueous propanoic acid.

(i)      Calculate the number of moles of propanoic acid in 50.0 cm3 of 0.125 mol dm–3 aqueous propanoic acid.

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(ii)     Use your answers to part (a)(iii) and part (c)(i) to calculate the number of moles of propanoic acid in the buffer solution formed when 2.00 cm3 of 0.500 mol dm–3 aqueous sodium hydroxide are added to 50.0 cm3 of 0.125 mol dm–3 aqueous propanoic acid.

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(iii)     Hence calculate the pH of this buffer solution at 298 K.

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**(Total 16 marks)**

**7.** An experiment was carried out to measure the e.m.f. of this cell.

Al(s)   Al3+(aq)   H+(aq)   H2(g)   Pt(s)

(a)     The aluminium used as the electrode is rubbed with sandpaper prior to use.

Suggest the reason for this.

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**(1)**

(b)     Draw a labelled diagram of a suitable apparatus for the right-hand electrode in this cell. You do **not** need to include the salt bridge or the external electrical circuit.

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(c)     A simple salt bridge can be prepared by dipping a piece of filter paper into potassium carbonate solution. Explain why such a salt bridge would **not** be suitable for use in this cell.

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**(Total 5 marks)**

**8.** White phosphorus (P4) is a hazardous form of the element. It is stored under water.

(a)     Suggest why white phosphorus is stored under water.

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(b)     Phosphorus(V) oxide is known as phosphorus pentoxide.
Suggest why it is usually represented by P4O10 rather than by P2O5

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(c)     Explain why phosphorus(V) oxide has a higher melting point than sulfur(VI) oxide.

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(d)     Write an equation for the reaction of P4O10 with water to form phosphoric(V) acid.
Give the approximate pH of the final solution.

Equation ........................................................................................................

pH ...................................................................................................................

**(2)**

(e)     A waste-water tank was contaminated by P4O10. The resulting phosphoric(V) acid solution was neutralised using an excess of magnesium oxide. The mixture produced was then disposed of in a lake.

(i)      Write an equation for the reaction between phosphoric(V) acid and magnesium oxide.

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**(1)**

(ii)     Explain why an excess of magnesium oxide can be used for this neutralisation.

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**(1)**

(iii)     Explain why the use of an excess of sodium hydroxide to neutralise the phosphoric(V) acid solution might lead to environmental problems in the lake.

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**(1)**

**(Total 9 marks)**

**9.**          Aqueous metal ions can be identified by test-tube reactions.

          For each of the following, describe what you would observe.

          Write an equation or equations for any reactions that occur.

(a)     The addition of aqueous sodium carbonate to a solution containing
[Fe(H2O)6]3+(aq) ions.

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(b)     The addition of aqueous sodium hydroxide, dropwise until in excess, to a solution containing [Al(H2O)6]3+(aq) ions.

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(c)     The addition of dilute aqueous ammonia, dropwise until in excess, to a solution containing [Cu(H2O)6]2+(aq) ions.

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**(4)**

(d)     The addition of concentrated hydrochloric acid, dropwise until in excess, to a solution containing [Cu(H2O)6]2+(aq) ions.

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**(2)**

**(Total 14 marks)**

**10.** (a)     Use data from the table below to explain why dilute hydrochloric acid cannot be used to acidify potassium manganate(VII) in a titration.

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| --- | --- | --- | --- | --- |
|    |  |  |  | ***E* / V** |
|   | MnO4–(aq) + 8H+(aq) + 5e– | → | Mn2+(aq) + 4H2O(I) | +1.51 |
|   | Cl2(aq) + 2e– | → | 2Cl–(aq) | +1.36 |
|   | 2H+(aq) + 2e– | → | H2(aq) | 0.00 |

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**(2)**

(b)     Use information from the table in part (a) to determine the minimum volume, in cm3, of 0.500 mol dm–3 sulfuric acid that is required for a titre of 25.0 cm3 of 0.0200 mol dm–3 potassium manganate(VII) solution.
Show your working.

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**(3)**

(c)     In each titration using potassium manganate(VII), a large excess of dilute sulfuric acid is used to avoid any possibility of the brown solid MnO2 forming.

(i)      Deduce a half-equation for the reduction of MnO4– ions in acidic solution to form MnO2.

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**(1)**

(ii)     Give **two** reasons why it is essential to avoid this reaction in a titration between potassium manganate(VII) and iron(II) ions.

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(d)     Potassium manganate(VII) is an oxidising agent.
Suggest **one** reason why a 0.0200 mol dm–3 solution of potassium manganate(VII) does **not** need to be kept away from flammable material.

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**(1)**

**(Total 9 marks)**

**11.** Ethanedioic acid is an important industrial chemical with a number of uses.

Ethanedioate ions, C2O42–, act as bidentate ligands with transition metal ions.

(a)     Write an equation for the ligand substitution reaction of an excess of ethanedioate ions with aqueous cobalt(II) ions.

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**(1)**

(b)     The table below shows some standard electrode potentials.

|  |  |  |  |
| --- | --- | --- | --- |
|   |  |  | **E / V** |
|   | Fe3+(aq) + e– | → Fe2+(aq) | +0.77 |
|   | 2CO2(g) + 2e– | → C2O42–(aq) | –0.49 |

Use *E* values from the table to explain why an iron(III) complex is **not** formed when solutions containing ethanedioate ions and iron(III) ions are mixed.

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**(2)**

**(Total 3 marks)**

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| **12.** |  |
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|  |  | (i) |  |
|  |  | (ii) |  |
|  |  | (iii) |  |
|  | (c) |  |
|  | (d) |  |
|  | (e) |  |
|  | (f) |  |
|  | (g) |  |
|  | (h) | **(Total 14 marks)** |