UNIVERSITI KUALA LUMPUR

MALAYSIAN INSTITUTE OF CHEMICAL AND BIO-ENGINEERING TECHNOLOGY

FINAL EXAMINATION

OCTOBER 2025 SEMESTER

COURSE CODE	:	CLB12203
COURSE TITLE	:	PHYSICAL CHEMISTRY
PROGRAMME NAME	:	TECHNICAL FOUNDATION - BACHELOR
DATE	:	02 FEBRUARY 2026
TIME	:	2:00PM - 5:00PM
DURATION	:	3 HOURS

INSTRUCTIONS TO CANDIDATES

1. Please read the instructions given in the question paper CAREFULLY.
2. This question paper is printed on both sides of the paper.
3. This question paper consist of ONE sections.
4. Section A consist of five questions. Answer FOUR (4) questions only.
5. Please write your answer on the answer booklet provided.
6. Please answer all questions in English only.
7. Refer to the attached Formula/ Appendies. ☑ *Tick if applicable*

THERE ARE 10 PAGES OF QUESTIONS INCLUDING THIS PAGE

SECTION A (Total: 100 marks)

Answer FOUR (4) questions.

Please use the answer booklet provided.

Question 1

Show all calculation steps clearly and state the units for all intermediate and final answers.

(a) Helium gas in a rigid container has an initial pressure of 600 kPa at 450 K. After cooling, the pressure drops to 360 kPa. Assuming ideal gas behavior, find the final temperature in °C.

(4 marks)

(b) The van der Waals' equation is given by:

$$\left(P + a(\frac{n}{V})^2\right)(V - nb) = nRT$$

0.50 mol of N_2 confined in a volume of 10.0 dm^3 at 27°C. The van der Waals constant for N_2 and the gas constant are: a = 1.39 atm. L^2/mol^2, b = 0.0391 L/mol and R = 0.0821 atm.L/K.mol. Determine the pressure exerted by N_2 molecules, if the gas obeys:

i. ideal gas equation

(3 marks)

ii. real gas equation

(3 marks)

(c) A sample of 700 g of octane (C_8H_{16}) is combusted in a bomb calorimeter containing 5.65 dm^3 of water at initial room temperature. After the reaction, the temperature rises to 55.4°C. Calculate the heat of combustion of octane in kJ/mol, if the heat capacity of the calorimeter is known to be 10.4 kJ/°C and the specific heat capacity of water is 4.184 J/g°C.

(7 marks)

(d) The following reaction presents the combustion of methane:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2 + 2H_2O(g)$$

Using the reaction equation and the data provided in Table below, calculate the standard Gibbs free energy change ($\Delta G°_{rxn}$) in kJ at 25°C.

Refer Below - Table1 : Standard Enthalpies of formation of several compounds at 25°C. .

(8 marks)

Table 1: Standard Enthalpies of formation of several compounds at 25°C.

Compound	Formula	$\Delta H°_f$ (kJ mol^{-1})	S° (J mol^{-1} K^{-1})
Methane	$CH_4(g)$	−74.8	186.3
Oxygen	$O_2(g)$	0	205.0
Carbon dioxide	$CO_2(g)$	−393.5	213.7
Water	$H_2O(g)$	−241.8	188.8

Question 2

Show all the steps in the calculations clearly and state the units for all intermediate and final answers.

(a) Write a balanced equation for each of the following chemical reactions:

i. $NOCl(g) \leftrightarrow NO(g) + Cl_2(g)$

(1 marks)

ii. $Cu(NO_3)_2(s) \leftrightarrow CuO(s) + NO_2(g) + O_2(g)$

(1 marks)

iii. $CH_3COOC_2H_5(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + C_2H_5OH(aq)$

(1 marks)

iv. $Fe^{3+}(aq) + OH^-(aq) \leftrightarrow Fe(OH)_3(s)$

(1 marks)

(b) Write the expression for the equilibrium constant (K_c) for any three chemical reactions in Question 2(a).

(6 marks)

(c) Consider the following equilibrium at 527°C:

$$2Na(g) + Cl_2(g) \leftrightarrow 2NaCl(g)$$

At equilibrium, the concentrations for each gas are [Na] = 0.030 M, [Cl_2] = 0.025 M and [NaCl] = 0.090 M. Calculate the equilibrium constants, K_c and K_p.

(5 marks)

(d) Hydrogen bromide can be formed from hydrogen and bromine according to the reaction:

$$H_2(g) + Br_2(g) \leftrightarrow 2HBr(g)$$

A 0.75 dm^3 container is filled with 1.20 mol of H_2 and 1.80 mol Br_2 at 448 °C. The equilibrium constant (K_c) for the reaction at this temperature is 3.50. Construct the ICE table with sufficient boxes for all substances and represent the unknown values using the variable x.

(10 marks)

Question 3

Show all the steps in the calculations clearly and state the units for all intermediate and final answers.

(a) Determine the pH of the NaOH solution with a concentration of 3.0×10^{-2} M.

(4 marks)

(b) Determine whether each of the following hydronium ion (H_3O^+) concentrations is acidic, basic, or neutral.

i. $[H_3O^+] = 3.7 \text{ X } 10^{-3}$ M

(2 marks)

ii. $[H_3O^+] = 6.7 \times 10^{-8}$ M

(2 marks)

iii. $[H_3O^+] = 1.0 \text{ X } 10^{-7}$ M

(2 marks)

(c) A solution is prepared by dissolving 0.10 mol of sodium acetate (CH_3COONa) in 1.0 dm^3 of water, assuming complete dissociation of the salt. Given that the acid dissociation constant (K_a) of acetic acid (CH_3COOH) is 1.8×10^{-5} , calculate the pH of the resulting solution.

(7 marks)

(d) If 0.05 mol of HCl is added to the solution in question 4(c), calculate the new pH.

(8 marks)

Question 4

Show all calculation steps clearly and in an organized sequence, whenever applicable.

(a) Determine the oxidation number of the bolded elements in each of the following compounds:

i. $K_2\mathbf{Cr}_2O_7$

(1 marks)

ii. $H\mathbf{S}O_4^-$

(1 marks)

iii. $\mathbf{P}Cl_5$

(1 marks)

iv. $\mathbf{N}H_4^+$

(1 marks)

(b) Consider the reaction shown in the Table below to answer the following questions:
Refer Below - Table2 : Chemical reaction equation .

Table 2: Chemical reaction equation

Reaction	Chemical equation
1	$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$
2	$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
3	$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

i. classify each of the reactions as a redox or non-redox reaction.

(3 marks)

ii. identify the oxidizing and reducing agents for any two (2) of the reactions.

(3 marks)

(c) Consider the diagram of a galvanic cell, which consists of $Ag^+(aq)/Ag(s)$ and $Pb^{2+}(aq)/Pb(s)$ half-cells under standard conditions at 25 °C and a $KNO_3(aq)$ salt bridge between $AgNO_3$ and $Pb(NO_3)_2$ electrolyte solution.

i. Sketch a suitable galvanic cell diagram that utilizes the resulting standard cell potential. Include the label of the anode and cathode, the electrodes and solutions, the salt bridge and show the direction of electron movement.

(5 marks)

ii. Using the table of half-cell potentials in the Appendix (Table A1), calculate the Electromotive Force (e.m.f.) of the electrochemical cell.

(2 marks)

(d) The Nernst equation is given by:

$$E_{cell} = E^{\circ}_{cell} - (\frac{RT}{nF})InQ$$

At 298 K, the cell is operating under non-standard conditions with $[Ag^{+}]$ = 0.551 M and $[Pb^{2+}]$ = 0.897 M. Using the Nernst equation, show whether the cell in Question 3(c) is spontaneous or non-spontaneous.

(8 marks)

Question 5

Show all the steps in the calculations clearly and state the units for all intermediate and final answers.

(a) The reaction between nitrogen monoxide and hydrogen gas is represented as:

$$2NO(g) + 2H_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

Express the reaction rate in terms of changes in the concentration of NO and N_2.

(2 marks)

(b) The following data were obtained in the experiment from the chemical reaction in question 5(a). Use the data in Table below to answer the following questions.
Refer Below - Table3 : Experiment data of the reaction between NO(g) and H2(g). .

Table 3: Experiment data of the reaction between NO(g) and H2(g).

Experiment	**Initial [NO] (M)**	**Initial [H_2] (M)**	**Initial Rate (M/s)**
1	0.10	0.10	1.2×10^{-3}
2	0.20	0.10	4.8×10^{-3}
3	0.10	0.20	1.2×10^{-3}
4	0.30	0.10	**x**

i. Determine the reaction order of the overall reaction.

(4 marks)

ii. Determine the average rate constant (k) with respect to the initial rate of N_2 formation.

(2 marks)

iii. Determine the value of *x* for Experiment 4.

(2 marks)

(c) At 230°C the rate constant (k) of a chemical reaction of AB is $6.30 \times 10^{-4}\ s^{-1}$ and at 251°C, the rate constant is $3.16 \times 10^{-3} s^{-1}$. [Hint: R= 8.314 J/mol.K].

i. Calculate the value of the activation energy (E_a) of the reaction.

(5 marks)

ii. Calculate the frequency factor (A) of the reaction based on the answer in 5(c)(i).

(5 marks)

(d) Consider the following reaction:

$$2A \rightarrow C + D$$

The rate constant, k is given as 3.18×10^{-4} $M^{-1}s^{-1}$. If the concentration was experimentally measured at 0.235 M after 3 minutes, calculate:

i. the initial concentration, [A].

(2 marks)

ii. the half-life of A in minutes.

(3 marks)

END OF EXAMINATION PAPER

APPENDIX A1

Half-Reaction	Standard Electrode Potential, $E°$/V
$F_2 + 2e^- \rightarrow 2F^-$	2.866
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.776
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.72
$Au^+ + e^- \rightarrow Au$	1.692
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.52
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.72
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.35827
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.232
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.224
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.229
$Pt^{2+} + 2e^- \rightarrow Pt$	1.18
$Br_2 + 2e^- \rightarrow 2Br^-$	1.0873
$Hg^{2+} + 2e^- \rightarrow Hg$	0.851
$Ag^+ + e^- \rightarrow Ag$	0.7996
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.7973
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.771
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.695
$I_2 + 2e^- \rightarrow 2I^-$	0.5355
$Cu^{2+} + 2e^- \rightarrow Cu$	0.3419
$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$, 0.1 m HCl	0.3337
$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$, saturated KCl	0.2412
$AgCl(s) + e^- \rightarrow Ag + Cl^-$	0.22233
$Cu^{2+} + e^- \rightarrow Cu^+$	0.153
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	0.151
$HgO + H_2O + 2e^- \rightarrow Hg + 2OH^-$	0.0977
$2H^+ + 2e^- \rightarrow H_2$	0.00 (by definition)
$Pb^{2+} + 2e^- \rightarrow Pb$	−0.1262
$Sn^{2+} + 2e^- \rightarrow Sn$	−0.1375
$Ni^{2+} + 2e^- \rightarrow Ni$	−0.257
$PbCl_2 + 2e^- \rightarrow Pb + 2Cl^-$	−0.2675
$Co^{2+} + 2e^- \rightarrow Co$	−0.280
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.447
$Cr^{3+} + 3e^- \rightarrow Cr$	−0.744
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.7618
$Al^{3+} + 3e^- \rightarrow Al$	−1.662
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.372
$Na^+ + e^- \rightarrow Na$	−2.71
$Ca^{2+} + 2e^- \rightarrow Ca$	−2.868
$K^+ + e^- \rightarrow K$	−2.931
$Li^+ + e^- \rightarrow Li$	−3.0401

APPENDIX A2

Mass

1 lbm = 0.453592 kg

1 ton = 2000 lbm

1 kg = 2.20462 lbm

Pressure

1 atm = 1 bar = 760 mmHg = 760 torr

1 atm = 101.325 kPa

Volume

1 ft^3 = 0.028317 m^3

1 L = 0.001 m^3

1 m^3 = 35.32 ft^3

1 cm^3 = 0.06102 in^3

1 dm^3 = 1 Liter

1 gal =0.0037854 m^3

1 gal/min = 6.31 x 10^{-5} m^3/s

Faraday Constant = 96 485 Coulombs/mol

Gas Constant, R = 0.0821 L.atm/K.mol

= 8.314 J/K.mol

= 8.314 Pa.m^3 /K.mol

$C_{p,\,water}$ = 4.184 J/g°C.

Standard enthalpy of reaction, $\Delta H^o_{rxn} = \Sigma n\Delta H^o_f$ (products) - $\Sigma m\Delta H^o_f$ (reactants)

Standard entropy of reaction, $\Delta S°_{rxn}=\Sigma n\Delta S°_{(products)}-\Sigma n\Delta S°_{(reactants)}$

Gibbs free energy, $\Delta G = \Delta H - T\Delta S$

$$pH = pK_a - \log\frac{[acid]}{[salt]}$$

$$pOH = pK_b - \log\frac{[base]}{[salt]}$$

Nerst equation, $E_{cell} = E°_{cell} - [\frac{RT}{nF}\ln Q]$

Activation energy, $\ln [k_1/k_2] = E_a/R(1/T_2 - 1/T_1)$

Arrhenius equation, $k_1 = Ae^{-E_a/RT_1}$

Integrated rate laws & half-life formula:

Zero order: $[A]t = -kt + [A]_0$, $t_{1/2} = [A]_0/2k$

1st order: $\ln[A]_t = -kt + \ln[A]_0$, $t_{1/2} = \ln 2/k$

2nd order: $\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$, $t_{1/2} = 1/k[A]_0$