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5. The following elementary steps have been proposed for a reaction:
$NO + ½ O_2 \rightarrow NO_2$
$NO_2 + H_2O + SO_2 \rightarrow H_2SO_4 + NO$
The catalyst in this process is

A NO
B O_2
C NO_2
D SO_2
E H_2SO_4

6. 0.16 mol of SO_2 and 0.12 mol of O_2 were introduced into a 1 dm^3 vessel at constant temperature. When the system reached equilibrium, 0.06 mol of SO_3 was present.
The reaction is: $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$.
Which set of values shows the concentration of each gas at equilibrium?

	$[SO_2]$/moldm^{-3}	$[O_2]$/moldm^{-3}	$[SO_3]$/moldm^{-3}
A	0.16	0.12	0.06
B	0.10	0.09	0.06
C	0.16	0.09	0.06
D	0.16	0.12	0.10
E	0.10	0.06	0.10

7. At a given temperature, T, some PCl_5 at an initial concentration of 1.0 M, was placed in a container and allowed to dissociate into PCl_3 and Cl_2. It was found that the PCl_5 was 40 % dissociated at equilibrium.
K_c for the process: $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ at temperature, T, is closest to

A 0.27
B 0.40
C 0.60
D 3.7
E 4.0

8. At 298 K, K_c for the process $CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s)$ is 1.29×10^{24}.
Which of the following **cannot** be deduced from the data?

A $K_c = 1/[CO_2]$
B Equilibrium position lies far to the right.
C When calcium oxide and carbon dioxide react, the limiting reagent is almost completely used up.
D The rate of the reaction between calcium oxide and carbon dioxide to form calcium carbonate is extremely fast.
E K_p for the reverse process is pCO_2.

9. In an equilibrium system, a catalyst increases

A the activation energy of the forward process whilst decreasing that of the reverse process.
B the rate of forward process whilst decreasing that of the reverse process.
C the activation energy of both the forward and reverse processes.
D the enthalpy change for the reverse process.
E the rates of both the forward and reverse processes.

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10. If the system : NH_4CONH_2 (s) ⇌ 2 NH_3 (g) + CO (g) is at equilibrium at constant temperature, and some of the ammonium carbamate (NH_4CONH_2) is removed, then

A The masses of ammonia and carbon dioxide increase until equilibrium is re-established.
B The masses of ammonia and carbon dioxide decrease until equilibrium is re-established.
C The mass of ammonium carbamate increases to re-establish equilibrium.
D The partial pressures of ammonia and carbon dioxide decrease.
E The partial pressures of ammonia and carbon dioxide remain unchanged.

11. Consider the process: P_4 (g) + 6 H_2 (g) ⇌ 4 PH_3 (g) $\Delta H = +110.5$ kJ at equilibrium.
The value of K_c can be increased by

A Using a suitable catalyst.
B Adding some H_2 to the equilibrium mixture.
C Increasing the concentration of PH_3.
D Increasing the temperature.
E Decreasing the volume of the container.

12. Ammonium hydrogen sulphide dissociates into ammonia and hydrogen sulphide:
NH_4HS (s) ⇌ NH_3 (g) + H_2S (g)
When ammonium hydrogen sulphide is introduced into a closed vessel at 282.5 K, the total pressure at equilibrium is 0.230 atm. K_p for the system at 282.5 K is closest to

A 0.230
B 0.115
C $(0.230)^2$
D $(0.115)^2$
E 2 x 0.230

13. For the reaction $PCl_5(g)$ ⇌ $PCl_3(g) + Cl_2(g)$, $K_p = 1.7$ at 298K. Five systems were set up with the initial partial pressure of each gas as shown in the table. In which system would the **reverse** reaction occur to establish equilibrium?

	Initial partial pressure /atm		
	PCl_5	PCl_3	Cl_2
A	1	1	1
B	2	2	2
C	1	0.5	1.5
D	2	2	1
E	3	1	2

14. According to the Bronsted-Lowry definition, an **acid** is a substance which donates a

A hydrogen atom.
B hydrogen ion.
C hydrogen molecule.
D hydride ion.
E hydroxide ion.

15. Which does **not** constitute an acid/base conjugate pair?

A H_2SO_4/ HSO_4^-
B $H_2PO_4^-$/HPO_4^{2-}
C NH_4^+/ NH_3
D HNO_2/ NO_2^-
E H_3O^+/ OH^-

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16. Which set shows the substances in order of **increasing** acid strength?

A $HClO$, $HClO_2$, $HClO_3$
B H_2SO_4, H_2SO_3, HSO_4^-
C HCl, HBr, HF
D HF, H_2O, NH_3
E H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}

17. The acidity constant for an acid, HA, is 3.2×10^{-5}. The pK_b of its conjugate base is

A 4.5
B 9.5
C 3.1×10^{-10}
D 1.0×10^{-14}
E 14

18. Which salt would be expected to produce a solution with the **lowest** pH? Assume all solutions have the same molar concentration.

A NaCl
B $MgCl_2$
C $CrCl_3$
D $CaCl_2$
E $BaCl_2$

Questions 19 to 23 refer to the following solutions.

A $1x10^{-4}$ M HI
B $1x10^{-4}$ M KOCl
C $1x10^{-4}$ M $Fe(NO_3)_3$
D $1x10^{-4}$ M KCl
E $1x10^{-4}$ M KOH

Select, from A to E,

19. the solution which would have the lowest pH.
20. the solution which would have the highest pH.
21. the solution which would have a pH closest to 7.
22. the solution which would have a pH between 4 and 7.
23. the solution which would have a pH between 7 and 10.

Questions 24 – 28 refer to the following titrations:

A The titration of 20.0 cm^3 of 0.1M HCl with 0.1 M NaOH
B The titration of 20.0 cm^3 of 0.1M HCl with 0.1 M NH_3
C The titration of 20.0 cm^3 of 0.1M CH_3COOH with 0.1 M NaOH
D The titration of 20.0 cm^3 of 0.1M KOH with 0.1 M HCl
E The titration of 20.0 cm^3 of 0.1M HNO_3 with 0.1 M KOH

For which titration

24. would there be a decreases in pH as the titrant is added?
25. would the pH be greater than 7 at the equivalence point?
26. would the pH be lower than 7 at the equivalence point?
27. would phenolphthalein (pH range 8.3 – 10.0) be unsuitable as an indicator?
28. would bromocresol green (pH range 3.8 – 5.4) be unsuitable as an indicator?

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29. The solubility of silver phosphate(Ag_3PO_4) is x moldm^{-3}. The solubility product of this compound is

A x
B $4x^2$
C $4x^3$
D $27x^4$
E $3x^2$

Questions 30 and 31 require the following information:

The solubility product of $PbCl_2$ is 1.7×10^{-5}.

30. The solubility of $PbCl_2$ in water is closest to

A 1.6×10^{-2} mol dm^{-3}
B 4.1×10^{-3} mol dm^{-3}
C 4.3×10^{-6} mol dm^{-3}
D 2.9×10^{-10} mol dm^{-3}
E 1.2×10^{-9} mol dm^{-3}

31. The solubility of $PbCl_2$ in 0.10M NaCl is closest to

A 1.7×10^{-3} mol dm^{-3}
B 1.6×10^{-4} mol dm^{-3}
C 1.7×10^{-7} mol dm^{-3}
D 2.9×10^{-1} mol dm^{-3}
E 1.2×10^{-5} mol dm^{-3}

Questions 32 –33 concern the following compounds:

A NH_4NO_3
B NaH
C H_2O_2
D $KHCO_3$
E HBr

32. In which compound does hydrogen carry an oxidation number of –1?

33. In which compound does oxygen carry an oxidation number of –1?

Questions 34-35 concern the following reactions:

A $NH_4Cl \rightarrow NH_3 + HCl$
B $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$
C $5\ HClO_2 \rightarrow 4\ ClO_2 + HCl + 2\ H_2O$
D $2\ KMnO_4 + 5\ SO_2 + 2\ H_2O \rightarrow 2\ MnSO_4 + K_2SO_4 + 2\ H_2SO_4$
E $S_2O_8^{2-} + 2\ I^- \rightarrow 2\ SO_4^{2-} + I_2$

34. Which is **not** a redox reaction?

35. Which is a disproportionation reaction?

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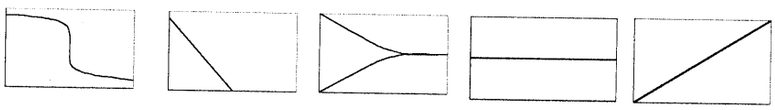
Questions 36 - 40 concern the following graphs:

A B C D E



Select, from A to E, the graph which best represents:

36. Concentration of X versus time for a reaction which is zero order in X.
37. Rate of reaction versus concentration of X for a reaction which is zero order in X.
38. Rate of reaction versus concentration of X for a reaction which is first order in X.
39. Rate of reaction versus time for a reversible process which attains equilibrium after some time.
40. The titration curve for the titration of a base with an acid.

SECTION B: Answer **all** questions in the spaces provided in the question paper.

You may use the following information wherever necessary:
$R = 8.31\ J\ mol^{-1}\ K^{-1} = 0.0821\ atm\ dm^3\ mol^{-1}\ K^{-1} = 0.0821\ atm\ L\ mol^{-1}\ K^{-1}$
$pK_w = 14$ at 298 K
$pK_a(HClO_2) = 1.96$
$pK_a(CH_3COOH) = 4.76$
$K_p = K_c (0.0821T)^{\Delta n(gas)}$

Remember to include units in your answers wherever appropriate.

1. The rate law for the reaction: $X + 2\,Y \rightarrow 4\,Z$ is
Rate = $k\,[X]^2\,[Y]$ with a rate constant of $1.2 \times 10^{-4}\ M^{-2}\ s^{-1}$ at 298 K.

a) Find the rate of the reaction the instant $20.0 cm^3$ of a 0.20 M solution of X is mixed with $30.0 cm^3$ of a 0.25 M solution of Y at 298 K. [2 marks]

b) What is the value of the ***rate constant*** if the concentration of X is doubled and the concentration of Y is kept constant at 298 K? [1mark]

c) The Arrenhius equation states that $k = Ae^{-Ea/RT}$.
What is represented by $e^{-Ea/RT}$? [1 mark]

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2. At 600 ^{0}C, the equilibrium constant, K_p, is 0.20 for the reaction :
$CO\ (g) + Cl_2\ (g) \rightleftharpoons COCl_2(g)$.

a) Find the equilibrium partial pressure of each gas when 0.15 mol each of CO and Cl_2 are admitted into a 1.50 dm^3 vessel at 298 K and the system reaches equilibrium. [5 marks]

b) Find the total pressure of the system at equilibrium. [1 mark]

c) What effect, if any, will the addition of 0.01mol of an inert gas have on the equilibrium position if the volume is kept constant? **Show your reasoning**. [2 marks]

d) What effect, if any, will the addition of 0.01mol of an inert gas have on the equilibrium position if the total pressure is kept constant? **Show your reasoning**. [2 marks]

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3. Find the pH of the following solutions:

a) 0.20 M $NaClO_2$ [4 marks]

b) a mixture of 20.0 cm^3 of 0.20 M $HClO_2$ + 30.0 cm^3 of 0.20 M NaOH. [3 marks]

c) a mixture of 30.0 cm^3 of 0.20 M $HClO_2$ + 20.0 cm^3 of 0.20 M NaOH. [3 marks]

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4. Use the following table of standard reduction potentials wherever necessary.

	E^0/V
MnO_4^- (aq) + 8 H^+ (aq) + 5 e^- → Mn^{2+} (aq) + 4 H_2O (l)	+1.51
Cl_2 (g)+ 2 e^- → 2Cl^- (aq)	+1.36
$Cr_2O_7^{2-}$ (aq) + 14 H^+ (aq) + 6 e^- → 2 Cr^{3+} (aq) + 7 H_2O (l)	+1.33
Ag^+ (aq) + e^- → Ag (s)	+0.80
Fe^{3+} (aq) + e^- → Fe^{2+} (aq)	+0.77
2 H^+ (aq) + 2e^- → H_2 (g)	0.00
Fe^{3+} (aq) + 3e^- → Fe (s)	-0.036
Fe^{2+} (aq) + 2e^- → Fe (s)	-0.44
Mg^{2+} (aq) + 2e^- → Mg (s)	-2.38

a) The cell notation represents a **standard** galvanic cell:
Mg (s) | $MgCl_2$ (aq) || $FeCl_3$ (aq), $FeCl_2$ (aq)|Pt (s)

i) Write a balanced **ionic** equation for the cell reaction.[1 mark]

ii) What is the cell potential? [1 mark]

iii) Draw a **fully labeled** diagram of the galvanic cell. Show the direction of flow of electrons, the polarity of the electrodes and the concentration of all solutions. [5 marks]

b) Use the table of standard reduction potentials to explain why the reaction of chlorine gas on iron metal produces iron(11) chloride instead of iron (111) chloride. [2 marks]

c) Given the Nernst Equation: $E = E^0 - \frac{0.059}{n}\log Q$, at 298 K, find the value of the equilibrium constant for the reaction:
5 Fe^{2+} (aq) + MnO_4^- (aq) + 8 H^+ (aq) ⇌ Mn^{2+} (aq) + 4 H_2O (l) + 5 Fe^{3+} (aq)
at 298 K. [3 marks]

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5. Derive a balanced **ionic** equation for each reaction by writing half equations and then combining them.

a) $PbS\ (s) + NO_3^-(aq) \rightarrow S\ (s) + Pb^{2+}\ (aq) + NO\ (g)$ (in **acid** medium) [3 marks]

b) $MnO_4^-\ (aq) + IO_3^-\ (aq) \rightarrow MnO_2\ (s) + IO_4^-\ (aq)$ (in **basic** medium) [3 marks]

6. Use the given K_c values for the processes X, Y and Z to find K_c for the process Q.

Process X: $N_2O_4(g) \rightleftharpoons 2\ NO_2\ (g)$	$K_c = K_x = 4.6 \times 10^{-3}$
Process Y: $N_2\ (g) + 2\ O_2\ (g) \rightleftharpoons 2\ NO_2\ (g)$	$K_c = K_y = 1.7 \times 10^{-17}$
Process Z: $2\ N_2O\ (g) + 3\ O_2\ (g) \rightleftharpoons 2\ N_2O_4\ (g)$	$K_c = K_z = 1.2 \times 10^{6}$
Process Q: $N_2\ (g) + ½\ O_2\ (g) \rightleftharpoons N_2O\ (g)$	

[2 marks]

END OF EXAMINATION