S5 Chemistry Practice Paper (Set 3): Rates of Reaction and Collision Theory, Chemical Equilibrium and the Equilibrium Constant, Weak Acid pH and Titration Curves, Electrolysis and Faraday Calculations, Enthalpy Change and Hess Law, Redox Titration

S5 Chemistry | 100 marks | 75 minutes

Question 1: Rates of Reaction and Collision Theory

Excess marble chips (calcium carbonate) are added to 50.0 cm3 of 1.0 M hydrochloric acid at 25 degrees Celsius. The volume of carbon dioxide collected is 0 cm3 at 0 s, 48 cm3 at 20 s, 72 cm3 at 40 s, 86 cm3 at 60 s, 94 cm3 at 80 s, 98 cm3 at 100 s and 98 cm3 at 120 s.

Calculate the average rate of reaction over the first 40 seconds in cm3 per second. *(5 marks)*

Answer: Average rate = (72 - 0) / 40 = 1.8 cm3 per second.

Compare the average rate in the first 20 s with that between 20 s and 40 s, and explain the difference. *(4 marks)*

Answer: First 20 s = 48 / 20 = 2.4 cm3/s; 20-40 s = (72 - 48) / 20 = 1.2 cm3/s. The later stage is slower because the acid concentration falls as it is used up, so collisions become less frequent.

If the same mass of marble is used as a powder, how does the curve change? Does the final gas volume change? *(4 marks)*

Answer: Powder has a larger surface area, so collisions are more frequent, the reaction is faster, the curve is steeper and reaches the plateau earlier. The amount of calcium carbonate is the same, so the final volume is still 98 cm3, unchanged.

Suggest two ways to speed up this reaction without changing the final gas volume, and explain why. *(3 marks)*

Answer: (1) Raise the temperature: particles have more kinetic energy so more collisions exceed the activation energy. (2) Add a catalyst: it provides an alternative pathway of lower activation energy. Neither changes the amount of reactants.

Question 2: Chemical Equilibrium and the Equilibrium Constant

Nitrogen and hydrogen are sealed in a container at 500 degrees Celsius: N2(g) + 3H2(g) <-> 2NH3(g), delta H = -92 kJ per mole. At equilibrium [N2] = 0.50 M, [H2] = 0.30 M and [NH3] = 0.40 M.

Write the expression for the equilibrium constant Kc and calculate its value and unit. *(5 marks)*

Answer: Kc = [NH3]^2 / ([N2][H2]^3) = 0.40^2 / (0.50 x 0.30^3) = 11.9, with units mol^-2 dm^6.

At the same temperature the container volume is halved. How do the position of equilibrium and Kc change? *(4 marks)*

Answer: Halving the volume doubles all concentrations, so the equilibrium shifts to the side with fewer gas moles (the right, delta n = -2) and the ammonia yield increases. Kc depends only on temperature, so it is unchanged.

The temperature is raised to 600 degrees Celsius. How do Kc and the equilibrium yield of ammonia change? Explain. *(4 marks)*

Answer: The forward reaction is exothermic (delta H = -92 kJ per mole), so raising the temperature shifts the equilibrium to the left (endothermic direction), lowering the ammonia yield and decreasing Kc.

What effect does adding an iron catalyst have on Kc and on the position of equilibrium? *(4 marks)*

Answer: A catalyst speeds up both the forward and reverse reactions, so equilibrium is reached sooner, but it does not change Kc or the position of equilibrium.

Question 3: Weak Acid pH and Titration Curves

At 25 degrees Celsius, 25.0 cm3 of 0.100 M ethanoic acid (CH3COOH, Ka = 1.8 x 10^-5 mol dm^-3) is titrated with 0.100 M sodium hydroxide solution.

Calculate the pH of the acid before any alkali is added. *(4 marks)*

Answer: [H+] = sqrt(Ka x c) = sqrt(1.8x10^-5 x 0.100) = 1.34e-03 M, so pH = 2.87.

Calculate the volume of sodium hydroxide needed to reach the equivalence point. *(3 marks)*

Answer: The mole ratio is 1:1, so V = (0.100 x 25.0) / 0.100 = 25.0 cm3.

Calculate the pH at the equivalence point (salt concentration 0.050 M). *(5 marks)*

Answer: Ethanoate hydrolyses: Kb = Kw / Ka = 5.56x10^-10, [OH-] = sqrt(Kb x 0.050) = 5.27e-06 M, pOH = 5.28, pH = 8.72.

Which indicator should be used? Explain. *(5 marks)*

Answer: Use phenolphthalein (pH 8.3-10.0) because the equivalence point pH is about 8.72, which is weakly alkaline and lies inside its range; methyl orange (pH 3.1-4.4) would change colour too early.

Question 4: Electrolysis and Faraday Calculations

A steady current of 2.0 A is passed through copper(II) sulphate solution for 30 minutes (1 F = 96500 C per mole, Cu = 63.5, molar volume 24.0 dm3 per mole at room temperature).

Calculate the quantity of charge passed and the number of moles of electrons. *(5 marks)*

Answer: Q = It = 2.0 x 1800 = 3600 C; moles of electrons = 3600 / 96500 = 0.0373 mol.

With copper as the cathode, calculate the mass gained at the cathode. *(5 marks)*

Answer: Cu2+ + 2e- -> Cu, so moles of copper = 0.0373 / 2 = 0.0187 mol and mass = 0.0187 x 63.5 = 1.18 g.

If inert electrodes are used to electrolyse dilute sulphuric acid, calculate the volume of oxygen released at the anode in cm3. *(4 marks)*

Answer: Anode: 2H2O -> O2 + 4H+ + 4e-, moles of oxygen = 0.0373 / 4 = 0.0093 mol, volume = 0.0093 x 24000 = 224 cm3 (hydrogen at the cathode is 448 cm3).

To electroplate 0.200 mol of silver at 5.0 A, how many seconds are needed? *(3 marks)*

Answer: Ag+ + e- -> Ag requires 0.200 mol of electrons, charge = 0.200 x 96500 = 19300 C, so t = 19300 / 5.0 = 3860 s.

Question 5: Enthalpy Change and Hess Law

Standard enthalpies of formation: CO2(g) -394, H2O(l) -286, C2H5OH(l) -277 and CO(g) -111 kJ per mole. Specific heat capacity of water is 4.2 J per gram per kelvin and the molar mass of ethanol is 46.0 g per mole.

Calculate the standard enthalpy of combustion of ethanol. *(5 marks)*

Answer: C2H5OH + 3O2 -> 2CO2 + 3H2O, so delta Hc = 2(-394) + 3(-286) - (-277) = -1369 kJ per mole.

How much energy is released by burning 4.6 g of ethanol? *(5 marks)*

Answer: Moles = 4.6 / 46.0 = 0.100 mol, so energy = 0.100 x 1369 = 136.9 kJ.

If all this energy heats water from 25 to 65 degrees Celsius, what mass of water can be heated? *(4 marks)*

Answer: delta T = 40 K, so m = 136900 / (4.2 x 40) = 815 g.

Use Hess law to find delta H for CO(g) + H2O(l) -> CO2(g) + H2(g), and state why the experimental value is often less exothermic than the theoretical one. *(3 marks)*

Answer: delta H = (-394 + 0) - (-111 - 286) = +3 kJ per mole (endothermic). Experimental values are smaller because of heat loss, incomplete combustion and energy used to vaporise water or ethanol.

Question 6: Redox Titration

25.0 cm3 of iron(II) sulphate solution is titrated with 0.0200 M acidified potassium manganate(VII) solution, mean titre 24.0 cm3. The same iron(II) solution is then used to titrate 25.0 cm3 of 0.0167 M potassium dichromate solution.

Write the balanced ionic equation and find the concentration of the iron(II) solution. *(5 marks)*

Answer: MnO4- + 8H+ + 5Fe2+ -> Mn2+ + 4H2O + 5Fe3+; MnO4- = 0.0200 x 0.0240 = 0.0005 mol, Fe2+ = 5 x 0.0005 = 0.0024 mol, c = 0.0024 / 0.0250 = 0.0960 M.

State the changes in oxidation number of manganese and iron. *(3 marks)*

Answer: Manganese falls from +7 to +2 (reduction, gaining 5 electrons); iron rises from +2 to +3 (oxidation, losing 1 electron each).

Why is no extra indicator needed in this titration? *(4 marks)*

Answer: Potassium manganate(VII) is deep purple and is reduced to almost colourless Mn2+ during the titration; at the end point one excess drop gives a persistent pale pink colour, so it is self-indicating.

Calculate the volume of iron(II) solution needed to titrate 25.0 cm3 of the dichromate solution. *(4 marks)*

Answer: Cr2O7^2- + 14H+ + 6Fe2+ -> 2Cr3+ + 7H2O + 6Fe3+; Cr2O7^2- = 0.0004 mol, Fe2+ = 6 x 0.0004 = 0.0025 mol, V = 0.0025 / 0.0960 x 1000 = 26.1 cm3.