S4 Chemistry Practice Paper: moles and formulae, molar volume, solution concentration and dilution, solubility and ppm, pH and neutralisation, titration, redox and electrolysis, calorimetry, Hess law and rates of reaction (100 marks)

Level: S4 | Subject: Chemistry | 75 minutes | 100 marks | 24 questions | answer key

Section A: Moles and Formulae (18 marks)

Answer all questions and show your working, units and significant figures.

#	Question	Answer	Marks
1	Copper(II) sulphate crystals, CuSO4.5H2O, are used as a fungicide. (Cu = 63.5, S = 32.1, O = 16.0, H = 1.0) (a) Calculate its relative molecular mass. (b) Find the percentage by mass of water in the crystal (1 d.p.). (c) If 24.96 g of the crystals are heated until all the water is lost, what mass remains?	(a) Mr = 63.5 + 32.1 + 64.0 + 90.0 = 249.6. (b) Water by mass = 90.0 / 249.6 x 100% = 36.1%. (c) 24.96 g is 0.100 mol, so 0.500 mol of water (9.00 g) is lost and 24.96 - 9.00 = 15.96 g of anhydrous salt remains.	5
2	12.5 g of calcium carbonate (CaCO3, Mr = 100.1) is heated until it decomposes completely: CaCO3 -> CaO + CO2. (a) Find the number of moles of calcium carbonate (4 d.p.). (b) Find the mass of carbon dioxide released (2 d.p.). (c) What volume does this carbon dioxide occupy at room conditions (molar volume 24.0 dm3/mol)?	(a) n = 12.5 / 100.1 = 0.1249 mol. (b) m(CO2) = 0.1249 x 44.0 = 5.49 g. (c) V = 0.1249 x 24.0 = 3.00 dm3.	5
3	An organic compound contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen by mass. Its relative molecular mass is 180. (a) Find its empirical formula. (b) Find its molecular formula. (c) How many moles of oxygen atoms are in 12.0 g of the compound?	(a) Mole ratio C : H : O = 3.33 : 6.70 : 3.33 = 1 : 2.01 : 1, so the empirical formula is CH2O (formula mass 30.0). (b) 180 / 30.0 = 6, so the molecular formula is C6H12O6. (c) 12.0 g = 0.0667 mol, each molecule has 6 oxygen atoms, giving 0.400 mol of oxygen atoms.	4
4	4.4 g of carbon dioxide (Mr = 44.0) is kept at room conditions (molar volume 24.0 dm3/mol). (a) Find the number of moles. (b) Find its volume in dm3. (c) Find the number of molecules (Avogadro constant 6.02 x 10^23).	(a) n = 4.4 / 44.0 = 0.10 mol. (b) V = 0.100 x 24.0 = 2.4 dm3. (c) Molecules = 0.100 x 6.02 x 10^23 = 6.02 x 10^22.	4

Section B: Solutions and Concentration (17 marks)

Answer all questions and show your working, units and significant figures.

#	Question	Answer	Marks
5	5.85 g of sodium chloride (Mr = 58.5) is dissolved in water to make 250.0 cm3 of solution. (a) Find the number of moles. (b) Find the molarity in mol/dm3. (c) 25.0 cm3 of this solution is diluted to 250.0 cm3; find the new concentration.	(a) n = 5.85 / 58.5 = 0.100 mol. (b) c = 0.100 / 0.250 = 0.400 mol/dm3. (c) Moles taken = 0.400 x 0.0250 = 0.0100 mol, new concentration = 0.0100 / 0.250 = 0.040 mol/dm3 (a tenfold dilution).	5
6	A laboratory has 2.0 mol/dm3 sulphuric acid. (a) What volume of the concentrated acid is needed to prepare 500.0 cm3 of 0.20 mol/dm3	(a) c1V1 = c2V2: 2.0 x V1 = 0.20 x 500.0, so V1 = 50.0 cm3. (b) Water added = 500.0 - 50.0 = 450.0 cm3. (c) Concentrated sulphuric	4

	dilute acid? (b) How much water must be added? (c) Explain why the acid is added to water and not water to the acid.	acid releases much heat on dilution; adding water to the acid can cause local boiling and spitting, so the acid is added slowly to water with stirring.	
7	At 20 C the solubility of potassium nitrate is 36.0 g per 100 g of water. (a) Find the percentage by mass of the saturated solution (1 d.p.). (b) If 136.0 g of this saturated solution is cooled to 0 C (solubility 13.0 g per 100 g of water), what mass of crystals separates?	(a) Percentage by mass = 36.0 / 136.0 x 100% = 26.5%. (b) The solution holds 36.0 g of solute in 100 g of water; at 0 C only 13.0 g dissolves, so 36.0 - 13.0 = 23.0 g crystallises out.	4
8	A 250.0 cm3 drinking water sample contains 5.0 mg of lead(II) ions (assume a density of 1.00 g/cm3). (a) Express the concentration in ppm (mg/kg). (b) If the limit for lead in drinking water is 10 ppb, does the sample exceed it? (c) Suggest one practical way to reduce the risk of lead in drinking water.	(a) Mass = 250.0 g = 0.250 kg, concentration = 5.0 mg / 0.250 kg = 20 ppm. (b) 20 ppm is 20 000 ppb, far above the 10 ppb limit, so it exceeds the standard. (c) Replace lead pipes or linings and flush the tap before use.	4

Section C: Acids, Bases and Titration (17 marks)

Answer all questions and show your working, units and significant figures.

#	Question	Answer	Marks
9	(a) Find the pH of 0.010 mol/dm3 hydrochloric acid (a strong acid, fully ionised). (b) Find the pH of 0.010 mol/dm3 sodium hydroxide solution. (c) When equal volumes of the two are mixed, what is the approximate pH? Explain.	(a) [H+] = 0.010, pH = -log 0.010 = 2.00. (b) [OH-] = 0.010, pOH = 2.00, pH = 14.00 - 2.00 = 12.00. (c) Equal moles neutralise to give NaCl and water, so the pH is about 7 (neutral) because neither Na+ nor Cl- hydrolyses.	5
10	25.0 cm3 of 0.20 mol/dm3 sodium hydroxide solution exactly neutralises 20.0 cm3 of dilute sulphuric acid: 2NaOH + H2SO4 -> Na2SO4 + 2H2O. (a) Find the moles of NaOH used. (b) Find the moles of H2SO4. (c) Find the concentration of the acid.	(a) n(NaOH) = 0.20 x 0.0250 = 0.0050 mol. (b) The ratio is 2 : 1, so n(H2SO4) = 0.00250 mol. (c) c = 0.00250 / 0.0200 = 0.125 mol/dm3.	4
11	25.0 cm3 of sodium carbonate solution is titrated with 0.100 mol/dm3 hydrochloric acid; the mean titre is 24.10 cm3: Na2CO3 + 2HCl -> 2NaCl + H2O + CO2. (a) Find the moles of HCl. (b) Find the moles and molarity of Na2CO3. (c) Find the mass of Na2CO3 in 25.0 cm3 of the solution (Mr = 106.0).	(a) n(HCl) = 0.100 x 0.02410 = 0.00241 mol. (b) Ratio 1 : 2, so n(Na2CO3) = 0.001205 mol and c = 0.001205 / 0.0250 = 0.0482 mol/dm3. (c) m = 0.001205 x 106.0 = 0.1277 g.	5
12	(a) Which indicator suits a titration of a strong acid with a weak base? (b) Which indicator suits a titration of a weak acid with a strong base? (c) Explain why a sodium carbonate solution has a pH greater than 7.	(a) Methyl orange (range pH 3.1-4.4, acidic equivalence point). (b) Phenolphthalein (range pH 8.3-10.0, basic equivalence point). (c) The carbonate ion hydrolyses: CO3^2- + H2O <-> HCO3- + OH-, producing hydroxide ions, so the solution is alkaline.	3

Section D: Redox, Metals and Electrolysis (16 marks)

Answer all questions and show your working, units and significant figures.

#	Question	Answer	Marks

13	A zinc strip is placed in blue copper(II) sulphate solution. (a) Write the ionic equation. (b) State which species is oxidised and which is reduced. (c) Describe two observable changes. (d) Explain how the reactivity of zinc relates to this reaction.	(a) Zn(s) + Cu2+(aq) -> Zn2+(aq) + Cu(s). (b) Zinc is oxidised (oxidation number 0 -> +2) and copper(II) ions are reduced (+2 -> 0). (c) The blue colour fades; a red-brown deposit forms on the zinc. (d) Zinc is more reactive than copper, and a more reactive metal can reduce the ions of a less reactive metal, so the reaction is spontaneous.	5
14	A current of 2.0 A is passed through copper(II) sulphate solution for 30 minutes and copper is deposited at the cathode (Cu = 63.5, Faraday constant 96500 C/mol). (a) Find the quantity of charge passed. (b) Find the moles of electrons (4 d.p.). (c) Find the mass of copper deposited (2 d.p.). (d) Write the cathode equation.	(a) Q = It = 2.0 x 1800 = 3600 C. (b) n(e-) = 3600 / 96500 = 0.0373 mol. (c) Cu2+ + 2e- -> Cu, so n(Cu) = 0.01865 mol and m = 0.01865 x 63.5 = 1.18 g. (d) Cathode: Cu2+(aq) + 2e- -> Cu(s).	5
15	Find the oxidation number of the stated element: (a) manganese in KMnO4; (b) chromium in K2Cr2O7; (c) oxygen in H2O2. State which of these compounds is commonly used as an oxidising agent.	(a) (+1) + x + 4(-2) = 0, so x = +7. (b) 2(+1) + 2x + 7(-2) = 0, so x = +6. (c) 2(+1) + 2x = 0, so x = -1. (d) Both KMnO4 and K2Cr2O7 are strong oxidising agents, and KMnO4 is the one most commonly used.	3
16	Iron gates rust easily in damp air. (a) Write the main oxidation reaction in rusting. (b) Explain why galvanising with zinc (sacrificial protection) prevents rusting better than tin plating. (c) Give one everyday method of rust prevention.	(a) Fe -> Fe2+ + 2e- (further oxidised to Fe3+ giving hydrated iron(III) oxide rust). (b) Zinc is more reactive than iron, so it is oxidised first and pushes electrons to the iron, which becomes the protected cathode; this still works when the coating is scratched. Tin is less reactive than iron, so once the coating breaks the iron becomes the anode and corrodes faster. (c) Paint or grease the surface, or keep it dry.	3

Section E: Energetics and Calorimetry (16 marks)

Answer all questions and show your working, units and significant figures.

#	Question	Answer	Marks
17	50.0 cm3 of 1.00 mol/dm3 hydrochloric acid is mixed with 50.0 cm3 of 1.00 mol/dm3 sodium hydroxide and the temperature rises by 6.7 C. Assume the solution density is 1.00 g/cm3 and the specific heat capacity is 4.18 J g-1 K-1. (a) Find the heat released in J. (b) Find the moles of water formed. (c) Find the enthalpy change of neutralisation in kJ/mol. (d) Explain why the experimental value is usually smaller in magnitude than the literature value of -57.3 kJ/mol.	(a) E = mcDT = 100.0 x 4.18 x 6.7 = 2801 J. (b) Both reactants give 0.0500 mol, so 0.0500 mol of water forms. (c) DH = -2800.6 / 0.0500 / 1000 = -56.0 kJ/mol. (d) Some heat is lost to the surroundings and the calorimeter, so the measured temperature rise is too small.	4
18	Burning 0.500 g of a fuel raises the temperature of 200.0 g of water by 12.5 C (specific heat capacity 4.18 J g-1 K-1). (a) Find the heat absorbed by the water. (b) Find the heat of combustion of the fuel in kJ/g. (c) If the fuel contains hydrogen, why can the water formed make the measured energy release appear higher?	(a) E = 200.0 x 4.18 x 12.5 = 10450 J. (b) Heat per gram = 10450 / 0.500 = 20900 J/g = 20.9 kJ/g. (c) Water vapour from combustion condenses inside the calorimeter and releases latent heat, which the water also absorbs, so the calculated heat of combustion comes out too high.	4

19	A 50.0 g metal block at 96.0 C is placed in 100.0 g of water at 22.0 C and the final temperature is 26.0 C (specific heat capacity of water 4.18 J g-1 K-1). (a) Find the heat gained by the water. (b) Find the specific heat capacity of the metal (3 d.p.). (c) If the true value is 0.900 J g-1 K-1, explain the difference.	(a) E = 100.0 x 4.18 x (26.0 - 22.0) = 1672 J. (b) The metal cools by 96.0 - 26.0 = 70.0 K, so c = 1672 / (50.0 x 70.0) = 0.478 J g-1 K-1. (c) Heat is lost to the air and the container, so the measured temperature rise of the water is too small and the calculated value is too low.	4
20	Given: C(s) + O2(g) -> CO2(g), DH = -394 kJ/mol; CO(g) + 1/2 O2(g) -> CO2(g), DH = -283 kJ/mol. (a) Use Hess law to find DH for C(s) + 1/2 O2(g) -> CO(g). (b) State the basis of Hess law. (c) Why does incomplete combustion of carbon release less energy?	(a) Subtract the second equation from the first: DH = -394 - (-283) = -111 kJ/mol. (b) Hess law follows from conservation of energy (the first law of thermodynamics): the total enthalpy change depends only on the initial and final states, not the route. (c) Incomplete combustion gives CO instead of CO2; CO can still burn to release more heat, so the first step releases less energy.	4

Section F: Rate of Reaction and Integrated Calculation (16 marks)

Answer all questions and show your working, units and significant figures.

#	Question	Answer	Marks
21	A magnesium ribbon is added to 1.0 mol/dm3 hydrochloric acid and 48.0 cm3 of hydrogen is collected in the first 20 s. (a) Find the average rate over the first 20 s in cm3/s. (b) Explain why the rate falls with time. (c) Write the ionic equation.	(a) Rate = 48.0 / 20 = 2.4 cm3/s. (b) The acid concentration falls as it is used up and the surface area of the magnesium decreases, so collisions become less frequent and the rate drops. (c) Mg(s) + 2H+(aq) -> Mg2+(aq) + H2(g).	4
22	The same mass of magnesium is used as a powder and as a ribbon with acid of the same concentration. (a) Which reacts faster? (b) Explain. (c) Name two other factors that would speed up the reaction, with one sentence of explanation each.	(a) The powder reacts faster. (b) Powder has a much larger surface area, so more particles are in contact with the acid and the frequency of effective collisions is higher. (c) Raising the temperature increases the average kinetic energy so more collisions exceed the activation energy; increasing the acid concentration raises the number of H+ ions per unit volume and hence the collision frequency.	4
23	(a) Use collision theory to explain how a rise in temperature speeds up a reaction. (b) How does a catalyst speed up a reaction without changing its enthalpy change? (c) Why can a catalyst not make a reaction happen that would otherwise not occur?	(a) A rise in temperature increases the average kinetic energy, so a much larger fraction of particles has energy above the activation energy and there are more effective collisions. (b) A catalyst provides an alternative route of lower activation energy; both forward and reverse reactions speed up, and the initial and final states are unchanged, so DH is unchanged. (c) A catalyst only lowers the activation energy; it cannot change the energy difference between reactants and products, so it does nothing for a reaction that is not energetically feasible.	4
24	100.0 g of calcium carbonate is heated for some time and 78.0 g of solid remains (CaCO3 -> CaO + CO2, Mr: CaCO3 = 100.1, CO2 = 44.0). (a) Find the moles of carbon dioxide released. (b) Find the mass of calcium carbonate decomposed. (c) Find the percentage decomposed (1 d.p.). (d) What mass	(a) Mass lost = 100.0 - 78.0 = 22.0 g, so n(CO2) = 22.0 / 44.0 = 0.500 mol. (b) The ratio is 1 : 1, so the mass decomposed = 0.500 x 100.1 = 50.05 g. (c) Percentage decomposed = 50.05 / 100.0 x 100% = 50.0%. (d) Complete decomposition would lose 0.999 x 44.0 = 44.0 g.	4

	would be lost on complete decomposition?		