S3 Chemistry Practice Paper 1 with Answers

Chemistry · S3 · Data-response

S3 chemistry practice paper with answers on Mr, percentage composition, combustion, concentration, neutralisation and metals. Full marking scheme included.

Part 1: Relative molecular mass and percentage composition

Data A (simulated): relative atomic masses are Ca = 40, C = 12, O = 16, Al = 27, S = 32, N = 14 and H = 1. Study the data and answer all parts.

#	Question	Answer and marking notes
1	Calculate the relative molecular mass of calcium carbonate (CaCO3).	Answer: Mr(CaCO3) = 40 + 12 + 3 x 16 = 100.
2	What percentage by mass of calcium carbonate is calcium? (1 decimal place)	Answer: percentage of calcium = 40 / 100 x 100 = 40.0 per cent.
3	Calculate the relative molecular mass of aluminium sulphate (Al2(SO4)3) and the percentage by mass of sulphur in it (2 decimal places).	Answer: Mr = 2 x 27 + 3 x (32 + 4 x 16) = 342; percentage of sulphur = 3 x 32 / 342 x 100 = 28.07 per cent.
4	Compare ammonium nitrate (NH4NO3) with ammonium sulphate ((NH4)2SO4): (1) calculate the percentage by mass of nitrogen in each (1 decimal place); (2) state which one has the higher nitrogen content; (3) calculate the mass of nitrogen in 50 g of ammonium nitrate (1 decimal place).	Answer: (1) Mr(NH4NO3) = 28 + 4 + 48 = 80, so nitrogen = 28 / 80 x 100 = 35.0 per cent; Mr((NH4)2SO4) = 28 + 8 + 32 + 64 = 132, so nitrogen = 28 / 132 x 100 = 21.2 per cent; (2) ammonium nitrate has the higher nitrogen content; (3) mass of nitrogen = 50 x 35.0 / 100 = 17.5 g.

Part 2: Combustion and conservation of mass

Data B (simulated): 4.8 g of magnesium ribbon is heated in a crucible. The magnesium reacts completely with oxygen in the air to form magnesium oxide (2Mg + O2 → 2MgO). Relative atomic masses: Mg = 24, O = 16.

#	Question	Answer and marking notes
5	Calculate the number of moles in 4.8 g of magnesium.	Answer: moles = 4.8 / 24 = 0.20 mol.
6	Using the equation, calculate the mass of magnesium oxide formed after complete reaction.	Answer: the mole ratio of magnesium to magnesium oxide is 2:2 (that is 1:1), so magnesium oxide = 0.20 mol x 40 = 8.0 g.
7	Calculate the mass of oxygen used up in the reaction.	Answer: mass of oxygen = 8.0 - 4.8 = 3.2 g.
8	(1) If the experiment actually collects only 6.4 g of magnesium oxide, calculate the percentage yield (1 decimal place). (2) Calculate the mass of magnesium oxide lost and its percentage of the theoretical yield (1 decimal place). (3) Explain why less product is collected than the theoretical mass when magnesium burns in an open crucible.	Answer: (1) yield = 6.4 / 8.0 x 100 = 80.0 per cent; (2) loss = 8.0 - 6.4 = 1.6 g, which is 1.6 / 8.0 x 100 = 20.0 per cent of the theoretical yield; (3) some white magnesium oxide powder is carried out of the crucible by the hot air current, and some magnesium also reacts with nitrogen in the air to form magnesium nitride, so the mass collected is below the theoretical value.

Part 3: Solution concentration and dilution

Data C (simulated): 25.0 g of sodium chloride is dissolved in water to make 500 cm3 of solution; the mass of water added is 475.0 g.

#	Question	Answer and marking notes
9	Calculate the concentration of the solution in g/dm3.	Answer: volume = 500 cm3 = 0.5 dm3, so concentration = 25.0 / 0.5 = 50 g/dm3.
10	Calculate the mass of sodium chloride contained in 120 cm3 of this solution.	Answer: 120 cm3 = 0.12 dm3, so mass = 50 x 0.12 = 6.0 g.
11	200 cm3 of the solution is diluted with water to 500 cm3. (1) Calculate the concentration of the diluted solution in g/dm3. (2) State whether the total mass of solute changes on dilution and give that mass.	Answer: (1) diluted concentration = 50 x 200 / 500 = 20 g/dm3; (2) the total mass of solute is unchanged at 50 x 0.2 = 10 g; only more solvent is added, so the volume rises and the concentration falls.
12	(1) Calculate the percentage concentration by mass of the solution (1 decimal place). (2) How much more water must be added to halve this percentage concentration by mass?	Answer: (1) total mass = 25.0 + 475.0 = 500.0 g, so percentage by mass = 25.0 / 500.0 x 100 = 5.0 per cent; (2) half of that is 2.5 per cent, so the total mass must be 25.0 / 0.03 = 1000.0 g and the water to add is 1000.0 - 500.0 = 500.0 g.

Part 4: Neutralisation and energy change

Data D (simulated): 0.1 mol/dm3 sodium hydroxide solution is used to neutralise 25.0 cm3 of hydrochloric acid; 25.0 cm3 of the same sodium hydroxide solution is needed for complete neutralisation (NaOH + HCl → NaCl + H2O). The mixture has a total mass of 50.0 g and its temperature rises by 6.0 °C. The specific heat capacity of water is 4.2 J g-1 °C-1.

#	Question	Answer and marking notes
13	Calculate the number of moles of NaOH in 25.0 cm3 of 0.1 mol/dm3 sodium hydroxide solution.	Answer: moles = 0.025 dm3 x 0.1 = 0.0025 mol.
14	The mole ratio of NaOH to HCl in the neutralisation is 1:1. Calculate the concentration of the 25.0 cm3 of hydrochloric acid used.	Answer: moles of HCl = 0.0025 mol, so concentration = 0.0025 / 0.025 = 0.10 mol/dm3.
15	If 20.0 cm3 of that hydrochloric acid is used instead, what volume of the same sodium hydroxide solution is needed for complete neutralisation?	Answer: moles of HCl = 0.020 x 0.10 = 0.0020 mol, the moles of NaOH are the same, so the volume needed is 0.0020 / 0.10 = 20.0 cm3, the same as the acid volume.
16	For the neutralisation of 25.0 cm3 of acid with 25.0 cm3 of sodium hydroxide solution: (1) calculate the energy released in joules; (2) calculate the energy released per mole of water formed in kJ/mol (1 decimal place); (3) explain why the measured energy is usually lower than the literature value.	Answer: (1) energy = 50.0 x 4.2 x 6.0 = 1260 J; (2) moles of water formed = 0.0025 mol, so energy per mole = 1260 / 0.0025 = 504000 J/mol = 504.0 kJ/mol; (3) part of the heat is absorbed by the calorimeter and part is lost to the surroundings, so the measured value is lower.

Part 5: Metal displacement and redox

Data E (simulated): 6.5 g of zinc is placed in copper(II) sulphate solution and the zinc reacts completely (Zn + CuSO4 → ZnSO4 + Cu). In a separate tube 4.0 g of copper(II) oxide is heated in a stream of hydrogen until fully reduced (CuO + H2 →

Cu + H2O). Relative atomic masses: Zn = 65, Cu = 64, O = 16, H = 1; molar volume of a gas at room conditions is 24 dm3/mol.

#	Question	Answer and marking notes
17	Calculate the number of moles in 6.5 g of zinc and the mass of copper deposited.	Answer: moles of zinc = 6.5 / 65 = 0.10 mol; the moles of copper are the same, so copper deposited = 0.10 x 64 = 6.4 g.
18	Calculate the change in mass of the solution (how many grams gained or lost) and explain the reason.	Answer: change in solution mass = 6.4 - 6.5 = -0.1 g, that is a loss of 0.1 g; the reason is that each mole of zinc dissolving (65 g) deposits only one mole of copper (64 g), so less mass enters the solution than leaves it.
19	If 6.5 g of zinc reacts completely with excess dilute hydrochloric acid (Zn + 2HCl → ZnCl2 + H2), calculate the volume of hydrogen produced at room conditions in cm3.	Answer: the mole ratio of zinc to hydrogen is 1:1, so moles of hydrogen = 0.10 mol; volume = 0.10 x 24 = 2.4 dm3 = 2400 cm3.
20	For the copper(II) oxide reduction: (1) calculate the moles in 4.0 g of copper(II) oxide; (2) calculate the masses of copper and of water formed; (3) calculate the volume of hydrogen needed at room conditions in dm3.	Answer: (1) Mr(CuO) = 64 + 16 = 80, so moles = 4.0 / 80 = 0.05 mol; (2) copper = 0.05 x 64 = 3.2 g and water = 0.05 x 18 = 0.9 g; (3) the moles of hydrogen are the same, so volume = 0.05 x 24 = 1.2 dm3.

Part 6: Air composition and combustion products

Data F (simulated): oxygen makes up about 21 per cent by volume of air. Burning 12 g of carbon gives carbon dioxide (C + O2 → CO2); burning 32 g of methane (CH4) also gives carbon dioxide (CH4 + 2O2 → CO2 + 2H2O). Relative atomic masses: C = 12, H = 1, O = 16.

#	Question	Answer and marking notes
21	Calculate the volume of oxygen in 200 cm3 of air.	Answer: volume of oxygen = 200 x 21 / 100 = 42 cm3.
22	Burning 12 g of carbon: calculate the mass of carbon dioxide formed.	Answer: moles of carbon = 12 / 12 = 1 mol, so the moles of carbon dioxide are the same and its mass = 1 x 44 = 44 g.
23	Burning 32 g of methane: calculate the moles of methane and the mass of carbon dioxide formed.	Answer: Mr(CH4) = 12 + 4 = 16, so moles = 32 / 16 = 2 mol and the mass of carbon dioxide = 2 x 44 = 88 g.
24	(1) Calculate the mass of carbon dioxide produced per gram of carbon and per gram of methane burnt (2 decimal places). (2) State which fuel releases more carbon dioxide per unit mass. (3) Suggest one practical way to reduce the carbon dioxide emitted by burning fossil fuels.	Answer: (1) per gram of carbon = 44 / 12 = 3.67 g and per gram of methane = 44 / 16 = 2.75 g; (2) per unit mass, carbon releases more carbon dioxide; (3) replace part of the fossil fuel used for electricity generation with renewable energy such as solar and wind power and improve the energy efficiency of buildings and industry, which directly cuts the carbon dioxide from combustion.