S4 Chemistry Practice Paper (Set 1): Mole Calculations and Stoichiometry, Solution Concentration and Dilution, Acid-Base Titration and pH, Redox and Electron Transfer, Reaction Rate and Half-life, Thermochemistry and Enthalpy (100 marks)

S4 Chemistry | Time: 75 minutes | Total: 100 marks | 24 questions

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Section A: Moles and Stoichiometry (16 marks)

Show all working. Molar masses: CaCO3 100, H2 2.0, O2 32, H2O 18, NH4NO3 80, Zn 65. Take the molar volume of a gas at room temperature as 24.0 dm3/mol.

1. 25.0 g of calcium carbonate (CaCO3, molar mass 100 g/mol) is heated until it decomposes completely into calcium oxide and carbon dioxide. Find the moles of calcium carbonate and the volume of carbon dioxide at room temperature.
Moles = 25.0 / 100 = 0.25 mol; 1 mol CaCO3 gives 1 mol CO2, so carbon dioxide = 0.25 mol; volume = 0.25 x 24.0 = 6 dm3.

2. 4.00 g of hydrogen burns completely in excess oxygen to form water (2H2 + O2 → 2H2O). Find the moles and mass of water formed, and the mass of oxygen used.
Hydrogen = 4.00 / 2.0 = 2 mol; water = 2 mol, mass = 2 x 18 = 36 g; oxygen = 1.00 mol, mass = 32 g.

3. Ammonium nitrate (NH4NO3, molar mass 80.0 g/mol) is a common fertiliser. Find the percentage by mass of nitrogen in it.
Each mole contains 2 nitrogen atoms, so nitrogen = 2 x 14.0 = 28.0 g; percentage = 28.0 / 80.0 x 100% = 35%.

4. 0.100 mol of zinc is added to 0.150 mol of hydrochloric acid (Zn + 2HCl → ZnCl2 + H2). Find the limiting reagent, the moles and volume of hydrogen formed, and the moles of zinc left over.
0.100 mol of zinc needs 0.200 mol of acid but only 0.150 mol is present, so the acid is limiting; hydrogen = 0.150 / 2 = 0.075 mol, volume = 0.075 x 24.0 = 1.8 dm3; zinc used = 0.075 mol, left = 0.100 − 0.075 = 0.025 mol.

Section B: Solution Concentration and Dilution (16 marks)

Give concentrations in mol/dm3. The moles of solute do not change on dilution, so use c1V1 = c2V2.

5. 5.85 g of sodium chloride (molar mass 58.5 g/mol) is dissolved in water and made up to 250.0 cm3. Find the concentration.
Moles = 5.85 / 58.5 = 0.100 mol; volume = 0.250 dm3; concentration = 0.100 / 0.250 = 0.4 mol/dm3.

6. 25.0 cm3 of the 0.400 mol/dm3 solution is diluted with water to 250.0 cm3. Find the moles of solute and the new concentration.

Solute = 0.400 x 0.0250 = 0.01 mol; new concentration = 0.01 / 0.250 = 0.04 mol/dm3.

7. 0.0120 g of pesticide is dissolved in 1.50 kg of water. Find the concentration in parts per million by mass.

Mass of water = 1.50 kg = 1500 g; ppm = 0.0120 / 1500 x 10^6 = 8 ppm.

8. A solution has a concentration of 0.150 mol/dm3. A 20.0 cm3 sample is taken and the molar mass of the solute is 40.0 g/mol. Find the moles and the mass of solute in the sample.

Moles = 0.150 x 0.0200 = 0.003 mol; mass = 0.003 x 40.0 = 0.12 g.

Section C: Acids, Bases and Titration (16 marks)

Assume 25 °C and Kw = 1.0 x 10^-14; give pH to two decimal places.

9. Find the hydrogen ion concentration and the pH of 0.010 mol/dm3 hydrochloric acid.

Hydrochloric acid is a strong acid and ionises completely, so [H+] = 0.010 mol/dm3; pH = -log(0.010) = 2.

10. Find the pOH and pH of a 0.0050 mol/dm3 sodium hydroxide solution.

[OH-] = 0.0050 mol/dm3; pOH = -log(0.0050) = 2.3; pH = 14 - 2.3 = 11.7.

11. 20.0 cm3 of 0.100 mol/dm3 hydrochloric acid exactly neutralises 25.0 cm3 of sodium hydroxide solution of unknown concentration. Find that concentration.

Moles of acid = 0.100 x 0.0200 = 0.002 mol; reaction is 1:1, so moles of base are equal; concentration = 0.002 / 0.0250 = 0.08 mol/dm3.

12. A weak monobasic acid HA has concentration 0.100 mol/dm3 and pH 3.00. Find [H+], the degree of ionisation and Ka.

[H+] = 10^-3 = 0.001 mol/dm3; degree of ionisation = 0.001 / 0.100 = 0.01 (1%); Ka = (0.001)^2 / 0.100 = 1e-05.

Section D: Redox and Electron Transfer (16 marks)

Show oxidation numbers or half-equations to justify the electron counts.

13. Find the oxidation number of the stated element: Mn in KMnO4, Cr in K2Cr2O7, S in H2SO4.

KMnO4: (+1) + x + 4(-2) = 0, so x = +7; K2Cr2O7: 2(+1) + 2x + 7(-2) = 0, so x = +6; H2SO4: 2(+1) + x + 4(-2) = 0, so x = +6.

14. The half-equation of permanganate in acid is MnO4- + 8H+ + 5e- → Mn2+ + 4H2O. Find the moles of electrons accepted by 0.0200 mol of MnO4-.

Each mole of MnO4- accepts 5 mol of electrons; electrons = 0.0200 x 5 = 0.1 mol.

15. How many moles of Fe2+ can be oxidised by 25.0 cm3 of 0.0200 mol/dm3 potassium permanganate (Fe2+ → Fe3+ + e-)?

Moles of MnO4- = 0.0200 x 0.0250 = 0.0005 mol; electrons = 0.0005 x 5 = 0 mol; each Fe2+ releases 1 mol of electrons, so Fe2+ = 0 mol.

16. How many moles of potassium permanganate are needed to oxidise 0.0500 mol of Fe2+ completely?

Electrons released by Fe2+ = 0.0500 mol; moles of MnO4- = 0.0500 / 5 = 0.01 mol.

Section E: Reaction Rate and Half-life (18 marks)

For a first-order reaction t1/2 = ln2 / k. A catalyst lowers the activation energy, and k2/k1 = exp((Ea1 - Ea2) / RT) with R = 8.314 J/(mol K).

17. The concentration of a reactant falls from 0.80 mol/dm3 to 0.20 mol/dm3 in 30.0 s. Find the average rate of reaction.

Change in concentration = 0.80 - 0.20 = 0.60 mol/dm3; average rate = 0.60 / 30.0 = 0.02 mol/(dm3 s).

18. An experiment shows that doubling the concentration of A, with all else unchanged, increases the initial rate four times. Find the order of reaction with respect to A and explain.

If rate = k[A]^n then 2^n = 4, so n = 2: the reaction is second order in A and the rate is proportional to [A]^2.

19. A first-order reaction has rate constant k = 0.0231 s-1. Find its half-life to three significant figures.

t1/2 = ln2 / k = 0.6931 / 0.0231 = 30 s.

20. At 300 K a reaction has an activation energy of 80.0 kJ/mol. With a catalyst the activation energy falls to 60.0 kJ/mol. Assuming the pre-exponential factor is unchanged, find the factor by which the rate constant increases.

dEa = 20.0 kJ/mol = 20000 J/mol; k2/k1 = exp(20000 / (8.314 x 300)) = exp(8.02) which is about 3.04e+03, i.e. roughly 3.0 x 10^3 times faster.

Section F: Thermochemistry and Enthalpy (18 marks)

Take the specific heat capacity of water as 4.18 J/(g K) and the density of solutions as 1.00 g/cm3; show heat released with a negative sign.

21. 100 g of water rises in temperature by 6.50 K. Find the heat absorbed in kJ.

q = mc(dT) = 100 x 4.18 x 6.50 = 2717 J = 2.72 kJ.

22. 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 6.70 K. Find the enthalpy change of neutralisation in kJ/mol.

Mass of solution is about 100 g; q = 100 x 4.18 x 6.70 = 2801 J; moles of water formed = 1.00 x 0.0500 = 0.0500 mol; dH = -2801 / 0.0500 = -56 kJ/mol.

23. Given standard enthalpies of formation: $CO_2(g)$ -394, $H_2O(l)$ -286 and $CH_4(g)$ -75 kJ/mol, find the standard enthalpy of combustion of methane ($CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$).

dH = (-394) + 2(-286) - (-75) = -394 - 572 + 75 = -891 kJ/mol.

24. Given bond enthalpies: H-H 436, Cl-Cl 243 and H-Cl 432 kJ/mol, find the enthalpy change of $H_2 + Cl_2 \rightarrow 2HCl$.

Bonds broken = 436 + 243 = 679 kJ/mol; bonds formed = 2 x 432 = 864 kJ/mol; dH = 679 - 864 = -185 kJ/mol.

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