S3 Chemistry Practice Paper: Separation Techniques, Atomic Structure and the Periodic Table, Ionic and Covalent Bonding, Chemical Equations, Acids and Alkalis, Metals and Redox (100 marks)

S3 | Chemistry | Time: 75 minutes | Total: 100 marks | 24 questions

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Section A Classification and Separation

Part A: 4 marks each, 16 marks in total.

1. Explain the difference between a compound and a mixture, and give one example of each. [4]

Answer: A compound is formed from two or more elements chemically combined in a fixed proportion; its components cannot be separated by physical methods, for example water. A mixture contains two or more substances in any proportion and can be separated by physical methods, for example salt solution.

2. In the laboratory, sand is separated from salt solution in two steps. Name the two steps and state the principle of each. [4]

Answer: First, filtration: sand is insoluble, so its particles are trapped by the filter paper while the salt solution passes through as the filtrate. Second, evaporation or crystallisation: heating the filtrate removes water, and the salt crystallises out once the solution becomes saturated.

3. Why can simple distillation separate water from ethanol? Describe the distillate collected. [4]

Answer: Ethanol boils at about 78 °C, lower than the 100 °C of water, so on heating ethanol vaporises first; its vapour is condensed in the condenser and collected as the distillate, which is richer in ethanol while more water remains in the flask.

4. State two experimental methods that can show whether a solid is pure. [4]

Answer: First, measure the melting point: a pure solid melts at a fixed temperature over a very narrow range, while impurities lower the melting point and widen the range. Second, use chromatography: a pure substance gives only one spot on the paper, whereas impurities give several spots.

Section B Atomic Structure and Periodic Table

Part B: 4 marks each, 16 marks in total.

5. Name the three subatomic particles in an atom and state the relative charge and location of each. [4]

Answer: A proton carries a charge of +1 and is found in the nucleus; a neutron carries no charge and is also in the nucleus; an electron carries a charge of -1 and moves rapidly in the electron shells around the nucleus.

6. The atomic number of sodium is 11. Write the electron arrangement of a sodium atom and state its period and group in the periodic table. [4]

Answer: The electron arrangement is 2, 8, 1, giving 11 electrons in total. With one electron in the outermost shell it lies in Period 3, Group I (the alkali metals).

7. Magnesium has three isotopes: Mg-24 79%, Mg-25 10% and Mg-26 11%. Calculate the relative atomic mass of magnesium. [4]

Answer: Relative atomic mass = 24 x 0.79 + 25 x 0.10 + 26 x 0.11 = 18.96 + 2.50 + 2.86 = 24.32.

8. The atomic number of magnesium is 12. Write the symbol and electron arrangement of a magnesium ion and explain why it is formed. [4]

Answer: A magnesium atom has the arrangement 2, 8, 2; by losing the two outermost electrons it becomes Mg2+ with the arrangement 2, 8, the same as neon, and so gains a stable structure.

Section C Ionic and Covalent Bonding

Part C: 4 marks each, 16 marks in total.

9. Compare how ionic bonds and covalent bonds are formed, giving one example of a compound for each. [4]

Answer: An ionic bond forms when a metal atom loses electrons and a non-metal atom gains them, so that oppositely charged ions attract each other, as in sodium chloride. A covalent bond forms when two non-metal atoms share a pair of electrons, as in carbon dioxide.

10. Write the chemical formulae of magnesium chloride and aluminium oxide, and explain briefly how the ionic charges lead to these formulae. [4]

Answer: Magnesium chloride is MgCl2: Mg2+ and Cl- have charges in the ratio 2:1, so two chloride ions pair with one magnesium ion. Aluminium oxide is Al2O3: the lowest common multiple of the charges of Al3+ and O2- is 6, so two aluminium ions pair with three oxide ions.

11. Sodium chloride melts at 801 °C, does not conduct when solid, but conducts when molten or dissolved. Explain these observations. [4]

Answer: Sodium chloride has a giant ionic lattice in which each ion is attracted to several ions of opposite charge; these strong electrostatic forces need a high temperature to overcome. In the solid the ions are held in fixed positions and cannot move, so it does not conduct, but when molten or dissolved the ions are free to move and carry charge.

12. Compare the structure and melting point of carbon dioxide and diamond, and explain the difference. [4]

Answer: Carbon dioxide has a simple molecular structure with only weak forces between its molecules, so it has a very low melting point and sublimes. Diamond has a giant covalent structure in which each carbon atom is bonded to four others; melting it requires breaking many strong covalent bonds, so its melting point is extremely high.

Section D Chemical Reactions and Equations

Part D: 4 marks each, 16 marks in total.

13. Calcium carbonate reacts with dilute hydrochloric acid. Write a balanced equation and name the gas produced. [4]

Answer: CaCO3 + 2HCl -> CaCl2 + H2O + CO2. The gas is carbon dioxide, which turns limewater milky.

14. 12 g of carbon burns completely in 32 g of oxygen to give 44 g of carbon dioxide. If only 24 g of oxygen is available, calculate the mass of carbon left unburnt and the mass of carbon dioxide formed. [4]

Answer: 24 g of oxygen is 24 / 32 = 0.75 mol. Carbon reacts with oxygen in a 1:1 ratio, so only 0.75 mol of carbon, that is 9.0 g, can burn; 12 - 9.0 = 3.0 g of carbon is left. The carbon dioxide formed is 0.75 x 44 = 33.0 g.

15. What is the role of a catalyst in a chemical reaction, and is it used up? [4]

Answer: A catalyst provides a reaction pathway of lower activation energy, so the reaction goes faster; its mass and chemical properties are unchanged before and after the reaction, so it is not used up.

16. Classify each reaction as combination, decomposition, displacement or neutralisation, giving one example of each: (a) iron with copper(II) sulphate solution; (b) heating calcium carbonate. [4]

Answer: (a) Displacement: Fe + CuSO4 -> FeSO4 + Cu, the more reactive iron displaces copper. (b) Decomposition: CaCO3 -> CaO + CO2, one compound breaks down into two substances.

Section E Acids, Alkalis and Neutralisation

Part E: 4 marks each, 16 marks in total.

17. Define an acid and an alkali, giving one common example of each and the pH range of each. [4]

Answer: An acid releases hydrogen ions when dissolved in water, for example hydrochloric acid, and has a pH below 7. An alkali releases hydroxide ions when dissolved in water, for example sodium hydroxide, and has a pH above 7.

18. State the colour of phenolphthalein and methyl orange in acidic, neutral and alkaline solutions. [4]

Answer: Phenolphthalein is colourless in acid and neutral solution and pink in alkali. Methyl orange is red in acid, orange in neutral solution and yellow in alkali.

19. What volume of 0.050 mol/dm3 hydrochloric acid is needed to neutralise 25.0 cm3 of 0.100 mol/dm3 sodium hydroxide solution? [4]

Answer: Moles of NaOH = 25.0 / 1000 x 0.100 = 0.0025 mol. HCl and NaOH react in a 1:1 ratio, so 0.0025 mol of HCl is needed; volume = 0.0025 / 0.050 x 1000 = 50.0 cm3.

20. When dilute hydrochloric acid is added to calcium carbonate, state two observations and name the type of reaction. [4]

Answer: The solid gradually dissolves and bubbles of carbon dioxide are given off, and the mixture warms slightly. This is the reaction of an acid with a carbonate, a form of neutralisation.

Section F Metals and Redox

Part F: 5 marks each, 20 marks in total.

21. A zinc strip is placed in copper(II) sulphate solution. State the observations, write the equation and name the type of reaction. [5]

Answer: The zinc becomes coated with a reddish-brown solid and the blue colour of the solution fades: Zn + CuSO4 -> ZnSO4 + Cu. Zinc is more reactive than copper in the reactivity series, so it displaces copper from its salt solution in a displacement reaction.

22. Explain why some metals can be extracted by reducing their oxides with carbon, while others must be extracted by electrolysis. [5]

Answer: Metals less reactive than carbon, such as iron and zinc, have oxides that carbon can reduce, for example Fe2O3 + 3CO -> 2Fe + 3CO2. Metals more reactive than carbon, such as aluminium and sodium, form very stable oxides that carbon cannot reduce, so they must be extracted by electrolysis.

23. State the two conditions needed for iron to rust, and describe three methods of preventing rusting. [5]

Answer: Iron rusts only when both oxygen and water are present. Methods of prevention: coat with paint or grease to keep out air and water; galvanise with zinc for sacrificial protection; use stainless steel or chromium plating so that a protective layer forms.

24. Using Zn + Cu2+ -> Zn2+ + Cu, explain oxidation and reduction in terms of electron transfer, and identify the oxidising and reducing agents. [5]

Answer: The zinc atom loses two electrons and is oxidised to Zn2+, so zinc is the reducing agent; the copper ion gains two electrons and is reduced to copper, so Cu2+ is the oxidising agent. Oxidation is loss of electrons and reduction is gain of electrons, and both occur at the same time.
