THE UNITED REPUBLIC OF TANZANIA

NATIONAL EXAMINATIONS COUNCIL OF TANZANIA

ADVANCED CERTIFICATE OF SECONDARY EDUCATION EXAMINATION

132/1 **CHEMISTRY 1**

(For Both School and Private Candidates)

Time: 3 Hours **ANSWERS** **Year: 2020**

Instructions

1. This paper consists of sections A and B with total of ten questions

2. Each question carries ten marks in section A and fifteen marks in section B.

1. (a) By using chemical equations, describe the following:

(i) Amphoteric oxides.

Amphoteric oxides react with both acids and bases to form salts and water. Examples include aluminum oxide (Al_2O_3) and zinc oxide (ZnO).

Reaction with acid:
$Al_2O_3 + 6HCl \text{---->} 2AlCl_3 + 3H_2O$

Reaction with base:
$Al_2O_3 + 2NaOH + 3H_2O \text{---->} 2Na[Al(OH)_4]$

(ii) Basic oxides.

Basic oxides react with acids to form salts and water. Examples include calcium oxide (CaO) and sodium oxide (Na_2O).

$CaO + 2HCl \text{---->} CaCl_2 + H_2O$
$Na_2O + H_2SO_4 \text{---->} Na_2SO_4 + H_2O$

(b) With the aid of chemical equation(s), explain how you can prepare soluble chlorides.

Soluble chlorides can be prepared by reacting metals, metal oxides, or metal carbonates with hydrochloric acid.

1. Metal + HCl
$Zn + 2HCl \text{---->} ZnCl_2 + H_2$

2. Metal oxide + HCl
$CuO + 2HCl \text{---->} CuCl_2 + H_2O$

3. Metal carbonate + HCl
$CaCO_3 + 2HCl \text{---->} CaCl_2 + CO_2 + H_2O$

(c) A student suggested a methodology to prepare aluminum trichloride ($AlCl_3$) by following the steps below:

(i) Add sodium hydroxide solution to aluminum sulfate solution.
(ii) Filter off the precipitate.
(iii) Convert the hydroxide into a chloride by adding hydrochloric acid.
(iv) Evaporate the solution to leave crystals of $AlCl_3$.

Could this process work? Justify.

This process will not work because $AlCl_3$ is highly deliquescent and hydrolyzes in water, forming $Al(OH)_3$ and HCl. Instead, $AlCl_3$ is prepared by reacting aluminum metal with dry chlorine gas.

$2Al + 3Cl_2$ ----> $2AlCl_3$

2. (a) Explain two detrimental effects of excessive salts in the soil.

1. Osmotic stress: Excess salts reduce the ability of plant roots to absorb water, leading to dehydration and stunted growth.
2. Toxicity: High concentrations of salts such as Na^+ and Cl^- interfere with nutrient uptake, leading to nutrient imbalances and plant toxicity.

(b) Give four advantages of adding organic manures in the soil.

1. Improves soil structure by enhancing aeration and water retention.
2. Increases microbial activity, promoting soil fertility.
3. Supplies essential nutrients gradually, reducing the need for chemical fertilizers.
4. Enhances soil pH buffering capacity, preventing extreme acidity or alkalinity.

(c) 20 g of a soil sample was shaken with 20 cm^3 of 0.1 M HCl solution. After filtering and washing the soil, the filtrate required 27 cm^3 of 0.1 M NaOH solution for complete neutralization. The total Cation Exchange Capacity (CEC) of the soil is 29 mill. eq per 100 g of the soil sample. Calculate the Percent Base Saturation (PBS) of the soil sample.

Step 1: Calculate milliequivalents of NaOH used.

Milliequivalents = Volume (mL) × Molarity
= 27 × 0.1
= 2.7 meq

Step 2: Calculate PBS.

PBS = (Milliequivalents of base / CEC) × 100
PBS = (2.7 / 29) × 100
PBS = 9.31%

3. (a) (i) Name two major natural sources of organic compounds.

1. Petroleum
2. Coal

(ii) Briefly, explain three properties of carbon element that makes it able to form a large number of compounds.

1. Catenation: Carbon atoms can form long chains and rings by bonding with other carbon atoms.
2. Tetravalency: Carbon has four valence electrons, allowing it to form stable single, double, and triple bonds.
3. Ability to form multiple bonds: Carbon can form strong covalent bonds with elements like hydrogen, oxygen, nitrogen, and halogens.

(b) A form six student gave the following names for various substituted aliphatic hydrocarbons:

(i) 2-methyl-3-bromobutane
(ii) 3,3-dimethyl-2-chlorobutane
(iii) 4-chloro-3-bromo-2-pentene
(iv) 2-methyl-4-butyne

The names do not follow IUPAC rules. Draw the structure suggested by the incorrect names and assign the correct name for each compound.

(i) Correct name: 3-Bromo-2-methylbutane
(ii) Correct name: 2-Chloro-2,3-dimethylbutane
(iii) Correct name: 3-Bromo-4-chloropent-2-ene
(iv) Correct name: 4-Methylpent-2-yne

(c) Explain each of the following observations:

(i) Methylpropane has a lower boiling point than butane although both have the same molecular mass.

Methylpropane is more branched, leading to weaker van der Waals forces compared to the linear structure of butane, resulting in a lower boiling point.

(ii) Ignition sources such as smoking are not allowed at petrol stations.

Petrol is highly volatile and produces flammable vapors that can ignite easily in the presence of an open flame or spark, causing fire or explosion.

4. (a) Diborane (B_2H_6) is very reactive such that it was once considered as a possible rocket fuel for U.S space programs. The overall equation for the synthesis of diborane is:

$2B(s) + 3H_2(g) \dashrightarrow B_2H_6(g)$

Use the following data to calculate the enthalpy change of formation of B_2H_6 from its elements:

$2B(s) + 3/2O_2(g) \text{ ----> } B_2O_3(s)$ $\Delta H = -1273$ kJ
$B_2H_6(g) + 3O_2(g) \text{ ----> } B_2O_3(s) + 3H_2O(g)$ $\Delta H = -2035$ kJ
$H_2(g) + 1/2O_2(g) \text{ ----> } H_2O(l)$ $\Delta H = -286$ kJ
$H_2O(l) \text{ ----> } H_2O(g)$ $\Delta H = -44$ kJ

Step 1: Reverse the combustion reaction of B_2H_6.

$B_2O_3(s) + 3H_2O(g) \text{ ----> } B_2H_6(g) + 3O_2(g)$ $\Delta H = +2035$ kJ

Step 2: Use Hess's Law to sum enthalpies.

$\Delta Hf(B_2H_6) = \Delta H_2 - \Delta H_1 - (3 \times \Delta H_3) - (3 \times \Delta H_4)$
$= 2035 - (-1273) - (3 \times -286) - (3 \times -44)$
$= 2035 + 1273 - 858 - 132$
$= 2318$ kJ

(b) When 100 cm³ of 1 M KOH and 100 cm³ of 1 M HCl were mixed in a calorimeter, temperature rose by 6.25 K. Given that the heat capacity of the calorimeter was 95 J/K and specific heat capacity of the solution mixture was 4.2 J/g·K, calculate the standard enthalpy of neutralization. Assume that the density of the solution is equal to the density of water.

Step 1: Calculate the total heat released.
$Q = (\text{mass} \times \text{specific heat} \times \Delta T) + (\text{calorimeter heat capacity} \times \Delta T)$

Total mass = 100 g + 100 g = 200 g

$Q = (200 \times 4.2 \times 6.25) + (95 \times 6.25)$
$= 5250 + 593.75$
$= 5843.75$ J

Step 2: Calculate enthalpy change per mole.

Moles of HCl = $0.1 \times 1 = 0.1$ mol

$\Delta H = Q / \text{moles}$
$= 5843.75 / 0.1$
$= -58.44$ kJ/mol

5. (a) (i) State four postulates of Dalton's atomic theory.

1. All matter is composed of small, indivisible particles called atoms.
2. Atoms of the same element are identical in mass and properties, while atoms of different elements have distinct masses and properties.
3. Atoms cannot be created, destroyed, or transformed into atoms of another element in a chemical reaction.
4. Atoms combine in simple whole-number ratios to form compounds.

(ii) Why different atoms have different chemical properties. Briefly explain.

Different atoms have different chemical properties due to variations in their electronic configurations, particularly the number of valence electrons. The number and arrangement of valence electrons determine an element's reactivity, bonding behavior, and chemical properties.

(b) Write the chemical symbol (Z X) and orbital electronic configuration for the atoms described in the following table:

S/N	Number of Neutrons	Number of Electrons
(i)	13	11
(ii)	7	8
(iii)	17	18
(iv)	16	16

(i) sodium (Na): $^{23}_{11}Na$, $1s^2\ 2s^2\ 2p^6\ 3s^1$
(ii) oxygen (O): $^{15}_{8}O$, $1s^2\ 2s^2\ 2p^4$
(iii) chloride ion (Cl^-): $^{35}_{17}Cl^-$, $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
(iv) sulfur (S): $^{32}_{16}S$, $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

(c) Calculate the minimum energy required to remove an electron from the hydrogen atom in its ground state.

Using the energy equation for hydrogen:

$E = -13.6\ eV/n^2$

For n = 1 (ground state):

$E = -13.6\ eV$

Since energy is required to remove the electron, the minimum energy needed is +13.6 eV.

6. (a) Differentiate:

(i) Electrovalent bond from octet rule.

An electrovalent bond (ionic bond) is formed when one atom donates electrons to another, creating oppositely charged ions that attract each other.

The octet rule states that atoms tend to gain, lose, or share electrons to achieve a stable configuration of eight valence electrons.

(ii) Lone pair from bonding pair of electrons.

A lone pair consists of valence electrons that are not shared with another atom in a bond, while a bonding pair consists of electrons shared between two atoms to form a covalent bond.

(b) Using sketches, briefly explain three possible overlaps that can occur during sigma bond formation.

1. s-s overlap: Two s orbitals from different atoms overlap, forming a sigma bond (e.g., H-H in H_2).
2. s-p overlap: An s orbital from one atom overlaps with a p orbital from another atom (e.g., H-Cl in HCl).
3. p-p overlap: Two p orbitals from different atoms overlap along the internuclear axis, forming a sigma bond (e.g., Cl-Cl in Cl_2).

(c) Give two reasons for the observed difference in bond strength between sigma and pi bonds in compounds.

1. Extent of orbital overlap: Sigma bonds result from head-on overlap of atomic orbitals, creating a stronger bond, while pi bonds result from side-by-side overlap, which is weaker.
2. Electron density distribution: In sigma bonds, electron density is concentrated along the bond axis, leading to greater stability, whereas in pi bonds, electron density is above and below the bond axis, making them more reactive.

(d) Predict the geometry of ammonia, basing on the Valence Shell Electron Pair Repulsion (VSEPR) theory.

Ammonia (NH_3) has trigonal pyramidal geometry due to the presence of three bonding pairs and one lone pair of electrons on nitrogen. The lone pair repels the bonding pairs, reducing the bond angle from the ideal 109.5° (tetrahedral) to approximately 107°.

7. (a) A 0.25 moles of air has entered a diesel engine at a pressure of 1.05×10^5 Pa and a temperature of 27°C. Assuming that air is ideal, calculate:

(i) the volume it occupies.

Using the ideal gas equation:

$PV = nRT$

Where:
$P = 1.05 \times 10^5$ Pa
$V = ?$
$n = 0.25$ moles
$R = 8.314$ J/mol·K
$T = 27 + 273 = 300$ K

$V = (nRT) / P$
$V = (0.25 \times 8.314 \times 300) / (1.05 \times 10^5)$
$V = (623.55) / (1.05 \times 10^5)$
$V = 5.93 \times 10^{-3}$ m^3 or 5.93 L

(ii) its temperature, immediately after compression to one twentieth of its original volume where the pressure rises to 7.0×10^6 Pa.

Using the combined gas law:

$(P_1V_1) / T_1 = (P_2V_2) / T_2$

Where:
$P_1 = 1.05 \times 10^5$ Pa
$V_1 = 5.93 \times 10^{-3}$ m^3
$T_1 = 300$ K
$P_2 = 7.0 \times 10^6$ Pa
$V_2 = V_1 / 20$

Solving for T_2:

$T_2 = (P_2V_2T_1) / (P_1V_1)$
$T_2 = (7.0 \times 10^6 \times (5.93 \times 10^{-3} / 20) \times 300) / (1.05 \times 10^5 \times 5.93 \times 10^{-3})$
$T_2 = (7.0 \times 10^6 \times 2.965 \times 10^{-4} \times 300) / (1.05 \times 10^5 \times 5.93 \times 10^{-3})$
$T_2 = (6.22 \times 10^2) / (6.23 \times 10^2)$
$T_2 \approx 3571$ K

(b) A 42 g of nitrogen gas and 8 g of hydrogen gas are mixed in a 10-litre vessel at 20°C. Calculate the partial pressure of each gas and the total pressure of the gas mixture.

Step 1: Calculate moles of each gas.

Moles of N_2 = 42 g / 28 g/mol = 1.5 mol

Moles of H_2 = 8 g / 2 g/mol = 4 mol

Step 2: Use the ideal gas equation to calculate partial pressures.

$PV = nRT$
$P = (nRT) / V$

For nitrogen:

$P_{N_2} = (1.5 \times 8.314 \times 293) / 10$
$P_{N_2} = (3651.46) / 10$
$P_{N_2} = 365.15$ kPa

For hydrogen:

$P_{H_2} = (4 \times 8.314 \times 293) / 10$
$P_{H_2} = (9737.57) / 10$
$P_{H_2} = 973.76$ kPa

Total pressure:

$P_{total} = P_{N_2} + P_{H_2}$
$P_{total} = 365.15 + 973.76$
$P_{total} = 1338.91$ kPa

8. (a) (i) Differentiate between cryoscopic constant and ebullioscopic constant.

Cryoscopic constant (K_f) is the proportionality constant used to calculate the freezing point depression of a solution.

Ebullioscopic constant (K_b) is the proportionality constant used to calculate the boiling point elevation of a solution.

(ii) Derive an expression relating the van't Hoff factor (i) and the degree of dissociation (α).

$i = 1 + (n - 1)\alpha$

where,
i = van't Hoff factor,
n = number of particles the solute dissociates into,
α = degree of dissociation.

(iii) Briefly, explain the effect of degree of dissociation of a solute on the boiling point of a solution.

A higher degree of dissociation increases the number of particles in solution, leading to a greater elevation in boiling point due to colligative properties.

(b) What mass of ethylene glycol, $C_2H_6O_2$, the main component of antifreezing agent which must be added to a 10.0 L of water to produce a solution for use in a car radiator, that freezes at -23.3°C? Assume that the density of water is exactly 1 g/mL and cryoscopic constant is 1.86°C·kg/mol.

Step 1: Use the freezing point depression formula.

$\Delta T\_f = Kf \times m$

$23.3 = 1.86 \times m$

$m = 23.3 / 1.86$
$m = 12.53$ mol/kg

Step 2: Calculate the mass of ethylene glycol required.

Molar mass of $C_2H_6O_2$ = 62 g/mol

Mass = m × molar mass × mass of solvent (kg)
Mass = 12.53 × 62 × 10
Mass = 7770 g or 7.77 kg

(c) A 0.120 g of haemoglobin, the protein which carries oxygen in the blood was dissolved in 200 cm^3 of benzene at 20°C. The solution exerted an osmotic pressure of 25.6 Pa. Find the relative molecular mass of haemoglobin.

Using the osmotic pressure formula:

$\pi = cRT$

$c = \pi / RT$

where,
$\pi = 25.6$ Pa,
$R = 8.314$ J/mol·K,
$T = 293$ K,
$V = 0.2$ L,
$n = m / M$.

Rearrange to find M:

$M = (mRT) / (\pi V)$

$M = (0.120 \times 8.314 \times 293) / (25.6 \times 0.2)$
$M = (292.55) / (5.12)$
M = 57177 g/mol

(d) A 1% solution of sodium chloride freezes at -0.604°C. Calculate the degree of dissociation of the sodium chloride if the molal freezing point depression constant for water is 1.86°C·kg/mol.

Using the formula:

$\Delta T\_f = i \times Kf \times m$

$0.604 = i \times 1.86 \times (0.17)$
Solving for i:

$i = 0.604 / (1.86 \times 0.17)$
i = 1.9

Using the dissociation formula:

$i = 1 + \alpha(n - 1)$
$1.9 = 1 + \alpha(2 - 1)$
$\alpha = 0.9$ or 90%

9. (a) (i) Briefly explain the effect of change of pressure to a system at equilibrium.

If pressure is increased, the equilibrium shifts toward the side with fewer gas molecules. If pressure is decreased, the equilibrium shifts toward the side with more gas molecules.

(ii) Why does a Coca-Cola soda fizz out when its bottle is opened? Give a reason.

When the bottle is opened, pressure decreases, causing dissolved CO_2 to escape as bubbles due to the shift in equilibrium.

(b) Consider the following equilibrium:
$2CrO_4^{2-} (aq) + 2H^+ (aq) \rightleftharpoons Cr_2O_7^{2-} (aq) + H_2O (l)$
yellow orange

What would be observed when:
(i) dilute NaOH is added to the equilibrium mixture?

The solution turns yellow as OH^- ions react with H^+, shifting equilibrium left.

(ii) dilute HCl is added to the equilibrium mixture?

The solution turns orange as H^+ concentration increases, shifting equilibrium right.

(c) The following equilibrium was established during the preparation of phosphorus(V) pentachloride:

(i) Calculate the equilibrium constant for this reaction at 127°C if the equilibrium concentrations observed at this temperature were 0.034 mol/L for ammonia, 0.85 mol/L for nitrogen and 0.031 mol/L for hydrogen.

$Kc = [N_2][H_2]^3 / [NH_3]^2$
$Kc = (0.85 \times 0.031^3) / (0.034^2)$
$Kc = 2.2$

(ii) Using the same equilibrium concentrations in (d) (i), calculate the equilibrium constant at 127°C for the equilibrium:
$2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$
Kc' = (Kc)^(1/2)

Kc' = (2.2)^(1/2)
Kc' = 1.48

10. (a) Why benzene though highly unsaturated, it does not undergo addition reactions? Briefly explain.

Benzene does not undergo addition reactions because of its delocalized π-electron system, which provides extra stability due to resonance. Addition reactions would disrupt this stable aromatic system, making the process unfavorable. Instead, benzene undergoes electrophilic substitution reactions to preserve its aromaticity.

(b) Name two examples in each of the following chemical groups:

(i) Ortho-para directors.

1. Methyl group ($-CH_3$)
2. Hydroxyl group (-OH)

(ii) Meta directors.

1. Nitro group ($-NO_2$)
2. Carboxyl group (-COOH)

(c) Why ortho-para directing groups are called activating groups and meta-directing groups are called deactivating groups? Briefly explain.

Ortho-para directing groups are called activating groups because they donate electron density to the benzene ring through inductive or resonance effects, making the ring more reactive toward electrophilic substitution.

Meta-directing groups are called deactivating groups because they withdraw electron density from the benzene ring, reducing its reactivity toward electrophilic substitution and directing incoming electrophiles to the meta position.

(d) Determine the structural formulae of the following compounds:

(i) 1,3,5-trimethylbenzene.

This structure consists of a benzene ring with three methyl ($-CH_3$) groups attached at positions 1, 3, and 5.

(ii) (1-methylethyl)benzene.

This structure consists of a benzene ring with an isopropyl ($-CH(CH_3)_2$) group attached at position 1.

(e) Give the IUPAC names of the following aromatic compounds:

(i) 1,3-Dimethylbenzene (m-xylene)
(ii) 4-Methyl-1-nitrobenzene (p-nitrotoluene)
(iii) 1,3,5-Trichlorobenzene