

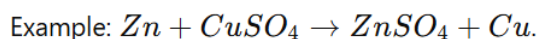
**CHEMISTRY ANSWER KEY**  
**STD IX Second Term model paper 1**

**Section A: Answer any 4 questions.**

(4 × 1 = 4 Marks)

1. Define redox reactions with an example:

A redox reaction is a chemical reaction in which one substance gets oxidized (loses electrons) and another gets reduced (gains electrons).



2. Write the formula for the rate of a chemical reaction:

$$\text{Rate} = \frac{\Delta[\text{Reactant/Product}]}{\Delta t}$$

3. Identify the oxidizing and reducing agents in  $2Mg + O_2 \rightarrow 2MgO$ :

- Oxidizing agent:  $O_2$
- Reducing agent:  $Mg$

4. Name the scientist who proposed the law of conservation of mass:

Antoine Lavoisier.

5. Element with the highest electronegativity from  $F, Cl, Br, I$ :

Fluorine (F).

**Section B: Answer any 4 questions.**

(4 × 2 = 8 Marks)

6. Differentiate between endothermic and exothermic reactions:

- **Endothermic Reaction:** Absorbs heat (e.g.,  $NH_4Cl + H_2O \rightarrow NH_4^+ + Cl^-$ ).
- **Exothermic Reaction:** Releases heat (e.g.,  $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + \text{heat}$ ).

7. Structure of water showing polar covalent bonds:

- Diagram:  $H : O : H$  with  $\delta^-$  on O and  $\delta^+$  on H atoms.

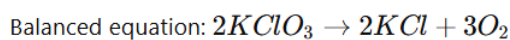
8. Define "activation energy" with a diagram:

- Activation energy is the minimum energy required for a chemical reaction to occur.
- Diagram: Energy profile showing reactants, activation energy peak, and products.

9. Effect of temperature on reaction rate:

- Higher temperature increases the kinetic energy of particles, leading to more frequent and energetic collisions, thus increasing the reaction rate.

10. Balance  $KClO_3 \rightarrow KCl + O_2$ :



## Section C: Answer any 4 questions.

(4 × 3 = 12 Marks)

### 11. Reaction mechanism:

- Step 1: Reactants combine to form an intermediate.
- Step 2: The intermediate decomposes to give the product.  
Example:  $NO_2 + CO \rightarrow NO + CO_2$ .

### 12. Oxidation number of sulfur in $H_2SO_4$ :

- $H = +1, O = -2$ .
- $2(+1) + S + 4(-2) = 0$ .
- $S = +6$ .

### 13. Catalysts and their role:

Catalysts increase the rate of reaction without being consumed.

- Example:  $MnO_2$  in the decomposition of  $H_2O_2$ .

### 14. Homogeneous vs. Heterogeneous Catalysts:

- **Homogeneous Catalyst:** Same phase as reactants (e.g.,  $H_2SO_4$  in esterification).
- **Heterogeneous Catalyst:** Different phase (e.g.,  $Pt$  in hydrogenation).

### 15. Differences between oxidation and reduction:

- Oxidation: Loss of electrons, increase in oxidation state.
- Reduction: Gain of electrons, decrease in oxidation state.
- Example:  $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$ .

## Section D: Answer any 4 questions.

(4 × 4 = 16 Marks)

16. Rate equation  $Rate = k[A]^2[B]$ :

- Order of reaction:  $2 + 1 = 3$ .
- Rate calculation:  
 $Rate = 0.5[1]^2[2] = 1 \text{ mol/L/s}$ .

17. Factors affecting reaction rates:

- Nature of reactants.
- Concentration of reactants.
- Temperature.
- Catalyst.
- Surface area (for solids).

18. Antoine Lavoisier's contribution:

- Law of conservation of mass: Mass is neither created nor destroyed in a chemical reaction.
- Example:  $2H_2 + O_2 \rightarrow 2H_2O$ , total mass of reactants equals total mass of products.

19. Redox reactions and electron transfer:

- $Fe + CuSO_4 \rightarrow FeSO_4 + Cu$ :
  - $Fe$  gets oxidized:  $Fe \rightarrow Fe^{2+} + 2e^-$ .
  - $Cu^{2+}$  gets reduced:  $Cu^{2+} + 2e^- \rightarrow Cu$ .

20. Energy profile diagram for an exothermic reaction:

- Diagram: Energy of products is lower than reactants, showing release of energy.
- Example: Combustion of methane:  $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + \text{heat}$ .