Redox Reactions JEE Main PYQ - 2

Total Time: 1 Hour **Total Marks:** 100

Instructions

Instructions

1. Test will auto submit when the Time is up.
2. The Test comprises of multiple choice questions (MCQ) with one or more correct answers.
3. The clock in the top right corner will display the remaining time available for you to complete the examination.

Navigating & Answering a Question

1. The answer will be saved automatically upon clicking on an option amongst the given choices of answer.
2. To deselect your chosen answer, click on the clear response button.
3. The marking scheme will be displayed for each question on the top right corner of the test window.

Redox Reactions

1. **The incorrect statements among the following is:** **(+4, -1)**

a. PH_3 shows lower proton affinity than NH_3.

b. PF_3 exists but NF_5 does not.

c. NO_2 can dimerise easily.

d. SO_2 can act as an oxidizing agent, but not as a reducing agent.

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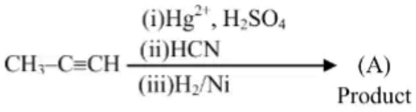
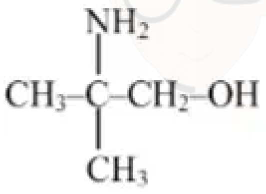
2. **The product (A) formed in the following reaction sequence is:** **(+4, -1)**



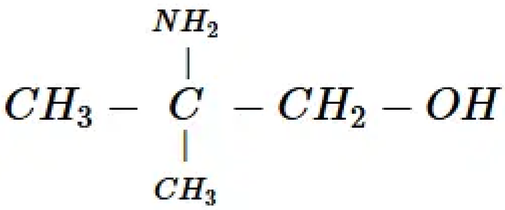
$$CH_3-C\equiv CH \xrightarrow[\text{(iii)}H_2/Ni]{\text{(i)}Hg^{2+},\ H_2SO_4 \quad \text{(ii)}HCN} \underset{\text{Product}}{(A)}$$

a.



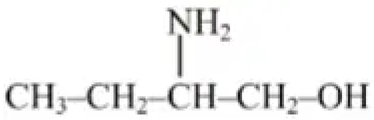
$$CH_3-\overset{\overset{NH_2}{|}}{\underset{\underset{CH_3}{|}}{C}}-CH_2-OH$$

b.



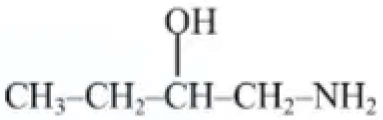
$$CH_3-\overset{\overset{NH_2}{|}}{\underset{\underset{CH_3}{|}}{C}}-CH_2-OH$$

c.



$$CH_3-CH_2-\overset{\overset{NH_2}{|}}{CH}-CH_2-OH$$

d.



$$CH_3-CH_2-\overset{\overset{OH}{|}}{CH}-CH_2-NH_2$$

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3. Which of the following arrangements with respect to their reactivity in nucleophilic addition reaction is correct? **(+4, -1)**

a. benzaldehyde $<$ acetophenone $<$ p-nitrobenzaldehyde $<$ p-tolualdehyde

b. acetophenone $<$ benzaldehyde $<$ p-tolualdehyde $<$ p-nitrobenzaldehyde

c. acetophenone $<$ p-tolualdehyde $<$ benzaldehyde $<$ p-nitrobenzaldehyde

d. p-nitrobenzaldehyde $<$ benzaldehyde $<$ p-tolualdehyde $<$ acetophenone

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4. The heat of formation of $\mathrm{SO_2}(g)$ is given by: **(+4, -1)**

$$S(g) + \frac{3}{2}O_2(g) \rightarrow SO_3(g) + 2x\,\text{kcal}$$

$$SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g) + y\,\text{kcal}$$

a. $\frac{2x}{y}\,\text{kcal}$

b. $x + y\,\text{kcal}$

c. $y - 2x\,\text{kcal}$

d. $2x + y\,\text{kcal}$

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5. Given below are two statements: **(+4, -1)**

Statement (I): The first ionization energy of Pb is greater than that of Sn.
Statement (II): The first ionization energy of Ge is greater than that of Si.

In light of the above statements, choose the correct answer from the options given below:

a. Statement I is true but Statement II is false

b. Both Statement I and Statement II are false

c. Statement I is false but Statement II is true

d. Both Statement I and Statement II are true

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6. For a $Mg|Mg^{2+}(aq)||Ag^{+}(aq)|Ag$, the correct Nernst Equation is: **(+4, -1)**

a. $E_{cell} = E^{o}_{cell} + \frac{RT}{2F}\ln\frac{[Ag^{+}]}{[Mg^{2+}]}$

b. $E_{cell} = E^{o}_{cell} - \frac{RT}{2F}\ln\frac{[Ag^{+}]}{[Mg^{2+}]}$

c. $E_{cell} = E^{o}_{cell} - \frac{RT}{2F}\ln\frac{[Mg^{2+}]}{[Ag^{+}]}$

d. $E_{cell} = E^{o}_{cell} + \frac{RT}{2F}\ln[Ag^{+}]^{2}$

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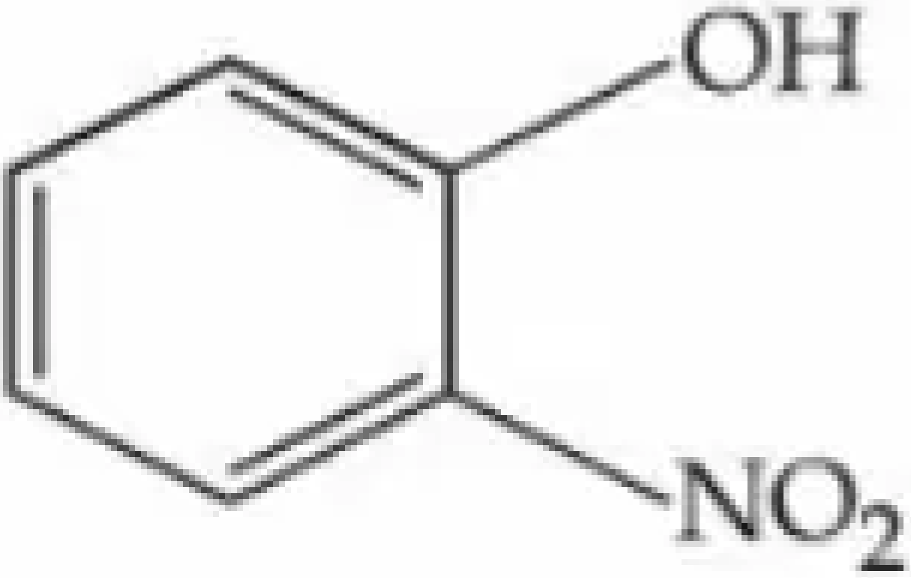
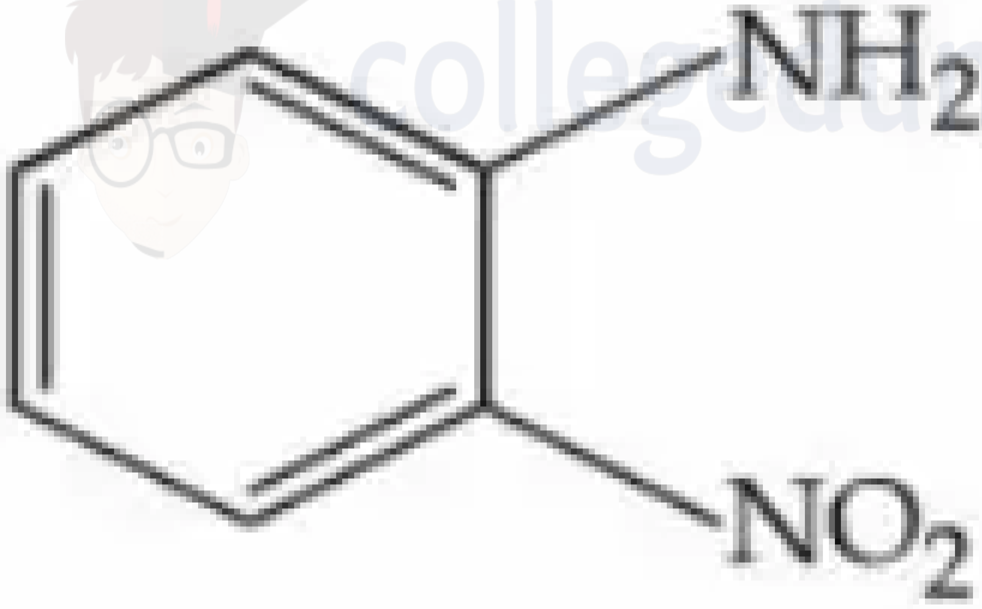
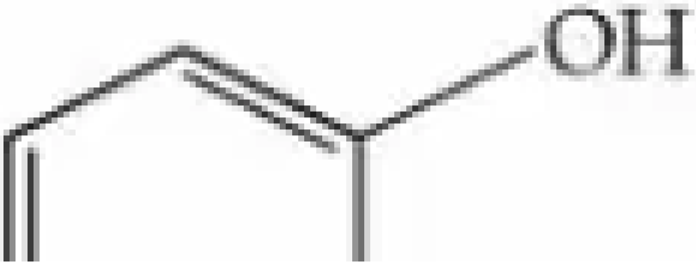
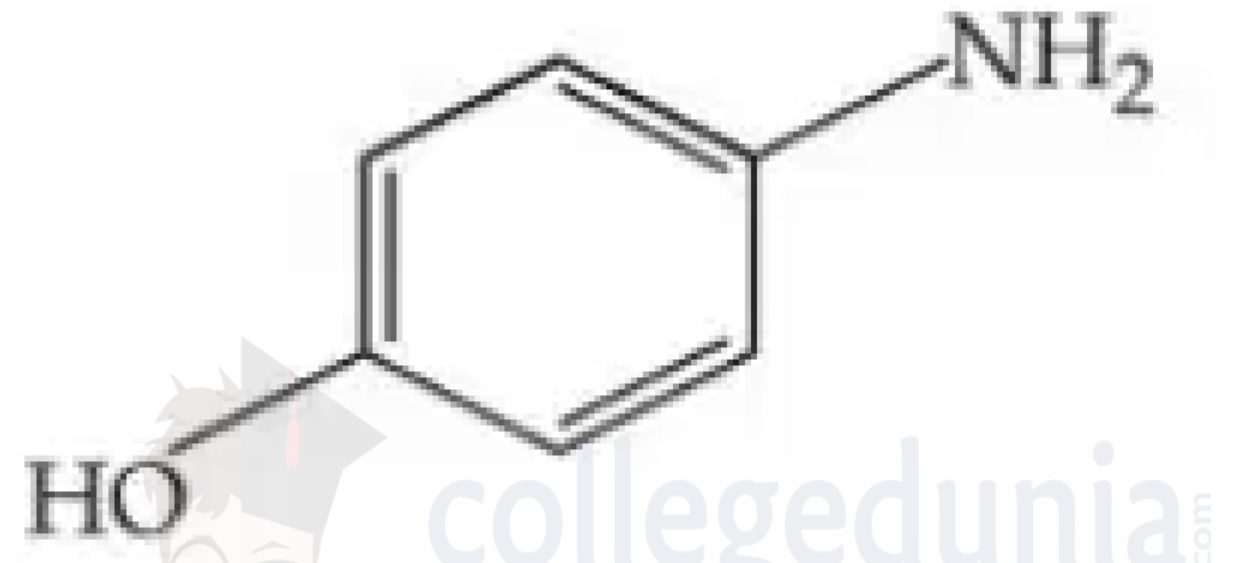
7. The steam volatile compounds among the following are: **(+4, -1)**


OH
NO_2


NH_2
NO_2


OH



Choose the correct answer from the options given below:

a. (A) and (B) Only

b. $(A), (B)$ and (C) Only

c. (B) and (D) Only

d. (A) and (C) Only

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8. **Match the LIST-I with LIST-II (Redox Reactions).** **(+4, -1)**

LIST-I (Redox Reaction)	LIST-II (Type of Redox Reaction)
$A. CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	$I.$ Disproportionation reaction
$B. 2NaH(s) \rightarrow 2Na(s) + H_2(g)$	$II.$ Combination reaction
$C. V_2O_5(s) + 5Ca(s) \rightarrow 2V(s) + 5CaO(s)$	$III.$ Decomposition reaction
$D. 2H_2O(aq) \rightarrow 2H_2(g) + O_2(g)$	$IV.$ Displacement reaction

Choose the correct answer from the options given below:

a. A-IV, B-I, C-II, D-III

b. A-II, B-III, C-IV, D-I

c. A-II, B-III, C-I, D-IV

d. A-III, B-IV, C-I, D-II

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9. Which of the following oxidation reactions are carried out by both $K_2Cr_2O_7$ and $KMnO_4$ in acidic medium? A. $I^- \to I_2$ B. $S^{2-} \to S$ C. $Fe^{2+} \to Fe^{3+}$ D. $I^- \to IO_3^-$ E. $S_2O_3^{2-} \to SO_4^{2-}$ **Choose the correct answer from the options given below:** **(+4, -1)**

a. A, D and E Only

b. A, B and C Only

c. B, C and D Only

d. C, D and E Only

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10. Consider the following redox reaction: **(+4, -1)**

$$\mathrm{MnO_4^- + H^+ + H_2C_2O_4 \rightleftharpoons Mn^{2+} + H_2O + CO_2}$$

The standard reduction potentials are given as below (E_{red}):

$$E^\circ_{\mathrm{MnO_4^-/Mn^{2+}}} = +1.51\,\mathrm{V}$$

$$E^\circ_{\mathrm{CO_2/H_2C_2O_4}} = -0.49\,\mathrm{V}$$

If the equilibrium constant of the above reaction is given as $K_{\mathrm{eq}} = 10^x$, then the value of (x = _________) (nearest integer).

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11. Given below are two statements: one is labelled as Assertion (A) and the other is labelled as Reason (R). **(+4, -1)**

Assertion (A): In aqueous solutions $\mathrm{Cr^{2+}}$ is reducing while $\mathrm{Mn^{3+}}$ is oxidising in nature.

Reason (R): Extra stability to half-filled electronic configuration is observed

than incompletely filled electronic configuration.
In the light of the above statement, choose the most appropriate answer from the options given below:

a. Both (A) and (R) are true and (R) is the correct explanation of (A)

b. Both (A) and (R) are true but (R) is not the correct explanation of (A)

c. (A) is false but (R) is true

d. (A) is true but (R) is false

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12. The number of ions from the following that have the ability to liberate hydrogen from a dilute acid is _______. Ti^{2+}, Cr^{2+} and V^{2+}. **(+4, -1)**

a. 0

b. 2

c. 3

d. 1

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13. Only $2\,\mathrm{mL}$ of $\mathrm{KMnO_4}$ solution of unknown molarity is required to reach the end point of a titration of $20\,\mathrm{mL}$ of oxalic acid $(2\,\mathrm{M})$ in acidic medium. The molarity of $\mathrm{KMnO_4}$ solution should be ______ M. **(+4, -1)**

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14. When $\mathrm{MnO_2}$ and $\mathrm{H_2SO_4}$ is added to a salt (A), the greenish yellow gas liberated as salt (A) is: **(+4, -1)**

a. NaBr

b. $\mathrm{CaI_2}$

c. $\mathrm{KNO_3}$

d. $\mathrm{NH_4Cl}$

15. Thiosulphate reacts differently with iodine and bromine in the reaction given below: **(+4, -1)**

$$2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$$

$$S_2O_3^{2-} + 5Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 4Br^- + 10H^+$$

Which of the following statements justifies the above dual behaviour of thiosulphate?

a. Bromine undergoes oxidation and iodine undergoes reduction by iodine in these reactions

b. Thiosulphate undergoes oxidation by bromine and reduction by iodine in these reactions

c. Bromine is a stronger oxidant than iodine

d. Bromine is a weaker oxidant than iodine

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16. Total number of species from the following which can undergo disproportionation reaction _______. **(+4, -1)**
$H_2O_2, ClO_3^-, P_4, Cl_2, Ag, Cu^{+1}, F_2, NO_2, K^+$

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17. Number of moles of H^+ ions required by 1 mole of MnO_4^- to oxidise oxalate ion to CO_2 is ____. **(+4, -1)**

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18. Given below are two statements: **(+4, -1)**
Statement I: S_8 solid undergoes disproportionation reaction under alkaline conditions to form S^{2-} and $S_2O_3^{2-}$.
Statement II: ClO_4^- can undergo disproportionation reaction under acidic condition.
In the light of the above statements, choose the most appropriate answer from the options given below:

a. Statement I is correct but statement II is incorrect.

b. Statement I is incorrect but statement II is correct

c. Both statement I and statement II are incorrect

d. Both statement I and statement II are correct

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19. In acidic medium, $K_2Cr_2O_7$ shows oxidizing action as represented in the half-reaction: $Cr_2O_7^{2-} + XH^+ + Ye^- \rightarrow 2A + ZH_2O$ **(+4, -1)**
X, Y, Z, and A are respectively:

a. 8, 6, 4, and Cr_2O_3

b. 14, 7, 6, and Cr^{3+}

c. 8, 4, 6, and Cr_2O_3

d. 14, 6, 7, and Cr^{3+}

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20. In alkaline medium, $\mathrm{MnO_4^-}$ oxidizes $\mathrm{I^-}$ to: **(+4, -1)**

a. $\mathrm{IO_4^-}$

b. $\mathrm{IO^-}$

c. $\mathrm{I_2}$

d. $\mathrm{IO_3^-}$

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21. Chlorine undergoes disproportionation in alkaline medium as shown below: **(+4, -1)**
$a\,\mathrm{Cl_2}(g) + b\,\mathrm{OH^-(aq)} \rightarrow c\,\mathrm{ClO^-(aq)} + d\,\mathrm{Cl^-(aq)} + e\,\mathrm{H_2O(l)}$
The values of a, b, c, and d in a balanced redox reaction are respectively:

a. 1, 2, 1, and 1

b. 2, 2, 1, and 3

c. 2, 2, 1, and 3

d. 2, 2, 1, and 3

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22. 1 mole of PbS is oxidised by “X” moles of O_3>to get “Y” moles of O2. X + Y = **(+4, -1)**

____

23. $$2\mathrm{MnO_4^-} + b\mathrm{I^-} + c\mathrm{H_2O} \rightarrow x\mathrm{I_2} + y\mathrm{MnO_2} + z\mathrm{OH^-}$$ **(+4, -1)**

If the above equation is balanced with integer coefficients, the value of z is ______.

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24. Consider the reaction between PbS and HNO_3. **(+4, -1)**

PbS + HNO_3 →

Which of the following is not formed.

- **a.** NO
- **b.** NO_2
- **c.** S
- **d.** $Pb(NO_3)_2$

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25. In neutral or faintly alkaline medium, $KMnO_4$ being a powerful oxidant can oxidize, thiosulphate almost quantitatively, to sulphate. In this reaction overall change in oxidation state of manganese will be **(+4, -1)**

- **a.** 5
- **b.** 1
- **c.** 0
- **d.** 3

Answers

1. **Answer: d**

Explanation:

Let's analyze each statement to determine which one is incorrect:

1. PH_3 shows lower proton affinity than NH_3:

This statement is true. This is because phosphorus (P) has a larger atomic size than nitrogen (N), causing PH_3 to have a weaker tendency to accept a proton compared to NH_3. Thus, PH_3 has a lower proton affinity.

1. PF_3 exists but NF_5 does not:

This statement is valid. PF_3 is a stable compound. However, NF_5 does not exist because nitrogen cannot expand its octet; it only has five valence electrons available for bonding, making it unable to form NF_5.

1. NO_2 can dimerize easily:

This statement is true. Due to the presence of unpaired electrons, NO_2 can dimerize to form N_2O_4 readily.

1. SO_2 can act as an oxidizing agent, but not as a reducing agent:

This statement is incorrect. SO_2 can act as both an oxidizing agent and a reducing agent. It can be reduced to sulfur, and it can also oxidize other substances to form SO_4^{2-} ions.

The correct answer is: **"SO_2 can act as an oxidizing agent, but not as a reducing agent."**

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2. **Answer: b**

Explanation:

To determine the product (A) formed in the given reaction sequence, let's analyze each step of the transformation:

1. **Initial Reaction:** The starting compound is propyne (CH_3-C≡CH).
2. **Step 1:** The reaction with Hg^{2+}, H_2SO_4 involves oxymercuration. In the presence of these reagents, propyne undergoes hydration, which typically leads to the formation of a ketone. In this case, it converts to acetone (CH_3-C(=O)-CH_3).
3. **Step 2:** The subsequent reaction with HCN introduces a cyanohydrin formation. The compound adds across the carbonyl group of acetone forming 2-hydroxy-2-methylpropanenitrile.
4. **Step 3:** The final step with H_2/Ni involves hydrogenation. The nitrile group (-CN) is reduced to an amine group (-CH_2NH_2).

So, the final product (A) formed is 2-amino-2-methylpropanol, depicted as:

This corresponds to the correct answer option.

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3. Answer: c

Explanation:

The question asks about the relative reactivity of different carbonyl compounds in a nucleophilic addition reaction. Understanding the electronic effects and steric factors that influence the reactivity is essential in predicting the outcome.

The general order of reactivity of carbonyl compounds towards nucleophilic addition is influenced by two main factors:

- Electronic effects: Electron-withdrawing groups enhance the electrophilicity of the carbonyl carbon, thereby increasing its reactivity.
- Steric effects: Smaller groups on the carbonyl carbon make it more accessible to nucleophiles, enhancing reactivity.

Let's analyze each compound:

1. **Acetophenone**: Contains a phenyl group which is electron-donating via resonance, reducing the electrophilicity of the carbonyl carbon. Moreover, the bulkiness of the phenyl group can cause steric hindrance, further reducing reactivity.

2. **p-Tolualdehyde**: Has a methyl group which is slightly electron-donating, but since the carbonyl group is still more accessible than in acetophenone, it is more reactive.
3. **Benzaldehyde**: Lacks additional bulky substitutions and electron-donating groups compared to acetophenone, thus increasing its reactivity compared to both acetophenone and p-tolualdehyde.
4. **p-Nitrobenzaldehyde**: Contains a nitro group, which is a strong electron-withdrawing group. This significantly enhances the electrophilicity of the carbonyl carbon, making it the most reactive among the given compounds.

Therefore, considering these effects, the correct order of reactivity for nucleophilic addition reactions is:

$$\text{acetophenone} < \text{p-tolualdehyde} < \text{benzaldehyde} < \text{p-nitrobenzaldehyde}$$

This matches with the given correct option:

acetophenone $<$ p-tolualdehyde $<$ benzaldehyde $<$ p-nitrobenzaldehyde

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4. Answer: c

Explanation:

The heat of formation of $\mathrm{SO_2}$ is the heat change when 1 mole of $\mathrm{SO_2}(g)$ is formed from its elements in their standard states. By using the given reactions, the heat of formation is found to be $y - 2x$. **Final Answer:** $y - 2x\,\mathrm{kcal}$.

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5. Answer: c

Explanation:

- Statement (I) is false: The ionisation energy of Pb is lower than that of Sn because Pb is lower in the periodic table and has a higher atomic size. - Statement (II) is true: Ge has a higher ionisation energy than Si because it is in the same group but higher in the periodic table, so its electrons are closer to the nucleus. **Final Answer:** Statement I is false but Statement II is true.

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6. **Answer: c**

Explanation:

The Nernst equation is used to calculate the cell potential at non-standard conditions. The equation is given by:

$$E_{cell} = E^o_{cell} - \frac{RT}{nF} \ln \frac{[products]}{[reactants]}$$

For the given reaction, the number of electrons (n) is 2. The cell potential depends on the ratio of concentrations of the ions involved. In this case, the correct form of the Nernst equation is:

$$E_{cell} = E^o_{cell} - \frac{RT}{2F} \ln \frac{[Mg^{2+}]}{[Ag^+]}$$

Thus, the correct option is (3).

7. **Answer: a**

Explanation:

Steam volatile compounds are typically those that are capable of being vaporized or evaporated easily at relatively low temperatures. Generally, compounds with hydrogen bonding or low molecular weight are steam volatile.
Among the options:
- (A) $C_6H_4OHNO_2$: This is a nitrophenol compound, which is steam volatile due to the phenolic group that can form hydrogen bonds and is low molecular weight.
- (B) $C_6H_4NH_2NO_2$: This is a nitroaniline compound, which can also be steam volatile due to the amine group that can form hydrogen bonds.
- (C) $C_6H_4OHNH_2$: This is an amphenol compound, and although it contains an amine and phenolic group, it has a higher molecular weight and is less likely to be steam volatile.
- (D) C_6H_5OH: This is phenol, which is steam volatile but not included in the right matching with other steam volatile candidates. Thus, the correct answer is (1).

8. **Answer: c**

Explanation:

Step 1: The reaction $CH_4 + O_2$ is a combination reaction (A-II).

Step 2: The reaction $2NaH \rightarrow 2Na + H_2$ is a decomposition reaction (B-III).

Step 3: The reaction $V_2O_5 + Ca$ is a redox displacement reaction (C-I).

Step 4: The reaction $2H_2O \rightarrow 2H_2 + O_2$ is a displacement reaction (D-IV).

Final Conclusion: The correct matching is Option (3), A-II, B-III, C-I, D-IV.

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9. **Answer: a**

Explanation:

Step 1: $KMnO_4$ and $K_2Cr_2O_7$ are both strong oxidizing agents and can oxidize I^- to I_2 and $S_2O_3^{2-}$ to SO_4^{2-}. Step 2: Both can also oxidize I^- to IO_3^-, but Fe^{2+} is only oxidized by $KMnO_4$, not $K_2Cr_2O_7$. Final Conclusion: The correct answer is Option (1), A, D, and E.

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10. **Answer: 338 - 338**

Explanation:

Step-by-step Calculation:
The overall cell potential E°_{cell} for the redox reaction is calculated as:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Here:
$E^\circ_{\text{cathode}} = +1.51\,\text{V}$ (Reduction potential for MnO_4^-)
$E^\circ_{\text{anode}} = -0.49\,\text{V}$ (Reduction potential for $\text{H}_2\text{C}_2\text{O}_4$)
Therefore:

$$E^\circ_{\text{cell}} = 1.51 - (-0.49) = 2.00\,\text{V}$$

The equilibrium constant K_{eq} is related to the cell potential by the Nernst equation:

$$\Delta G^\circ = -nFE^\circ_{\text{cell}} \quad \text{and} \quad \Delta G^\circ = -RT\ln K_{\text{eq}}$$

Equating the two expressions:

$$nFE^\circ_{\text{cell}} = RT\ln K_{\text{eq}}$$

Rearranging to find K_{eq}:

$$\ln K_{\text{eq}} = \frac{nFE^\circ_{\text{cell}}}{RT}$$

Given:
$n = 5$ (number of electrons transferred)
$F = 96500\,\text{C mol}^{-1}$
$R = 8.314\,\text{J K}^{-1}\text{mol}^{-1}$
$T = 298\,\text{K}$
$E^\circ_{\text{cell}} = 2.00\,\text{V}$
Substituting the values:

$$\ln K_{\text{eq}} = \frac{5 \times 96500 \times 2.00}{8.314 \times 298}$$

$$\ln K_{\text{eq}} \approx 778.19$$

Converting to base 10:

$$\log_{10} K_{\text{eq}} = \frac{\ln K_{\text{eq}}}{\ln 10} \approx \frac{778.19}{2.303} \approx 337.78$$

Rounding to the nearest integer:

$$x \approx 338$$

Conclusion: The value of x is approximately 338 or 339.

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11. Answer: a

Explanation:

Let's analyze the assertion and reason given in the question:

1. **Assertion (A):** In aqueous solutions, Cr^{2+} is reducing while Mn^{3+} is oxidising in nature.
2. **Reason (R):** Extra stability to half-filled electronic configuration is observed than incompletely filled electronic configuration.
3. Both Cr^{2+} and Mn^{3+} are transition metal ions. Their behavior in redox reactions is influenced by their electronic configurations.
4. Cr^{2+} has the electronic configuration of $[\mathrm{Ar}]\,3d^4$. It tends to lose one electron to form the more stable $[\mathrm{Ar}]\,3d^5$ configuration, where all five 3d subshells are half-filled. Therefore, Cr^{2+} acts as a reducing agent.
5. Mn^{3+} has the electronic configuration of $[\mathrm{Ar}]\,3d^4$. By gaining an electron, it attains the stable half-filled configuration $[\mathrm{Ar}]\,3d^5$. Hence, Mn^{3+} acts as an oxidizing agent.
6. The reason provided is that the stability of the half-filled 3d subshell is significant in determining the redox behavior of these ions.

From the analysis above, we can conclude:

- Both the assertion and reason are true.
- The reason correctly explains why the assertion is true (i.e., the redox behavior of these ions is indeed due to the stability of their electronic configurations).

Therefore, the most appropriate answer is: **Both (A) and (R) are true and (R) is the correct explanation of (A)**.

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12. Answer: c

Explanation:

To determine the number of ions that can liberate hydrogen from a dilute acid, we need to analyze their reducing abilities. Strong reducing agents can donate electrons to H^+, reducing it to H_2.
Step 1: Evaluate Ti^{2+}
Titanium(II) (Ti^{2+}) has a strong tendency to get oxidized to Ti^{3+}, making it a strong reducing agent.
Ti^{2+} can react with H^+ from dilute acids to liberate H_2.
Step 2: Evaluate Cr^{2+}
Chromium(II) (Cr^{2+}) is a strong reducing agent and can be oxidized to Cr^{3+}.
Cr^{2+} reacts with H^+ from dilute acids to liberate H_2:

$$2Cr^{2+}(aq) + 2H^+(aq) \rightarrow 2Cr^{3+}(aq) + H_2(g).$$

Step 3: Evaluate V^{2+}
Vanadium(II) (V^{2+}) is also a strong reducing agent and can be oxidized to V^{3+}.
V^{2+} reacts with H^+ from dilute acids to liberate H_2.
Conclusion:
All three ions $(Ti^{2+}$, Cr^{2+}, and $V^{2+})$ can liberate H_2 from dilute acids.
Final Answer: (3).

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13. Answer: 8 - 8

Explanation:

In an acidic medium, the reaction between potassium permanganate $(KMnO_4)$ and oxalic acid $(H_2C_2O_4)$ can be represented as:

$$2KMnO_4 + 5H_2C_2O_4 + 6H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

Using the Concept of Equivalents:

According to the principle of equivalents:

$$\text{equivalents of } KMnO_4 = \text{equivalents of } H_2C_2O_4$$

Calculate Equivalents for Each Solution:

For oxalic acid $(H_2C_2O_4)$:

$$\text{Molarity} \times \text{Volume} \times \text{n-factor} = 2 \times 20 \times 2 = 80\,\text{meq}$$

where $\text{n-factor} = 2$ for oxalic acid.

For KMnO_4:

$$M \times 2 \times 5 = 10M \text{ meq}$$

where $\text{n-factor} = 5$ for KMnO_4.

Equating Equivalents:

$$10M = 80$$

Solving for M:

$$M = \frac{80}{10} = 8\,\text{M}$$

Conclusion:

The molarity of the KMnO_4 solution is $8\,\text{M}$.

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14. Answer: d

Explanation:

The reaction involved is:

$$2\text{NH}_4\text{Cl} + \text{MnO}_2 + 2\text{H}_2\text{SO}_4 \xrightarrow{\Delta} \text{MnSO}_4 + (\text{NH}_4)_2\text{SO}_4 + 2\text{H}_2\text{O} + \text{Cl}_2 \uparrow$$

In this reaction, chlorine gas (Cl_2) is liberated, which has a greenish-yellow color.

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15. Answer: c

Explanation:

The given reactions show the behavior of thiosulphate with iodine and bromine:

- With iodine: $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$
- With bromine: $S_2O_3^{2-} + 5Br_2 + 5H_2O \rightarrow 2SO_4^{2-} + 4Br^- + 10H^+$

Let's analyze the reactions:

Thiosulphate with Iodine: In the first reaction, iodine (I_2) is reduced to iodide (I^-), while thiosulphate $(S_2O_3^{2-})$ is oxidized to tetrathionate $(S_4O_6^{2-})$. This indicates that iodine acts as an oxidizing agent.

Thiosulphate with Bromine: In the second reaction, bromine (Br_2) is reduced to bromide (Br^-). Here, the thiosulphate $(S_2O_3^{2-})$ is completely oxidized to sulfate (SO_4^{2-}). This shows that bromine acts as a stronger oxidizing agent compared to iodine.

Conclusion:

The difference in behavior is due to the relative strengths of iodine and bromine as oxidizing agents. Bromine is a stronger oxidant than iodine, which is why it can oxidize thiosulphate to a greater extent, all the way to sulfate.

Based on this analysis, the correct statement that justifies the dual behavior of thiosulphate is:

> *Bromine is a stronger oxidant than iodine.*

---

16. **Answer: 6 - 6**

Explanation:

A disproportionation reaction is a type of redox reaction where a single substance is simultaneously oxidized and reduced, forming two different products. Only species in an intermediate oxidation state can undergo such reactions.

The given species are analyzed as follows:

- **H_2O_2** (Hydrogen peroxide) undergoes disproportionation to form water (H_2O) and oxygen (O_2).
- **ClO_3^-** (Chlorate ion) does not commonly undergo disproportionation.
- **P_4** (Phosphorus) can undergo disproportionation, forming phosphoric acid and phosphine in acidic conditions.

- **Cl_2** (Chlorine) can disproportionate to form chloride (Cl^-) and hypochlorite (ClO^-).
- **Ag** (Silver) in metallic form does not undergo disproportionation.
- **Cu^+** (Copper(I) ion) undergoes disproportionation to form copper(II) (Cu^{2+}) and copper metal (Cu).
- **F_2** (Fluorine) does not undergo disproportionation as it is the most electronegative element.
- **NO_2** (Nitrogen dioxide) undergoes disproportionation to form nitric acid (HNO_3) and nitric oxide (NO).
- **K^+** (Potassium ion) does not undergo disproportionation as it is already in a stable oxidation state.

From the list, the species that can undergo disproportionation reactions are:

H_2O_2, P_4, Cl_2, Cu^+, NO_2

Total number of species = 6.

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17. **Answer: 8 - 8**

Explanation:

To solve this problem, we need to determine the number of moles of H^+ ions required for the reaction where MnO_4^- oxidizes oxalate ions $(C_2O_4^{2-})$ to carbon dioxide (CO_2). The balanced redox reaction in an acidic medium is:

$$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$$

Let's break down the stoichiometric coefficients:

1. 1 mole of MnO_4^- reacts with 5/2 = 2.5 moles of $C_2O_4^{2-}$.
2. For every 2 moles of MnO_4^-, 16 moles of H^+ are needed. Therefore, for 1 mole of MnO_4^-, 16/2 = 8 moles of H^+ are required.

Hence, the number of moles of H^+ ions required by 1 mole of MnO_4^- is 8.

This result is verified as falling within the provided range of 8 to 8.

---

18. Answer: a

Explanation:

To evaluate the given statements and determine which is correct, let's analyze each statement based on known chemical reactions and principles.

1. **Statement I:** S_8 solid undergoes disproportionation reaction under alkaline conditions to form S^{2-} and $S_2O_3^{2-}$.
    - Disproportionation is a specific type of redox (reduction-oxidation) reaction in which a single substance is simultaneously reduced and oxidized to form two different products.
    - In alkaline conditions, elemental sulfur (S_8) can actually disproportionate. The reaction is:

$$3S_8 + 6OH^- \rightarrow 2S_2O_3^{2-} + 4S^{2-} + 3H_2O$$

- This confirms that under alkaline conditions, sulfur can indeed form S^{2-} (sulfide) and $S_2O_3^{2-}$ (thiosulfate).

1. **Statement II:** \(ClO_4^-\\) can undergo disproportionation reaction under acidic conditions.
    - Perchlorate ion (ClO_4^-) is known for being very stable, especially under acidic conditions. It is already in its highest oxidation state, where chlorine is +7.
    - Generally, in a disproportionation reaction, a substance must be capable of existing in a higher as well as a lower oxidation state. However, chlorine in ClO_4^- cannot oxidize further (higher than +7 is unfeasible under normal conditions).
    - Therefore, no disproportionation reaction of ClO_4^- occurs under acidic conditions, confirming that Statement II is incorrect.

Based on the analysis above:

- Statement I is correct as sulfur in S_8 can undergo disproportionation to form S^{2-} and $S_2O_3^{2-}$ under alkaline conditions.
- Statement II is incorrect as \(ClO_4^-\\) does not undergo disproportionation under acidic conditions.

Therefore, the correct answer is: Statement I is correct but statement II is incorrect.

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19. Answer: d

Explanation:

The problem asks to balance the given half-reaction for the reduction of dichromate ion $(K_2Cr_2O_7)$ in an acidic medium and identify the unknown coefficients X, Y, Z and the product species A.

Concept Used:

The balancing of a redox half-reaction in an acidic medium follows a set of sequential rules:

a. Balance all atoms other than oxygen (O) and hydrogen (H).
b. Balance the oxygen atoms by adding water (H_2O) molecules to the side deficient in oxygen.
c. Balance the hydrogen atoms by adding hydrogen ions (H^+) to the side deficient in hydrogen.
d. Balance the net charge on both sides by adding electrons (e^-) to the side with the greater positive charge.

We also need to know the product of the reduction of dichromate. In an acidic medium, the dichromate ion $(Cr_2O_7^{2-})$, where chromium is in the +6 oxidation state, is reduced to the chromium(III) ion (Cr^{3+}).

Step-by-Step Solution:

Step 1: Identify the species A and write the skeleton equation.

In an acidic medium, the dichromate ion $(Cr_2O_7^{2-})$ is a strong oxidizing agent and gets reduced to the chromium(III) ion (Cr^{3+}). Therefore, the species A is Cr^{3+}. The skeleton half-reaction is:

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$$

The chromium atoms are already balanced, with 2 on each side.

Step 2: Balance the oxygen atoms.

There are 7 oxygen atoms on the left-hand side (LHS) and none on the right-hand side (RHS). To balance the oxygen atoms, we add 7 water molecules (H_2O) to the RHS.

$$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$$

Comparing this with the given format $\cdots \rightarrow 2A + ZH_2O$, we can identify $Z = 7$.

Step 3: Balance the hydrogen atoms.

Now, there are $7 \times 2 = 14$ hydrogen atoms on the RHS and none on the LHS. To balance the hydrogen atoms in an acidic medium, we add 14 hydrogen ions (H^+) to the LHS.

$$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$$

Comparing this with the given format $Cr_2O_7^{2-} + XH^+ + \ldots$, we can identify $X = 14$.

Step 4: Balance the charge.

Calculate the total charge on both sides of the equation.

- Total charge on LHS = $(1 \times -2) + (14 \times +1) = -2 + 14 = +12$
- Total charge on RHS = $(2 \times +3) + (7 \times 0) = +6$

The charge is not balanced. To balance the charge, we add electrons (e^-) to the side with the higher (more positive) charge. We need to add $12 - 6 = 6$ electrons to the LHS.

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

Comparing this with the given format $Cr_2O_7^{2-} + XH^+ + Ye^- \rightarrow \ldots$, we can identify $Y = 6$.

Final Result:

The completely balanced half-reaction is:

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$

By comparing this with the given equation $Cr_2O_7^{2-} + XH^+ + Ye^- \rightarrow 2A + ZH_2O$, we find:

- $X = 14$

- $Y = 6$
- $A = Cr^{3+}$
- $Z = 7$

Therefore, X, Y, Z, and A are respectively: **14, 6, 7, and** Cr^{3+}.

---

20. Answer: d

Explanation:

To understand this question, we need to look at the redox reaction between permanganate ion $\mathrm{MnO_4^-}$ and iodide ion $\mathrm{I^-}$ in an alkaline medium.

The permanganate ion acts as a strong oxidizing agent. In an alkaline medium, it can oxidize iodide ions to iodate ions according to the balanced chemical equation:

$$2\mathrm{MnO_4^-} + \mathrm{I^-} + \mathrm{H_2O} \rightarrow 2\mathrm{MnO_2} + \mathrm{IO_3^-} + 2\mathrm{OH^-}$$

This reaction shows that in an alkaline medium, $\mathrm{I^-}$ is oxidized to $\mathrm{IO_3^-}$ (iodate ion) when reacted with permanganate ion. The presence of hydroxide ions from the alkaline medium helps stabilize the formation of manganese dioxide $(\mathrm{MnO_2})$.

Let's analyze the options:

1. $\mathrm{IO_4^-}$ - Periodate ion, not typically formed in alkaline conditions with this oxidation.
2. $\mathrm{IO^-}$ - Hypoiodite ion, not the product in this specific reaction.
3. $\mathrm{I_2}$ - Molecular iodine, not a product of this reaction in alkaline medium.
4. $\mathrm{IO_3^-}$ - Iodate ion, correct product of this reaction.

Thus, the correct answer is that in an alkaline medium, $\mathrm{MnO_4^-}$ oxidizes $\mathrm{I^-}$ to $\mathrm{IO_3^-}$.

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21. Answer: a

Explanation:

To find the values of a, b, c, and d in the balanced equation of chlorine disproportionation in an alkaline medium, let's follow these steps:

1. Disproportionation involves the simultaneous oxidation and reduction of the same element. In this case, chlorine $(\mathrm{Cl_2})$ is both reduced to chloride ions $(\mathrm{Cl^-})$ and oxidized to hypochlorite ions $(\mathrm{ClO^-})$.
2. First, write the half-reactions for oxidation and reduction:

- Reduction: $\mathrm{Cl_2} + 2e^- \rightarrow 2\mathrm{Cl^-}$
- Oxidation: $\mathrm{Cl_2} + 2\mathrm{OH^-} \rightarrow 2\mathrm{ClO^-} + 2\mathrm{H_2O} + 2e^-$

1. Combine the half-reactions. Make sure the number of electrons lost is equal to the number of electrons gained to maintain electron balance:

- Overall equation after canceling electrons: $\mathrm{Cl_2} + 2\mathrm{OH^-} \rightarrow \mathrm{ClO^-} + \mathrm{Cl^-} + \mathrm{H_2O}$

1. Balancing the atoms and charges, we have:

From the combined reaction:

- The coefficient of $\mathrm{Cl_2}$ (a) is 1.
- The coefficient of $\mathrm{OH^-}$ (b) is 2.
- The coefficient of $\mathrm{ClO^-}$ (c) is 1.
- The coefficient of $\mathrm{Cl^-}$ (d) is 1.

1. Thus, the balanced redox reaction coefficients are $a = 1$, $b = 2$, $c = 1$, and $d = 1$.

Therefore, the correct set of coefficients is 1, 2, 1, and 1, which matches with the given option: **1, 2, 1, and 1**.

---

22. Answer: 8 - 8

Explanation:

To solve this problem, we need to determine the number of moles of O_3 and O_2 involved when 1 mole of PbS is oxidized. The chemical reaction for oxidizing PbS using O_3 is:

$\mathrm{PbS} + 4\mathrm{O_3} \rightarrow \mathrm{PbSO_4} + 4\mathrm{O_2}$

Step-by-step explanation:

From the balanced reaction, 1 mole of PbS requires 4 moles of O_3 to be oxidized to $\mathrm{PbSO_4}$.

During this process, it produces 4 moles of O_2.

Given these stoichiometric relationships, the number of moles of O_3 (“X” moles) used is 4, and the number of moles of O_2 (“Y” moles) produced is also 4.

The problem asks for the sum X + Y, which is:

X + Y = 4 + 4 = 8

Verification:

The computed value of X + Y = 8 falls within the expected range of [8,8].

---

23. **Answer: 8 - 8**

Explanation:

Reduction Half Reaction:

$$2\mathrm{MnO}_4^- \rightarrow 2\mathrm{MnO}_2$$

$$2\mathrm{MnO}_4^- + 4\mathrm{H}_2\mathrm{O} + 6e^- \rightarrow 2\mathrm{MnO}_2 + 8\mathrm{OH}^-$$

Oxidation Half Reaction:

$$2\mathrm{I}^- \rightarrow \mathrm{I}_2 + 2e^-$$

$$6\mathrm{I}^- \rightarrow 3\mathrm{I}_2 + 6e^-$$

Adding the oxidation half and reduction half, we get the net reaction as:

$$2\mathrm{MnO}_4^- + 6\mathrm{I}^- + 4\mathrm{H}_2\mathrm{O} \rightarrow 3\mathrm{I}_2 + 2\mathrm{MnO}_2 + 8\mathrm{OH}^-$$

Thus, $z = 8$.

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24. **Answer: b**

Explanation:

The Correct answer is option is (B) : NO_2

25. **Answer: d**

Explanation:

The question involves determining the change in the oxidation state of manganese in the reaction of potassium permanganate ($KMnO_4$) with thiosulfate in a neutral or faintly alkaline medium. This can be solved by understanding the oxidation-reduction behavior of manganese in this chemical reaction.

In a neutral or faintly alkaline medium, potassium permanganate ($KMnO_4$) acts as a strong oxidizing agent. The oxidation state of manganese in $KMnO_4$ is +7. During the reaction with thiosulfate, manganese is reduced. We need to determine the oxidation state to which it is reduced.

The overall balanced reaction in a neutral or faintly alkaline medium is: $2\,\mathrm{KMnO_4} + 3\,\mathrm{Na_2S_2O_3} + 2\,\mathrm{H_2O} \rightarrow 2\,\mathrm{MnO_2} + 3\,\mathrm{Na_2SO_4} + 2\,\mathrm{KOH}$

In the above reaction, manganese goes from an oxidation state of +7 in $KMnO_4$ to an oxidation state of +4 in MnO_2. Thus, the change in oxidation state of manganese is: $+7 - (+4) = +3$.

The overall change in the oxidation state of manganese in this reaction is **3**.

Hence, the correct answer is **3**.

Concepts:

1. **Redox Reactions:**

Redox Reaction:

Redox reactions are chemical reactions where oxidation and reduction take place simultaneously. In this type of reaction, there is a gain of electrons for one chemical species while the other loses electrons or simply involves transfer of electrons. The species that loses electrons is oxidized while the one that gains electrons is reduced.

Types of Redox Reactions:

Redox reactions can be differentiated into 4 categories namely combination reactions, decomposition reactions, displacement reactions, and disproportionation reactions. Each is explained separately below:

Combination Reaction:

In this, the molecules combine to form new compounds. For example, when magnesium reacts to nitrogen.

Decomposition Reaction:

Opposite to the combination reaction, here there is a breakdown of compounds to simpler substances. For example, electrolysis of water.

Displacement Reaction:

In this, the more reactive metal will displace the less reactive one in a chemical reaction. The reactivity of an element is represented in a series called the reactivity series (arranged in decreasing order of reactivity) which makes it easier to determine the chemical reaction and its products.

Disproportionation Reaction:

This is a peculiar type of reaction where an element showing a particular oxidation state will be oxidized and reduced simultaneously. Another thing to note is that these reactions will always have an element that can exhibit three oxidation states.