KCET 2026 Chemistry

Question Paper with Solutions

Conducted by KEA

General Instructions

(i) **Duration:** The total duration of the examination is 80 minutes.

(ii) **Total Marks:** The complete paper carries a maximum of 60 marks.

(iii) **Compulsory Questions:** All 60 questions are compulsory.

(iv) Each question has four options. Only **one** option is correct.

(v) Correct Answer: +1 marks.

(vi) **Incorrect Answer:** There is no Negative marking for incorrect answers.

1. During the electrolysis of acidified water, 16 g of O_2 gas is formed. The volume of H_2 gas liberated at cathode under STP conditions is

(1) 22.4 L

(2) 11.2 L

(3) 2.24 L

(4) 1.12 L

Correct Answer: (1) 22.4 L

Solution:

Step 1: Understanding the Question:

The question asks to determine the volume of hydrogen gas (H_2) liberated at the cathode during the electrolysis of acidified water, given the mass of oxygen gas (O_2) formed at the anode. The conditions are STP (Standard Temperature and Pressure).

Step 2: Key Formula or Approach:

1. Write the balanced chemical equation for the electrolysis of water to determine the stoichiometric ratio between O_2 and H_2.
2. Calculate the moles of O_2 from its given mass.
3. Use the stoichiometric ratio to find the moles of H_2.
4. At STP, 1 mole of any ideal gas occupies 22.4 L. Use this to find the volume of H_2.

Step 3: Detailed Explanation:

The overall reaction for the electrolysis of acidified water is:

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

From the balanced equation, 2 moles of H_2 gas are produced for every 1 mole of O_2 gas formed.

Given:
Mass of $O_2 = 16$ g
Molar mass of $O_2 = 2 \times 16.0$ g/mol $= 32.0$ g/mol

Calculate moles of O_2:

$$\text{Moles of O}_2 = \frac{\text{Mass of O}_2}{\text{Molar mass of O}_2} = \frac{16\text{ g}}{32\text{ g/mol}} = 0.5\text{ mol}$$

Using the stoichiometric ratio (2 moles H_2 : 1 mole O_2):

$$\text{Moles of H}_2 = 2 \times \text{Moles of O}_2 = 2 \times 0.5\text{ mol} = 1.0\text{ mol}$$

At STP, 1 mole of any ideal gas occupies 22.4 L.

$$\text{Volume of H}_2 = \text{Moles of H}_2 \times 22.4\text{ L/mol} = 1.0\text{ mol} \times 22.4\text{ L/mol} = 22.4\text{ L}$$

Step 4: Final Answer:

The volume of H_2 gas liberated at the cathode under STP conditions is 22.4 L.

Quick Tip: Always remember the stoichiometric ratio for the electrolysis of water ($2H_2 : 1O_2$) and the molar volume of a gas at STP (22.4 L/mol). These are fundamental constants in electrochemistry and gas laws.

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2. $\Lambda_m^\circ(NH_4OH)$ **is equal to ____**

(1) $\Lambda_m^\circ(NH_4OH) + \Lambda_m^\circ(NH_4Cl) - \Lambda_m^\circ(HCl)$

(2) $\Lambda_m^\circ(NH_4Cl) + \Lambda_m^\circ(NaOH) - \Lambda_m^\circ(NaCl)$

(3) $\Lambda_m^\circ(NH_4Cl) + \Lambda_m^\circ(NaCl) - \Lambda_m^\circ(NaOH)$

(4) $\Lambda_m^\circ(NaOH) + \Lambda_m^\circ(NaCl) - \Lambda_m^\circ(NH_4Cl)$

Correct Answer: (2) $\Lambda_m^\circ(NH_4Cl) + \Lambda_m^\circ(NaOH) - \Lambda_m^\circ(NaCl)$

Solution:

Step 1: Understanding the Question:

The question asks to express the limiting molar conductivity of a weak base, ammonium hydroxide ($\Lambda_m^\circ(NH_4OH)$), using Kohlrausch's Law of Independent Migration of Ions. This law allows calculating the limiting molar conductivity of a weak electrolyte from those of strong electrolytes.

Step 2: Key Formula or Approach:

Kohlrausch's Law states that Λ_m° of an electrolyte can be written as the sum of the limiting molar conductivities of its constituent ions:

$$\Lambda_m^\circ(A_xB_y) = x\lambda_{A^{y+}}^\circ + y\lambda_{B^{x-}}^\circ$$

For weak electrolytes like NH_4OH, which do not dissociate completely, Λ_m° cannot be directly measured by extrapolation. Instead, it is calculated by combining Λ_m° values of strong electrolytes that share the relevant ions.

We need to find a combination that yields $\lambda_{NH_4^+}^\circ + \lambda_{OH^-}^\circ$.

Step 3: Detailed Explanation:

Consider the following strong electrolytes:

1. Ammonium chloride (NH_4Cl, a strong salt):

$$\Lambda_m^\circ(NH_4Cl) = \lambda_{NH_4^+}^\circ + \lambda_{Cl^-}^\circ$$

2. Sodium hydroxide ($NaOH$, a strong base):

$$\Lambda_m^\circ(NaOH) = \lambda_{Na^+}^\circ + \lambda_{OH^-}^\circ$$

3. Sodium chloride ($NaCl$, a strong salt):

$$\Lambda_m^\circ(NaCl) = \lambda_{Na^+}^\circ + \lambda_{Cl^-}^\circ$$

To obtain $\Lambda_m^\circ(NH_4OH) = \lambda_{NH_4^+}^\circ + \lambda_{OH^-}^\circ$, we can sum the conductivities of NH_4Cl and $NaOH$ and then subtract the conductivity of $NaCl$:

$$(\lambda_{NH_4^+}^\circ + \lambda_{Cl^-}^\circ) + (\lambda_{Na^+}^\circ + \lambda_{OH^-}^\circ) - (\lambda_{Na^+}^\circ + \lambda_{Cl^-}^\circ)$$

$$= \lambda_{NH_4^+}^\circ + \lambda_{OH^-}^\circ$$

Thus,

$$\Lambda_m^\circ(NH_4OH) = \Lambda_m^\circ(NH_4Cl) + \Lambda_m^\circ(NaOH) - \Lambda_m^\circ(NaCl)$$

Step 4: Final Answer:

The correct expression for $\Lambda_m^\circ(NH_4OH)$ is given by option (2).

Quick Tip: To construct the Λ_m° for a weak electrolyte (e.g., AB), you generally need Λ_m° of its strong salt (AX), a strong base containing its cation (BX), and a strong acid containing its anion (BY), then $\Lambda_m^\circ(AB) = \Lambda_m^\circ(AX) + \Lambda_m^\circ(BY) - \Lambda_m^\circ(XY)$. In this case, we have NH_4OH, so we use NH_4Cl (strong salt), $NaOH$ (strong base), and $NaCl$ (strong salt from remaining ions).

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3. Given below are the half-cell reactions:

$Mn^{2+} + 2e^- \rightarrow Mn$ **(E° = -1.18 V)**

$Mn^{3+} + e^- \rightarrow Mn^{2+}$ **(E° = +1.51 V)**

The E°_{cell} for $3Mn^{2+} \rightarrow Mn + 2Mn^{3+}$ **will be ____**

(1) - 2.69 V, the reaction will not occur (Non-Spontaneous)

(2) - 2.69 V, the reaction will occur (Spontaneous)

(3) - 0.33 V, the reaction will not occur (Non-Spontaneous)

(4) - 0.33 V, the reaction will occur (Spontaneous)

Correct Answer: (1) - 2.69 V, the reaction will not occur (Non-Spontaneous)

Solution:

Step 1: Understanding the Question:

The question asks to calculate the standard cell potential (E°_{cell}) for a given redox reaction involving manganese ions and determine its spontaneity. We are provided with the standard reduction potentials for two relevant half-reactions.

Step 2: Key Formula or Approach:

1. Identify the oxidation and reduction half-reactions in the target overall reaction.

2. Assign the standard reduction potentials to the appropriate cathode and anode half-reactions. If a given half-reaction needs to be reversed for the target reaction, its potential's sign must be flipped (though it's often safer to use ΔG° for non-standard combinations or to consistently assign E°_{red} for cathode and E°_{ox} for anode).

3. Calculate $E^\circ_{cell} = E^\circ_{\text{cathode}} + E^\circ_{\text{anode}}$ (where E°_{anode} is the standard oxidation potential). Or, $E^\circ_{cell} = E^\circ_{\text{red,cathode}} - E^\circ_{\text{red,anode}}$.

4. Determine spontaneity: If $E^\circ_{cell} > 0$, the reaction is spontaneous. If $E^\circ_{cell} < 0$, it is non-spontaneous.

Step 3: Detailed Explanation:

Target reaction: $3Mn^{2+} \rightarrow Mn + 2Mn^{3+}$

This is a disproportionation reaction where Mn^{2+} is simultaneously reduced to Mn and oxidized to Mn^{3+}.

Let's break down the overall reaction into half-reactions:

1. **Reduction part:** $Mn^{2+} + 2e^- \rightarrow Mn$

The standard reduction potential for this is given as $E^\circ_{\text{red},1} = -1.18$ V.

2. **Oxidation part:** $Mn^{2+} \rightarrow Mn^{3+} + e^-$

The given reaction is $Mn^{3+} + e^- \rightarrow Mn^{2+}$, with $E^\circ_{\text{red},2} = +1.51$ V.

To get the oxidation $Mn^{2+} \rightarrow Mn^{3+} + e^-$, we must reverse the second given half-reaction. The standard oxidation potential for this reaction will be the negative of its standard reduction potential: $E^\circ_{\text{ox}} = -E^\circ_{\text{red},2} = -1.51$ V.

Now, combine these two half-reactions to match the overall stoichiometry. The overall reaction needs 2 moles of Mn^{3+} formed, so the oxidation half-reaction must occur twice, leading to a transfer of 2 electrons. The reduction half-reaction involves 2 electrons.

So, the balanced half-reactions are:

Cathode (Reduction): $Mn^{2+} + 2e^- \rightarrow Mn \quad (E^\circ_{\text{red}} = -1.18 \text{ V})$

Anode (Oxidation): $2Mn^{2+} \rightarrow 2Mn^{3+} + 2e^- \quad (E^\circ_{\text{ox}} = -1.51 \text{ V})$

(Note: E° values are intensive properties and do not change when the stoichiometric coefficients are multiplied.)

Now, calculate E°_{cell}:

$$E^\circ_{cell} = E^\circ_{\text{red}} + E^\circ_{\text{ox}}$$

$$E^\circ_{cell} = (-1.18 \text{ V}) + (-1.51 \text{ V})$$

$$E^\circ_{cell} = -2.69 \text{ V}$$

Spontaneity:

Since E°_{cell} is negative (-2.69 V), the Gibbs free energy change ($\Delta G^\circ = -nFE^\circ_{cell}$) will be positive. A positive ΔG° indicates a non-spontaneous reaction under standard conditions.

Step 4: Final Answer:

The E°_{cell} for the reaction is -2.69 V, and the reaction will not occur (Non-Spontaneous). This matches option (1).

Quick Tip: Always ensure consistency when adding potentials: either sum reduction potential of cathode and oxidation potential of anode, or subtract reduction potential of anode from reduction potential of cathode. For spontaneity, a positive E°_{cell} (or negative ΔG°) is required.

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4. The conductivity of centimolar solution of KCl at 298 K is 0.021 Ω^{-1} cm^{-1}. The resistance of the cell containing the solution at 298 K is 60 Ω. The value of cell constant is

(1) 3.28 cm^{-1}

(2) 1.26 cm^{-1}

(3) 3.54 cm^{-1}

(4) 1.34 cm^{-1}

Correct Answer: (2) 1.26 cm^{-1}

Solution:

Step 1: Understanding the Question:

The question provides the conductivity (κ) and resistance (R) of a centimolar KCl solution in a conductivity cell at a specific temperature. We need to calculate the cell constant (G^*) of the conductivity cell.

Step 2: Key Formula or Approach:

The relationship between conductivity, resistance, and cell constant is given by:

$$\kappa = \frac{1}{R} \times G^*$$

Where:

- κ is the conductivity (units: Ω^{-1} cm^{-1} or S cm^{-1}).
- R is the resistance (units: Ω).
- G^* is the cell constant (units: cm^{-1}).

Rearranging the formula to solve for the cell constant:

$$G^* = \kappa \times R$$

Step 3: Detailed Explanation:

Given values:

- Conductivity, $\kappa = 0.021\ \Omega^{-1}\ \text{cm}^{-1}$
- Resistance, $R = 60\ \Omega$

Substitute these values into the formula for cell constant:

$$G^* = (0.021\ \Omega^{-1}\ \text{cm}^{-1}) \times (60\ \Omega)$$

$$G^* = 1.26\ \text{cm}^{-1}$$

The concentration ("centimolar") and temperature (298 K) are relevant for conductivity but not directly used in calculating the cell constant, as the cell constant is a geometric property of the cell itself, independent of the solution.

Step 4: Final Answer:

The value of the cell constant is 1.26 cm^{-1}.

Quick Tip: The cell constant is a physical property of the conductivity cell and remains constant for a given cell, regardless of the solution placed inside it or the temperature. It is often determined using a solution of known conductivity (like KCl).

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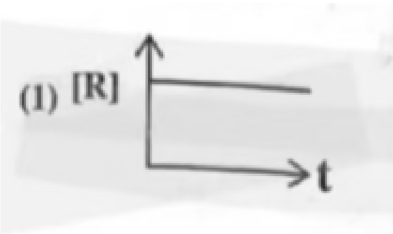
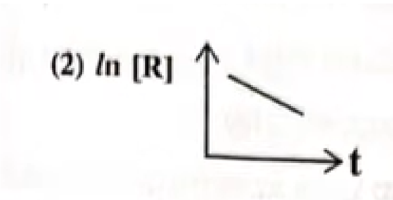
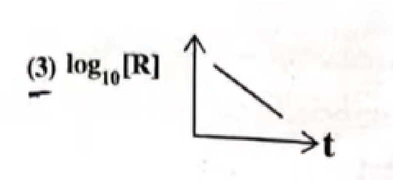
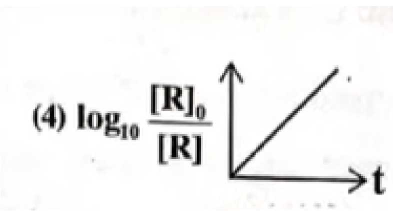
5 Which one of the following graph is not applicable for a 1st order reaction ($R \rightarrow P$)?









(1) Graph 1

(2) Graph 2

(3) Graph 3

(4) Graph 4

Correct Answer: (1) Graph 1

Solution:

Step 1: Understanding the Question:

The question asks to identify the graph that does NOT represent the kinetic behavior of a first-order reaction.

Step 2: Key Formula or Approach:

For a first-order reaction $R \rightarrow P$, the integrated rate law is:

$$\ln[R]_t - \ln[R]_0 = -kt$$

This can be rewritten in various forms for graphical analysis.

Step 3: Detailed Explanation:

Let's analyze each type of graph for a first-order reaction:

1. **Graph of** $\ln[R]$ **vs** t**:**

The integrated rate law is $\ln[R]_t = -kt + \ln[R]_0$.

This is in the form $y = mx + c$, where $y = \ln[R]_t$, $m = -k$ (a negative slope), and $c = \ln[R]_0$.

Therefore, a plot of $\ln[R]$ vs t is a straight line with a negative slope. This matches Graph (2) and is applicable.

2. **Graph of** $\log_{10}[R]$ **vs** t**:**

Dividing the integrated rate law by 2.303 gives $\log_{10}[R]_t = -\frac{k}{2.303}t + \log_{10}[R]_0$.

This is also a straight line with a negative slope. This matches Graph (3) and is applicable.

3. **Graph of** $\log_{10}\left(\frac{[R]_0}{[R]}\right)$ **vs** t**:**

From $\ln\left(\frac{[R]_0}{[R]_t}\right) = kt$, we get $\log_{10}\left(\frac{[R]_0}{[R]_t}\right) = \frac{k}{2.303}t$.

This is a straight line passing through the origin with a positive slope. This matches Graph (4) and is applicable.

4. **Graph of** $[R]$ **vs** t**:**

The concentration of reactant as a function of time is given by $[R]_t = [R]_0 e^{-kt}$.

This represents an exponential decay, meaning the concentration of reactant decreases over time, but not linearly. It is never a constant value unless $k = 0$ (no reaction) or $t = 0$.

Graph (1) shows $[R]$ as a horizontal line, implying constant concentration over time. This is incorrect for a reaction where reactant is consumed.

Step 4: Final Answer:

Graph (1) is not applicable for a 1st order reaction.

Quick Tip: Always remember the characteristic plots for different reaction orders.

For 1st order: $\ln[R]$ vs t is linear (negative slope).

For 0th order: $[R]$ vs t is linear (negative slope).

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6. For a reaction having three steps, the overall rate constant is $K = \frac{k_1 k_2}{k_3}$. The values E$_{a1}$, E$_{a2}$ and E$_{a3}$ (activation energies stepwise) are 40, 50 and 60 kJ mol^{-1} respectively. Then the overall Ea (activation energy) of the reaction is ____.

(1) 30 kJ mol^{-1}

(2) 40 kJ mol^{-1}

(3) 50 kJ mol^{-1}

(4) 60 kJ mol^{-1}

Correct Answer: (1) 30 kJ mol^{-1}

Solution:

Step 1: Understanding the Question:

The question asks to calculate the overall activation energy for a multi-step reaction, given the expression for its overall rate constant in terms of individual step rate constants and their respective activation energies.

Step 2: Key Formula or Approach:

The Arrhenius equation relates the rate constant k to the activation energy E_a:

$$k = Ae^{-E_a/RT}$$

Taking the natural logarithm of both sides:

$$\ln k = \ln A - \frac{E_a}{RT}$$

For an overall reaction, if its rate constant K is expressed as a product or quotient of individual step rate constants, the overall activation energy E_a can be found by summing and subtracting the individual activation energies.

Step 3: Detailed Explanation:

Given the overall rate constant expression:

$$K = \frac{k_1 k_2}{k_3}$$

Take the natural logarithm of both sides:

$$\ln K = \ln(k_1 k_2) - \ln k_3$$

$$\ln K = \ln k_1 + \ln k_2 - \ln k_3$$

Now, substitute the Arrhenius equation (in logarithmic form) for each rate constant:

$$\left(\ln A_{\text{overall}} - \frac{E_{a,\text{overall}}}{RT}\right) = \left(\ln A_1 - \frac{E_{a1}}{RT}\right) + \left(\ln A_2 - \frac{E_{a2}}{RT}\right) - \left(\ln A_3 - \frac{E_{a3}}{RT}\right)$$

Equating the terms containing the activation energies (which depend on $-1/RT$):

$$-\frac{E_{a,\text{overall}}}{RT} = -\frac{E_{a1}}{RT} - \frac{E_{a2}}{RT} + \frac{E_{a3}}{RT}$$

Multiply by $-RT$ to solve for the overall activation energy:

$$E_{a,\text{overall}} = E_{a1} + E_{a2} - E_{a3}$$

Given values:

$E_{a1} = 40 \text{ kJ mol}^{-1}$

$E_{a2} = 50 \text{ kJ mol}^{-1}$

$E_{a3} = 60 \text{ kJ mol}^{-1}$

Substitute these values:

$$E_{a,\text{overall}} = 40 \text{ kJ mol}^{-1} + 50 \text{ kJ mol}^{-1} - 60 \text{ kJ mol}^{-1}$$

$$E_{a,\text{overall}} = 90 \text{ kJ mol}^{-1} - 60 \text{ kJ mol}^{-1} = 30 \text{ kJ mol}^{-1}$$

Step 4: Final Answer:

The overall activation energy of the reaction is 30 kJ mol^{-1}.

Quick Tip: For overall rate constants that are products or quotients of individual rate constants (e.g., $K = \frac{k_1 k_2}{k_3}$), the overall activation energy E_a is found by summing the E_a values for terms in the numerator and subtracting those in the denominator. That is, $E_a = \sum E_a(\text{numerator}) - \sum E_a(\text{denominator})$.

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7. For a 1st order change R → P, the concentration of Reactant R changes from 0.1 M to 0.025 M in 40 minutes. The rate of reaction when the concentration of R is 0.01 M is ____.

(1) 1.73×10^{-5} M min^{-1}

(2) 3.47×10^{-4} M min^{-1}

(3) 3.47×10^{-5} M min^{-1}

(4) 1.73×10^{-4} M min^{-1}

Correct Answer: (2) 3.47×10^{-4} M min^{-1}

Solution:

Step 1: Understanding the Question:

The question asks to calculate the instantaneous rate of a first-order reaction at a specific reactant concentration, given the initial and final concentrations over a time interval.

Step 2: Key Formula or Approach:

1. **Integrated Rate Law for First-Order Reaction:**

$$k = \frac{2.303}{t}\log_{10}\left(\frac{[R]_0}{[R]_t}\right)$$

Where:

- k is the rate constant.
- t is the time.
- $[R]_0$ is the initial concentration of reactant.
- $[R]_t$ is the concentration of reactant at time t.

2. **Rate Law for First-Order Reaction:**

$$\text{Rate} = k[R]$$

Step 3: Detailed Explanation:

Part 1: Calculate the rate constant (k)

Given:

Initial concentration, $[R]_0 = 0.1\ \text{M}$

Concentration at time t, $[R]_t = 0.025\ \text{M}$

Time, $t = 40$ minutes

Substitute these values into the integrated rate law:

$$k = \frac{2.303}{40\ \text{min}}\log_{10}\left(\frac{0.1\ \text{M}}{0.025\ \text{M}}\right)$$

$$k = \frac{2.303}{40}\log_{10}(4)$$

Using $\log_{10}(4) \approx 0.6021$:

$$k = \frac{2.303}{40} \times 0.6021$$

$$k \approx 0.057575 \times 0.6021$$

$$k \approx 0.03467 \text{ min}^{-1}$$

Rounding to a reasonable number of significant figures, $k \approx 0.0347 \text{ min}^{-1}$.

Part 2: Calculate the rate of reaction when $[R] = 0.01$ M

Using the rate law for a first-order reaction:

$$\text{Rate} = k[R]$$

Substitute the calculated k and the given concentration:

$$\text{Rate} = (0.0347 \text{ min}^{-1}) \times (0.01 \text{ M})$$

$$\text{Rate} = 0.000347 \text{ M min}^{-1}$$

Expressing in scientific notation:

$$\text{Rate} = 3.47 \times 10^{-4} \text{ M min}^{-1}$$

Step 4: Final Answer:

The rate of reaction when the concentration of R is 0.01 M is 3.47×10^{-4} M min^{-1}.

Quick Tip: Always calculate the rate constant k first using the given concentration change over time. Then use this k with the specific concentration for which the rate is requested. Pay attention to units (M, min^{-1}).

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8. The activation energy for the reaction X → Y is 150 kJ mol^{-1}. The change in enthalpy for the above reaction is -135 kJ mol^{-1}. Then the activation energy for Y → X is ____.

(1) 280 kJ mol^{-1}

(2) 285 kJ mol^{-1}

(3) 270 kJ mol^{-1}

(4) 15 kJ mol^{-1}

Correct Answer: (2) 285 kJ mol^{-1}

Solution:

Step 1: Understanding the Question:

The question asks to determine the activation energy for the reverse reaction (Y → X), given the activation energy for the forward reaction (X → Y) and the enthalpy change for the forward reaction.

Step 2: Key Formula or Approach:

The relationship between the activation energy of the forward reaction ($E_{a,f}$), the activation energy of the reverse reaction ($E_{a,b}$), and the enthalpy change of the reaction (ΔH) is:

$$\Delta H = E_{a,f} - E_{a,b}$$

This relationship can be visualized using a reaction coordinate diagram. For an exothermic reaction ($\Delta H < 0$), the products are at a lower energy level than the reactants.

Step 3: Detailed Explanation:

Given values for the forward reaction $X \rightarrow Y$:

Activation energy for forward reaction, $E_{a,f} = 150 \text{ kJ mol}^{-1}$

Change in enthalpy, $\Delta H = -135 \text{ kJ mol}^{-1}$

We need to find the activation energy for the reverse reaction $Y \rightarrow X$, which is $E_{a,b}$.
Substitute the given values into the formula:

$$-135 \text{ kJ mol}^{-1} = 150 \text{ kJ mol}^{-1} - E_{a,b}$$

Now, solve for $E_{a,b}$:

$$E_{a,b} = 150 \text{ kJ mol}^{-1} + 135 \text{ kJ mol}^{-1}$$

$$E_{a,b} = 285 \text{ kJ mol}^{-1}$$

Step 4: Final Answer:

The activation energy for the reverse reaction Y → X is 285 kJ mol^{-1}.

Quick Tip: Always draw a simple reaction coordinate diagram to visualize the relationship. For an exothermic reaction ($\Delta H < 0$), the products are lower in energy than reactants. The energy barrier from products to reactants ($E_{a,b}$) will always be larger than the barrier from reactants to products ($E_{a,f}$) if ΔH is negative.
