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CHEMISTRY WORKSHEET

REACTION MECHANISMS

Course: SCH4U Grade 12 Chemistry

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50 Practice Problems with Answers and Detailed Solutions

FULL worksheet contains 50 representative problems on: a

✓ Elementary steps, molecularity, catalysts, intermediates, ✓ Rate-determining steps and mechanism-based rate laws, ✓ Pre-equilibrium and advanced mechanism reasoning,

All problems are designed to reflect style and difficulty commonly seen on quizzes, tests, and final exams. Detailed solutions are included to help students understand the reasoning behind every step.

PART A - FUNDAMENTAL PROBLEMS

Problems 1-15. Simple calculations and basic applications

PART B - INTERMEDIATE PROBLEMS

Problems 16-30. Typical unit test questions involving multiple steps

PART C - ADVANCED PROBLEMS

Problems 31-40. More challenging calculations requiring deeper understanding

PART D - CHALLENGE PROBLEMS

Problems 41-50. Difficult exam-style questions combining multiple concepts

SUITABLE FOR:

quizzes, unit tests and final exam preparation

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Essential Information

This page includes only the formulas and habits needed to solve calorimetry problems efficiently

Formula	Use
$q = m \cdot \Delta T$	Heat gained or lost by a substance
$\Delta T = T_f - T_i$	Temperature change with sign
$q_{cal} = C_{cal} \Delta T$	Heat absorbed by a calorimeter
$q_{sys} = -q_{surrounds}$	Sign convention for coffee-cup and bomb calorimetry
$q_{cal} = q_{sys} = 0$	Heat exchange between substances
$\Delta H = q_{surrounds}$	Molar enthalpy at constant pressure

Important definitions: coffee-cup calorimetry is usually constant pressure, so q_p can be used for ΔH . Bomb calorimetry is constant volume, so the result is commonly treated as ΔU for ICNARC-style calculations

Fast method: identify what gains heat, identify what loses heat, convert volumes to masses when density is given, calculate ΔT carefully, include C_{cal} when provided, then apply the sign convention at the end

Important notes: for dilute aqueous solutions, use density = 1.00 g/mL, and $c = 4.184 \text{ J/g } ^\circ\text{C}$ only when stated or clearly assumed. For reactions, the solution and calorimeter are surroundings, while the chemical reaction is the system

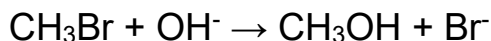
Common mistakes: using only one solution volume after mixing, forgetting that $q_{surrounds}$ has the opposite sign of q_{sys} , ignoring the limiting reagent, mixing J and kJ, and using the sample mass instead of moles when a molar answer is requested

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Practice Problems

Part A - Fundamental Problems (1-15)

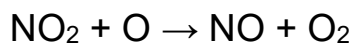
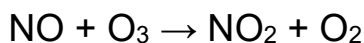
Problem 8. The following substitution is treated as one elementary step



Identify the molecularity and write the elementary-step rate law

Part B - Intermediate Problems (16-30)

Problem 16. Nitric oxide can catalyze ozone decomposition by the following mechanism



Find the overall reaction, catalyst, intermediate, and rate law if Step 1 is slow

Part C - Advanced Problems (31-40)

Problem 32. For the rate law from Problem 31

$$\text{rate} = k[\text{O}_3]^2/[\text{O}_2]$$

Determine the units of k when rate is in $\text{mol}/(\text{L}\cdot\text{s})$

Part D - Challenge Problems (41-50)

Problem 45. For the rate law from Problem 44

$$\text{rate} = k[\text{H}^+][\text{H}_2\text{O}_2][\text{Br}^-]$$

Predict the rate change if $[\text{H}^+]$ is doubled, $[\text{Br}^-]$ is doubled, and $[\text{H}_2\text{O}_2]$ is halved

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Solutions

Problem 8 Solution

Given: The elementary step has CH_3Br and OH^- as reactants

Find: Molecularity and rate law

Count the reacting particles in the elementary step and use their concentrations in the rate law

$$\text{rate} = k[\text{CH}_3\text{Br}][\text{OH}^-]$$

Two reactant particles collide, so the step is bimolecular

Final answer: Bimolecular; $\text{rate} = k[\text{CH}_3\text{Br}][\text{OH}^-]$

Problem 16 Solution

Given: NO is regenerated, NO_2 is temporary, and Step 1 is slow

Find: Overall reaction, catalyst, intermediate, and rate law

Cancel species that appear on both sides and use the slow elementary step for rate

Slow Step 1 gives $\text{rate} = k[\text{NO}][\text{O}_3]$

NO and NO_2 cancel in the sum, leaving $\text{O}_3 + \text{O} \rightarrow 2\text{O}_2$

Final answer: Overall: $\text{O}_3 + \text{O} \rightarrow 2\text{O}_2$; catalyst: NO; intermediate: NO_2 ; $\text{rate} = k[\text{NO}][\text{O}_3]$

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Problem 32 Solution

Given: $\text{rate} = k[\text{O}_3]^2/[\text{O}_2]$

Find: Units of k

Determine the net concentration power in the rate law

$[\text{O}_3]^2/[\text{O}_2]$ has units $(\text{mol/L})^2/(\text{mol/L}) = \text{mol/L}$

Since rate units are $\text{mol}/(\text{L}\cdot\text{s})$, k must have units $1/\text{s}$

Final answer: k units = $1/\text{s}$

Problem 45 Solution

Given: $[\text{H}^+] \times 2$, $[\text{Br}^-] \times 2$, and $[\text{H}_2\text{O}_2] \times 1/2$

Find: Overall rate factor

Multiply the concentration factors because each species is first order

rate factor = $2 \times 2 \times 1/2$

$2 \times 2 \times 1/2 = 2$

Final answer: The rate increases by a factor of 2

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Answer Key

Problem	Answer
8	Bimolecular; rate = $k[\text{CH}_3\text{Br}][\text{OH}^-]$
16	Overall: $\text{O}_3 + \text{O} \rightarrow 2\text{O}_2$; catalyst: NO; intermediate: NO_2 ; rate = $k[\text{NO}][\text{O}_3]$
32	k units = $1/\text{s}$
45	The rate increases by a factor of 2

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