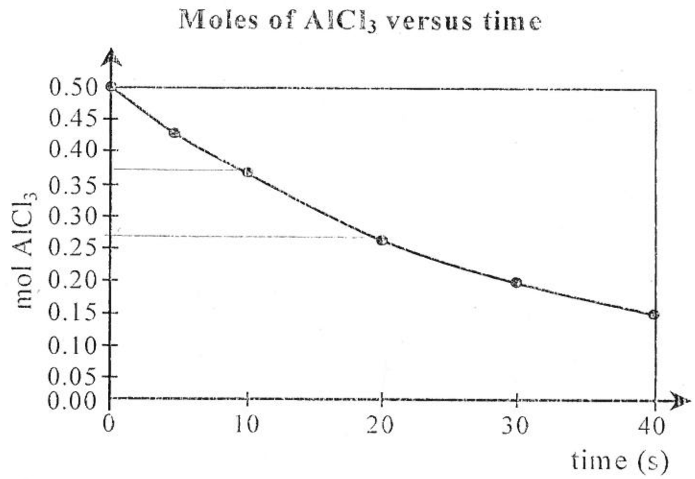
26/28 = 93%

Reaction Rates Assignment

1. A student was studying the electrolysis of 1 L of aqueous aluminum chloride ($AlCl_{3(aq)}$) according to the following equation:

$$2\ AlCl_{3(aq)} \longrightarrow 2Al_{(s)} + 3\ Cl_{2(g)}$$



Using the graph above, calculate the average rate of formation of chlorine gas from 10 s to 20 s.

SHOW ALL YOUR WORK

1. $\Delta[AlCl_3] = 0.37 mol - 0.27 mol = -0.1 mol$

2. $-0.1 mol \times \frac{3\ mol\ Cl_2}{2\ mol\ AlCl_3} = {}^{+}0.15\ mol\ Cl_2$

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3. $Rate = \frac{\Delta\ mol\ Cl_2}{\Delta t}$ $Rate = \frac{0.15\ mol\ Cl_2}{10 s}$ $Rate = 0.015\ mol/s$

Answer: The average rate of formation of $Cl_{2(g)}$ between 10 s and 20 s was 0.015 mol/s

2. Study the following hypothetical chemical reaction:

$$A_{(s)} + 2\ BC_{(aq)} \rightarrow AC_{2(aq)} + B_{2(g)}$$

The table below shows the consumption of the reactant BC. Complete the table and find the average rate of formation of the product B_2 between 5 and 30 seconds in mol/L/s.

Time (seconds)	Concentration of BC (mol/L)	Concentration of B_2 (mol/L)
0	4.0	0
5	3.0	0.50
10	2.3	0.85
15	1.7	1.15
20	1.4	1.30
25	1.2	1.40
30	1.1	1.45 ✓

a) Show the work for one interval.
b) Determine the rate of formation of B_2 for the interval 5s to 25s.

a) Δ mol/L BC = 3.0 mol − 4.0 mol = −1.0 mol ✓

$-1.0\ \text{mol} \times \dfrac{1\ \text{mol}\ B_2}{2\ \text{mol BC}} = +0.5\ \text{mol/L}\ B_2$ ✓

$+0.5\ \text{mol/L} + 0\ \text{mol/L} = 0.5\ \text{mol/L}\ B_2$ ✓

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b) $\text{Rate} = \dfrac{\Delta\ \text{mol/L}\ B_2}{\Delta t}$ $\quad \text{Rate} = \dfrac{1.4\ \text{mol/L} - 0.5\ \text{mol/L}}{20\ \text{s}}$ $\quad \text{Rate} = \dfrac{0.045\ \text{mol}}{L \cdot s}$

3. A two-step reaction mechanism has been suggested for the following overall reaction:

$$2\ H_{2\,(g)} + 2\ NO_{(g)} \rightarrow N_{2\,(g)} + 2\ H_2O_{(g)}$$

First reaction: $H_{2\,(g)} + 2\ NO_{(g)} \rightarrow N_2O_{(g)} + H_2O_{(g)}$

Second reaction: $H_{2\,(g)} + N_2O_{(g)} \rightarrow N_{2\,(g)} + H_2O_{(g)}$

The first step involves a collision between one molecule of hydrogen (H_2) and two molecules of nitrogen monoxide (NO). It has a reaction rate of 3.45×10^{-3} mol/(L·s).

The second step involves a collision between another molecule of hydrogen and a molecule of nitrous oxide (N_2O). It occurs at a rate of 0.45 mol/(L·s).

a) What reaction intermediate is formed by this mechanism and what does it become when the reaction is complete?

The reaction intermediate is $N_2O_{(g)}$ and it becomes $N_{2\,(g)} + 2H_2O_{(g)}$.

b) Using collision theory, explain why the first elementary reaction has a lower rate than the second reaction.

The 1st elementary reaction has a lower rate because 2NO is more complex than N_2O so the bonds are harder to break.

↳explain.

c) What is the rate-determining step in this mechanism?

The rate determining step is the first step (first reaction).

d) In terms of energy, compare the activated complexes formed during each of the steps.

The activated complex of the 1st reaction is higher than the 2nd reaction because it requires a higher activation energy and the 1st reaction is more complex.

e) What will the rate of the overall reaction be in moles per litre-second (mol/(L·s))?

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Answer: 3.45×10^{-3} mol/(L·s) → slowest step

f) Which of the following energy diagrams best represents this chemical reaction? Circle the correct answer.

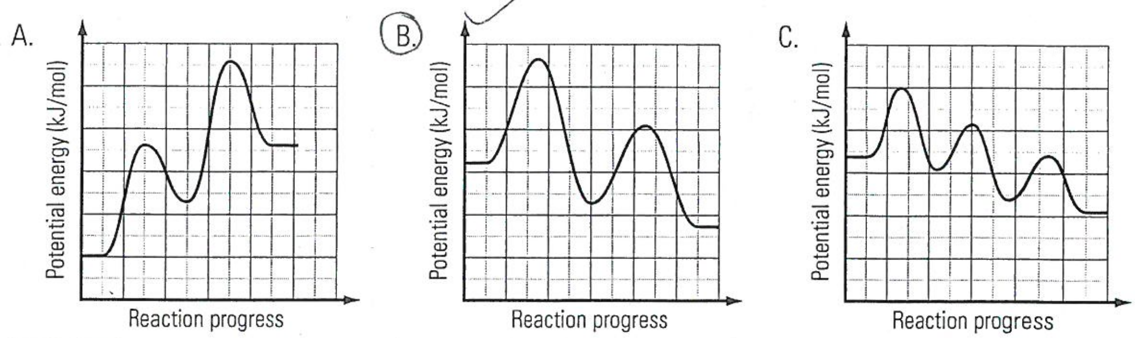

Figure 2 Energy diagrams

$\frac{L^{\square}}{mol^{\triangle} \cdot s}$

4. Express the rate laws for the following reactions and find the units of the rate constant, assuming that these are elementary reactions:

$K = \frac{\frac{mol}{K \cdot s}}{\frac{mol^2}{L^2}}$ $\frac{L}{mol}$

a) $Ba_2XeO_{6\,(s)} + 2\,H_2SO_{4\,(aq)} \rightarrow 2\,BaSO_{4\,(s)} + 2\,H_2O_{(l)} + XeO_{4\,(g)}$

Rate = $K\,[H_2SO_4]^2$ units: $L/mol \cdot s$ = K

b) $IF_{(g)} + 2\,F_{2\,(g)} \rightarrow IF_{5\,(g)}$

Rate = $K\,[IF]^1[F_2]^2$ units: $L^2/mol^2 \cdot s$ = K

c) $2\,At_{(s)} + SO_{2\,(g)} + 2H_2O_{(l)} \rightarrow 2\,At^-_{(aq)} + 3\,H^+_{(aq)} + HSO_4^-{}_{(aq)}$

Rate = $K\,[SO_2]^1$ units: $L^0/mol^0 \cdot s = \frac{1}{s}$ = K

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5. Consider the following elementary reaction:

$$SiCl_{4\,(l)} + 6\,NH_{3\,(g)} \rightarrow Si(NH)_{2\,(s)} + 4\,NH_4Cl_{(s)}$$

After 45 minutes of reacting in a 2.0-L container, the reaction rate is 2.3×10^{-4} mol/(L·s). The quantities of silicon tetrachloride ($SiCl_4$) and ammonia (NH_3) are then 0.25 mol and 0.40 mol.

a) What is the rate constant of this reaction?

$r = 2.3 \times 10^{-4}$ mol/L·s

$C = \frac{n}{V}$

$C = \frac{0.40\,mol}{2.0\,L}$

$C = 0.2\,mol/L$

$r = K\,[NH_3]^6$

$2.3 \times 10^{-4}\,mol/L \cdot s = K\,[NH_3]^6$

$K = \frac{2.3 \times 10^{-4}\,mol/L \cdot s}{(0.2\,mol/L)^6}$

$K = 3.60\,L^5/mol^5 \cdot s$

Answer: $3.60\,L^5/mol^5 \cdot s$

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b) The volume of the container is then increased to 4.0-L. What is the reaction rate in mol/L·s?

$C = \frac{n}{V}$

$C = \frac{0.40 mol}{4.0 L}$

$C = 0.1 mol/L$

$r = K [NH_3]^6$

$r = 3.60\ L^5/mol^5 \cdot s\ [0.1 mol/L]^6$

$r = 3.6 \times 10^{-6}$ mol/L·s

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