Name Answer Key

Practice Exam: Paper 2
Topic 6: Kinetics

SL Score	HL Score
/28	/62

SL

1. Factors that affect the rate of a chemical reaction include particle size, concentration of reactants and the temperature of the reaction.

(i) Define the term *rate of a chemical reaction*.

Increase in concentration of products or decrease in concentration of reactants per unit of time

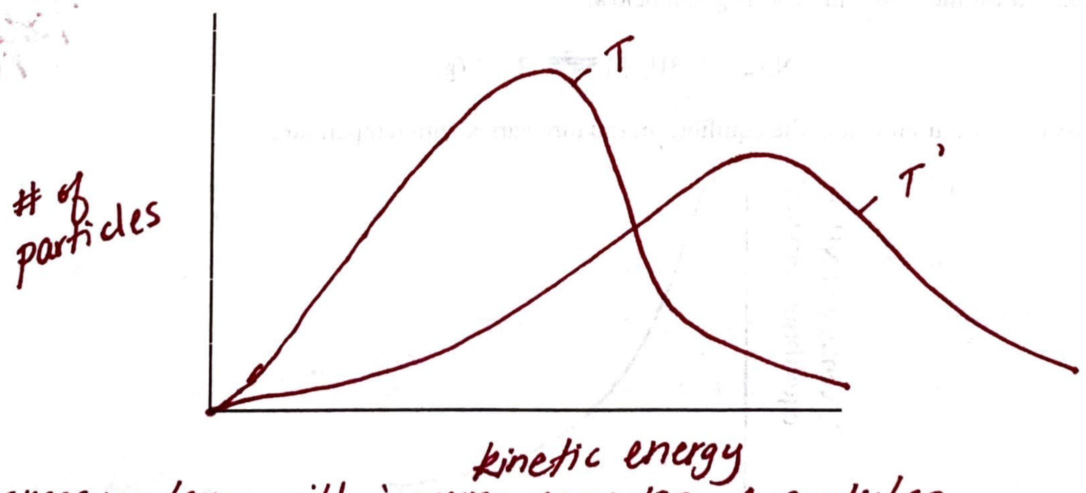
(1)

(ii) List the **three** characteristic properties of reactant particles which affect the rate of reaction as described by the collision theory.

- Frequency of collisions
- kinetic energy of particles
- collision geometry
- proportion of particles with $E > E_a$

(3)

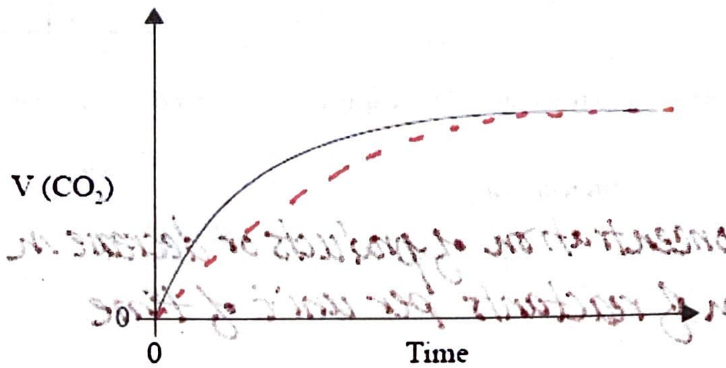
2. On the axes below sketch **two** Maxwell-Boltzmann energy distribution curves for the same sample of gas, one at a temperature *T* and another at a higher temperature T'. Label both axes. Explain why raising the temperature increases the rate of a chemical reaction.



Increasing temp will increase proportion of particles with $E > E_a$

(5)

3. The graph below shows how the volume of carbon dioxide formed varies with time when a hydrochloric acid solution is added to **excess** calcium carbonate in a flask.



(i) Explain the shape of the curve. **(3)**

The initial rate of the rxn is highest due to the concentration of HCl being at its highest. As the reactant HCl becomes more dilute, the production of CO_2 decreases, so the slope decreases and eventually levels out, indicating no more reactant particles are available for collision.

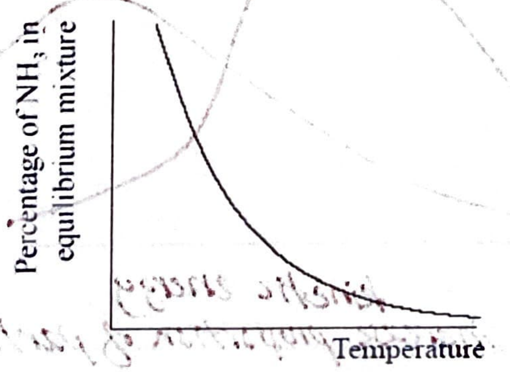
(ii) The experiment is repeated using a sample of hydrochloric acid with double the volume, but half the concentration of the original acid. Draw a second line on the graph to represent this change. Explain why the shape of the curve is different. **(3)**

Even though there's half the concentration of HCl, because there is double the volume, there is the same # of moles. This means the volume of CO_2 produced will be the same, but because the concentration is lower, the initial rate of the reaction.

4. The Haber process enables the large-scale production of ammonia needed to make fertilizers. The equation for the Haber process is given below.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

The percentage of ammonia in the equilibrium mixture varies with temperature.



(i) Use the graph to deduce whether the forward reaction is exothermic or endothermic and explain your choice. **(2)**

An increase in temp causes less NH_3 to be produced, indicating an exothermic rxn.

(4 cont.)

(ii) State and explain the effect of increasing the pressure on the yield of ammonia.

There are more moles of gas on the reactants, so increasing pressure will shift equilibrium to the right (fewer # of moles), which would ↑ the yield of ammonia.

(2)

(iii) Explain the effect of increasing the temperature on the rate of reaction.

Increasing the temp. increases the frequency of collisions and the # of particles with E>Ea, causing an increase in the reaction rate.

(2)

5. (a) A solution of hydrogen peroxide, H_2O_2, is added to a solution of sodium iodide, NaI, acidified with hydrochloric acid, HCl. The yellow colour of the iodine, I_2, can be used to determine the rate of reaction.

$$H_2O_2(aq) + 2NaI(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + I_2(aq) + 2H_2O(l)$$

The experiment is repeated with some changes to the reaction conditions. For each of the changes that follow, predict, stating a reason, its effect on the rate of reaction.

(i) The concentration of H_2O_2 is increased at constant temperature.

(2)

This would cause an ↑ in the rxn rate due to more reactant particles, thus ↑ the frequency of collisions.

(ii) The solution of NaI is prepared from a fine powder instead of large crystals.

(2)

No effect because the NaI is still in solution, so it is completely dissolved. Initial state of NaI has no relevance to rxn rate.

(b) Explain why the rate of a reaction increases when the temperature of the system increases.

(3)

↑ temp ↑ rxn rate b/c it ↑ proportion of particles with E>Ea & ↑ frequency of successful collisions.

1. Hydrogen and nitrogen(II) oxide react according to the following equation.

$$2H_2(g) + 2NO(g) \rightleftharpoons N_2(g) + 2H_2O(g)$$

At time = t seconds, the rate of the reaction is: rate = $k[H_2][NO]^2$

(i) Explain precisely what the square brackets around nitrogen(II) oxide, [NO], represent in this context.

Brackets represent the concentration of NO used in the rxn.

(1)

(ii) Deduce the units for the rate constant k.

This is a third-order rxn, so units of k are $mol^{-2}\,dm^{6}\,s^{-1}$

$mol\,dm^{-3}s^{-1} = k(mol\,dm^{-3})(mol\,dm^{-3})^2$ $\quad k = \frac{mol\,dm^{-3}s^{-1}}{mol^{3}dm^{-9}}$

$mol\,dm^{-3}s^{-1} = k(mol^{3}dm^{-9})$

(1)

2. Nitrogen monoxide reacts at 1280 °C with hydrogen to form nitrogen and water. All reactants and products are in the gaseous phase.

(i) The kinetics of the reaction were studied at this temperature. The table shows the initial rate of reaction for different concentrations of each reactant.

experiment	[NO(g)]/ mol dm^{-3} × 10^{-3}	[H$_2$(g)]/ mol dm^{-3} × 10^{-3}	Initial rate/ mol dm^{-3} s^{-1} × 10^{-5}
1	5.00	2.00	1.25
2	10.00	2.00	5.00
3	10.00	4.00	10.00

Deduce the order of the reaction with respect to NO and H_2, and explain your reasoning.

$\left(\frac{10.00}{5.00}\right)^x = \frac{5.00\times10^{-5}}{1.25\times10^{-5}}$ $\qquad \left(\frac{4.00}{2.00}\right)^x = \frac{10.00\times10^{-5}}{5.00\times10^{-5}}$

$2^x = 4$ $\qquad 2^x = 2$

$\therefore x = 2$ $\qquad x = 1$

second-order w/respect to NO $\qquad$ first-order w/respect to H_2

(4)

(ii) Deduce the rate expression for the reaction.

(1)

rate = $k[NO]^2[H_2]$

(iii) Determine the value of the rate constant for the reaction from Experiment 3 and state its units.

$10.00\times10^{-5}\,mol\,dm^{-3}s^{-1} = k[10.00\times10^{-3}\,mol\,dm^{-3}]^2[4.00\times10^{-3}\,mol\,dm^{-3}]$

$10.00\times10^{-5} = 4.00\times10^{-7}k$

(2)

$k = 250\,mol^{-2}dm^{6}s^{-1}$

3. The gas-phase decomposition of dinitrogen monoxide is considered to occur in two steps.

Step 1: $N_2O(g) \longrightarrow N_2(g) + O(g)$

Step 2: $N_2O(g) + O(g) \longrightarrow N_2(g) + O_2(g)$

The experimental rate expression for this reaction is rate = $k\,[N_2O]$.

(i) Identify the rate-determining step. **(1)**

first step / step 1

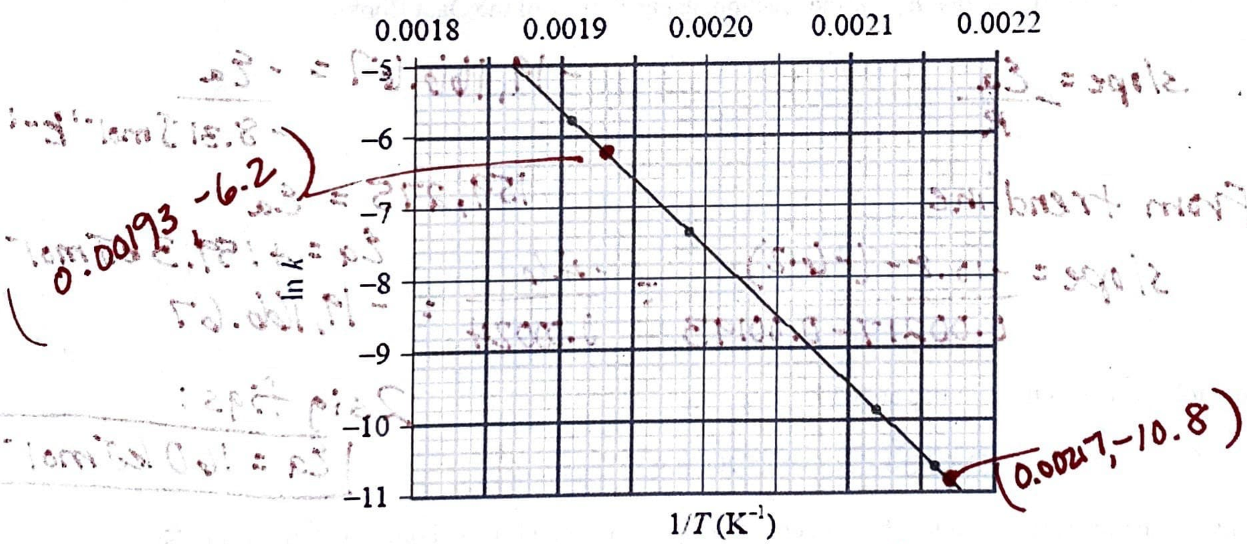
(ii) Identify the intermediate involved in the reaction. **(1)**

O(g)

4. The conversion of CH_3NC into CH_3CN is an exothermic reaction which can be represented as follows.

$CH_3{-}N{\equiv}C \longrightarrow$ transition state $\quad CH_3{-}C{\equiv}N$

This reaction was carried out at different temperatures and a value of the rate constant, k, was obtained for each temperature. A graph of ln k against $1/T$ is shown below.



(i) Define the term *activation energy*, E_a. **(1)**

minimum energy required for reactants to be converted to products in a rxn.

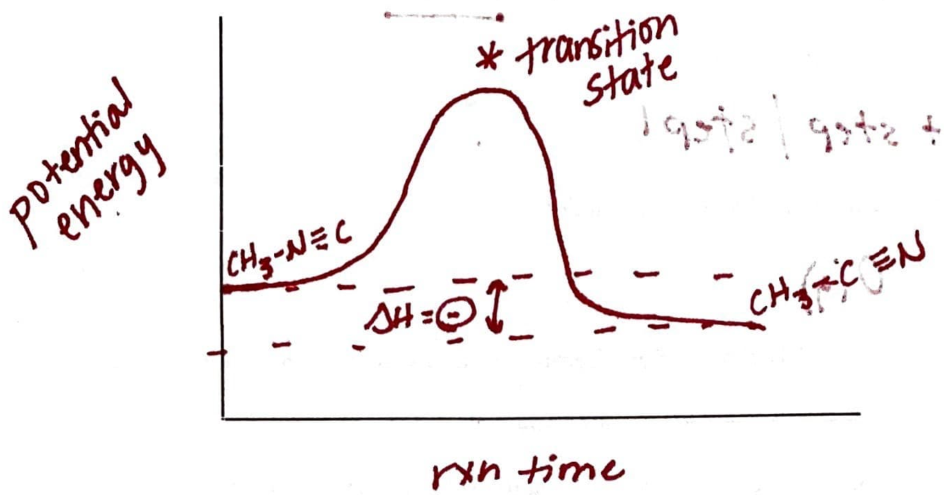
(ii) Describe qualitatively the relationship between the rate constant, k, and the temperature, T. **(1)**

they are ~~directly~~ exponentially proportional, so as T ↑, k ↑.
However, because ln k vs 1/T produces a linear function, as T ↑, k ↑ in an exponential manner.

(4 cont.)

exothermic rxn

(iii) Construct the enthalpy level diagram and label the activation energy, E_a, the enthalpy change, ΔH, and the position of the transition state.



(3)

(iv) Calculate the activation energy, E_a, for the reaction, using Table 1 of the Data Booklet.

slope = $-\frac{Ea}{R}$

From trendline

slope = $\frac{-10.8-(-6.2)}{0.00217-0.00193} = \frac{-4.6}{0.00024} = -19{,}166.67$

$-19{,}166.67 = \frac{-Ea}{-8.31\ J\ mol^{-1}K^{-1}}$

$159{,}275 = Ea$

$Ea = +159.3\ kJ\ mol^{-1}$

2 sig figs:

$Ea = 160\ kJ\ mol^{-1}$

(4)

5. Consider the following reaction studied at 263 K.

$$2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$$

It was found that the forward reaction is first order with respect to Cl_2 and second order with respect to NO. The reverse reaction is second order with respect to NOCl.

(i) State the rate expression for the forward reaction. (1)

rate = $k[NO]^2[Cl_2]$

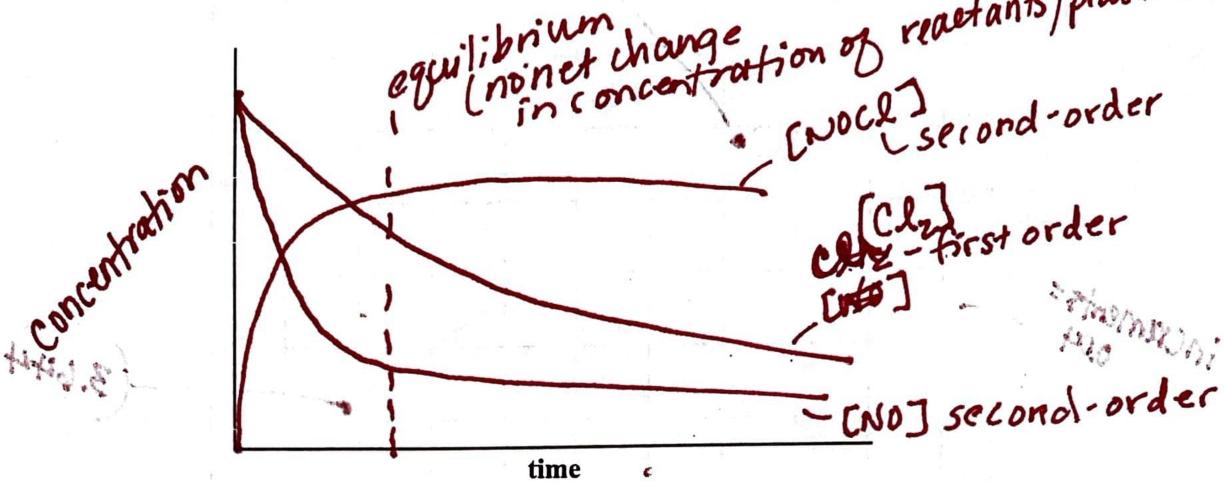
(ii) Predict the effect on the rate of the forward reaction and on the rate constant if the concentration of NO is halved. (2)

If the [NO] is halved, the rate of the forward rxn will ↓ by a factor of 4 since it is second order w/respect to NO.

Rate constant will not be affected, as k is only affected by temperature.

(5 cont.)

(iii) 1.0 mol of Cl_2 and 1.0 mol of NO are mixed in a closed container at constant temperature. Sketch a graph to show how the concentration of NO and NOCl change with time until after equilibrium has been reached. Identify the point on the graph where equilibrium is established.



(4)

6. Consider the following reaction.

$$NO_2(g) + CO(g) \rightarrow NO(g) + CO_2(g)$$

Possible reaction mechanisms are:

Above 775 K:	$NO_2 + CO \rightarrow NO + CO_2$	*slow*
Below 775 K:	$2NO_2 \rightarrow NO + NO_3$	*slow*
	$NO_3 + CO \rightarrow NO_2 + CO_2$	*fast*

Based on the mechanisms, deduce the rate expressions above and below 775 K.

Above 775K : rate = $k[NO_2][CO]$

Below 775K: rate = $k[NO_2]^2$

(2)

7. State **two** situations when the rate of a chemical reaction is equal to the rate constant.

1) rxn is zero-order w/respect to the reactants

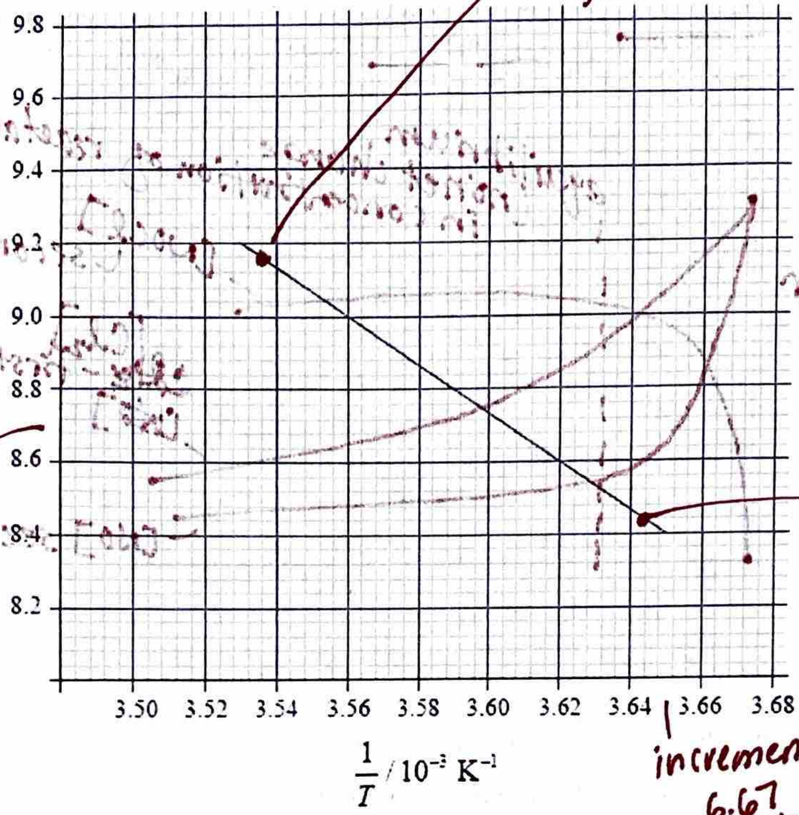
2) when concentrations of all reactants in rate law are equal to 1 mol dm^{-3}

(2)

8. Consider the following graph of ln k against $\frac{1}{T}$ for the first order decomposition of N_2O_4 into NO_2. Determine the activation energy in kJ mol^{-1} for this reaction.

(3.~~500~~ 3.536, 9.16)

increments = 0.04



(3.644, 8.44)

increments = 0.004

(3)

$$\text{slope} = -\frac{E_a}{R}$$

$$-6.67 \times 10^3 = -\frac{E_a}{8.31\ J\ mol^{-1}\ K^{-1}}$$

$$\text{slope} = \frac{8.44 - 9.16}{3.644 - 3.536} = \frac{-0.72}{0.108 \times 10^{-3}} = -6.67 \times 10^3$$

$= 55{,}400\ J\ mol^{-1}$

$= 55.4\ kJ\ mol^{-1}$