Name ____________________

SL Score	HL Score
/28	/62

Key

Practice Exam: Paper 2
Topic 6: Kinetics

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 product(s) per unit of time. or Decrease in concentration of reactants per unit of time.

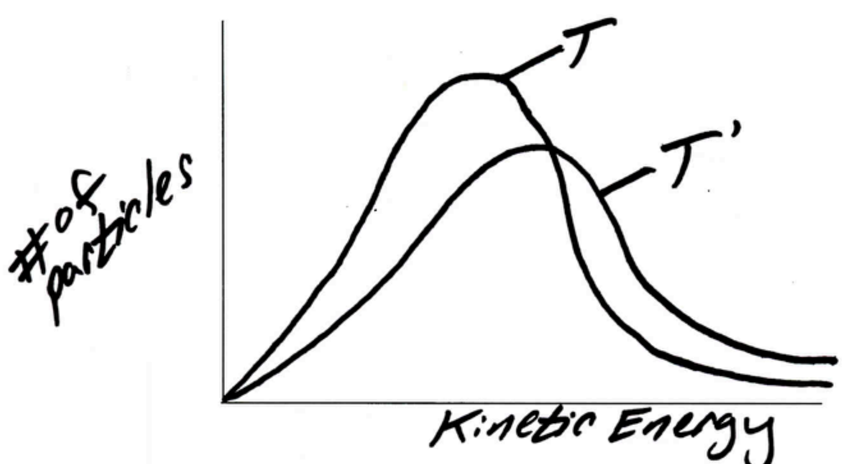
(1)

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

1.) Frequency of collisions
2.) Kinetic energy of particles
3.) collision geometry (orientation)

(3)

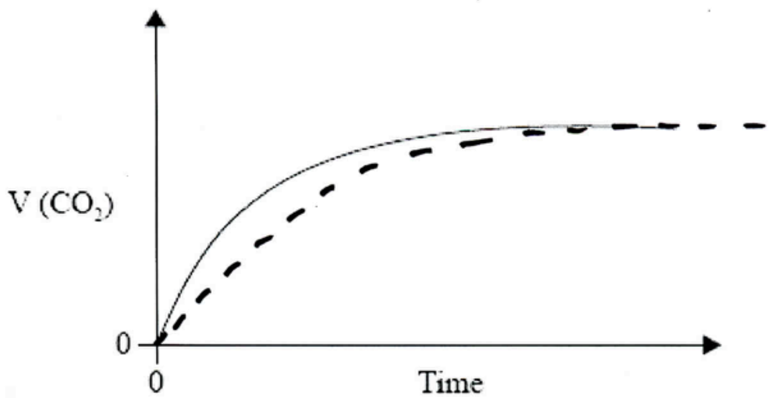
2. On the axes below sketch **two** Maxwell-Boltzmann energy distribution curves for the same sample of gas, one at a temperature T and another at a higher temperature T'. Label both axes. Explain why raising the temperature increases the rate of a chemical reaction.



At higher temperatures more particles will have sufficient activation energy for a reaction to occur.

(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)

Initially the concentration of HCl is the highest and the rate of production of CO_2 at a maximum. As HCl becomes more dilute the production of CO_2 slows and the curve flattens.

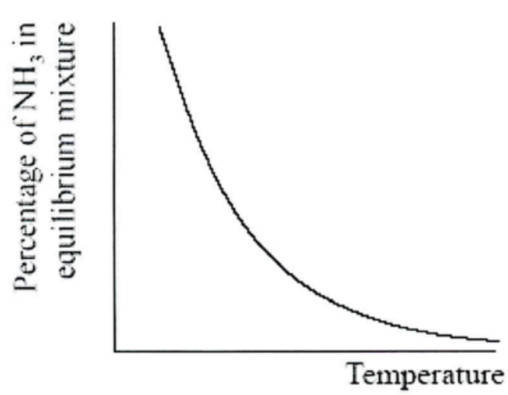
(ii) The experiment is repeated using a sample of hydrochloric acid with double the volume, but half the concentration of the original acid. Draw a second line on the graph to represent this change. Explain why the shape of the curve is different.

The lower concentration of HCl will result in a slower rate of reaction, but will produce the same volume of CO_2 as the total number of moles of HCl is unchanged. (3)

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

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{Heat}$$

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)

Exothermic; as the temperature increases the rate of production of NH_3 decreases.

(4 cont.)

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

Increasing pressure favors the reaction with fewer moles of gaseous products, so the reaction shifts to the right.

(2)

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

Increasing temperature both increases the number of collisions* and the number of colliding molecules with sufficient activation energy (E_a). *per unit of time!

(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)

Rate increases due to a greater number of H_2O_2 molecules, resulting in more collisions per unit of time.

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

(2)

No effect; solid NaI is not reacting, only an aqueous solution of NaI.

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

(3)

(same answer as #4(iii) above)

HL

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.

The brackets represent the concentration of NO.

(1)

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

$$K = \frac{rate}{[H_2][NO]^2} = \frac{mol/L\cdot s}{mol^3/L^3} = \frac{L^2}{mol^2\cdot s} = \boxed{dm^6\,mol^{-2}\,s^{-1}}$$

recall: 1L = 1dm^3

(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.

From experiments 1 and 2 at constant $[H_2]$, the rate quadruples when [NO] doubles. Therefore second order with respect to NO. From experiments 2 and 3 at constant [NO], the rate doubles when $[H_2]$ doubles. Therefore first order with respect to $[H_2]$

(4)

(ii) Deduce the rate expression for the reaction.

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

(1)

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

$$K = \frac{rate}{[NO]^2[H_2]} = \frac{(10.00\,\frac{mol}{L\cdot s})}{(10.00\,\frac{mol}{L})^2(4.00\,\frac{mol}{L})} = 250\,\frac{L^2}{mol^2\cdot s}$$

x10^{-5}

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

(2)

4

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.

Step 1 (unimolecular!) (1)

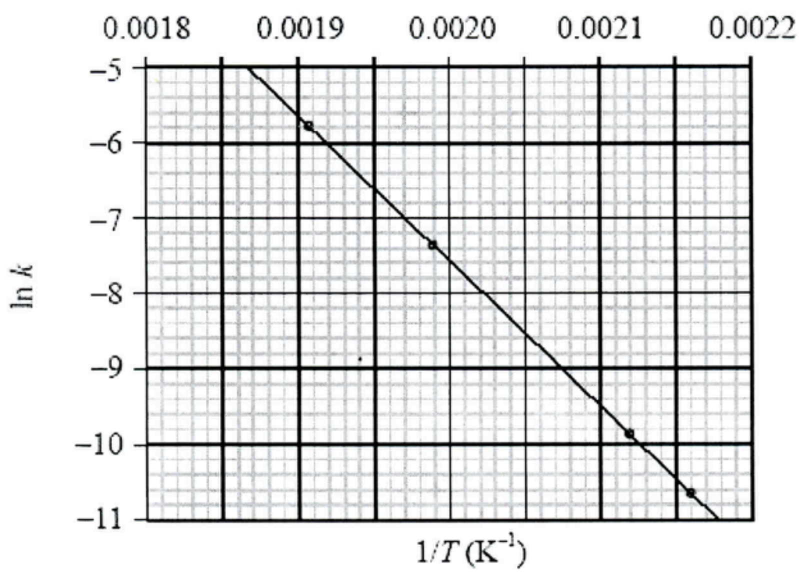
(ii) Identify the intermediate involved in the reaction.

O(g) (1)

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 $\longrightarrow CH_3{-}C{\equiv}N$

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.

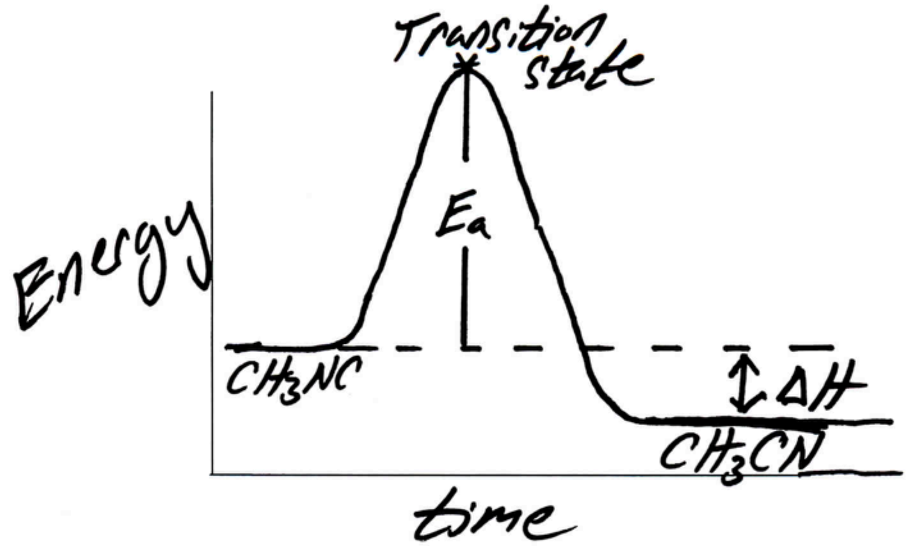
minimum energy needed for a reaction to occur. (1)

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

As the temperature increases, the rate constant (k) increases (exponetially) (1)

(4 cont.)

(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.

$\ln K = -\frac{E_a}{RT} + \ln A$

From the graph,

slope$(m) = -\frac{E_a}{R}$

$m = \frac{(-10.6)-(5.8)}{(0.00216)-(0.00191)} = -19{,}200$

approximate!

$E_a = -mR$

$= -(19{,}200)(8.31)$

$= \boxed{160 \text{ kJ mol}^{-1}}$

(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 $= K[NO]^2[Cl_2]$

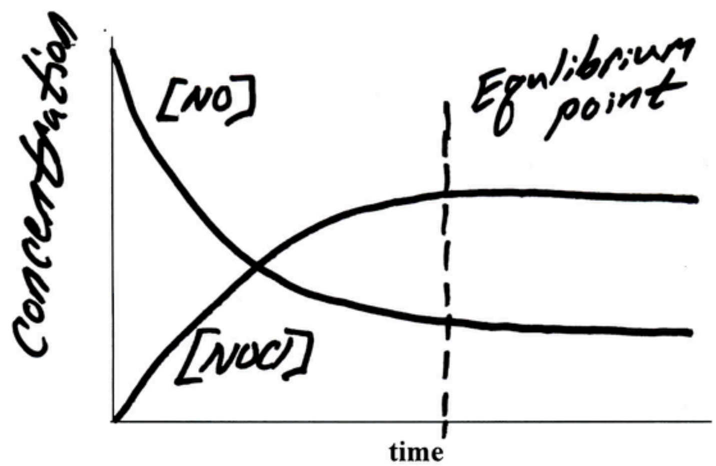
(ii) Predict the effect on the rate of the forward reaction and on the rate constant if the concentration of NO is halved. **(2)**

The rate will decrease by a factor of 4. (rate $\propto (\frac{1}{2})^2 = 0.25$)

The rate constant will not be affected.

(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 775: rate $= k[NO_2]^2$

(2)

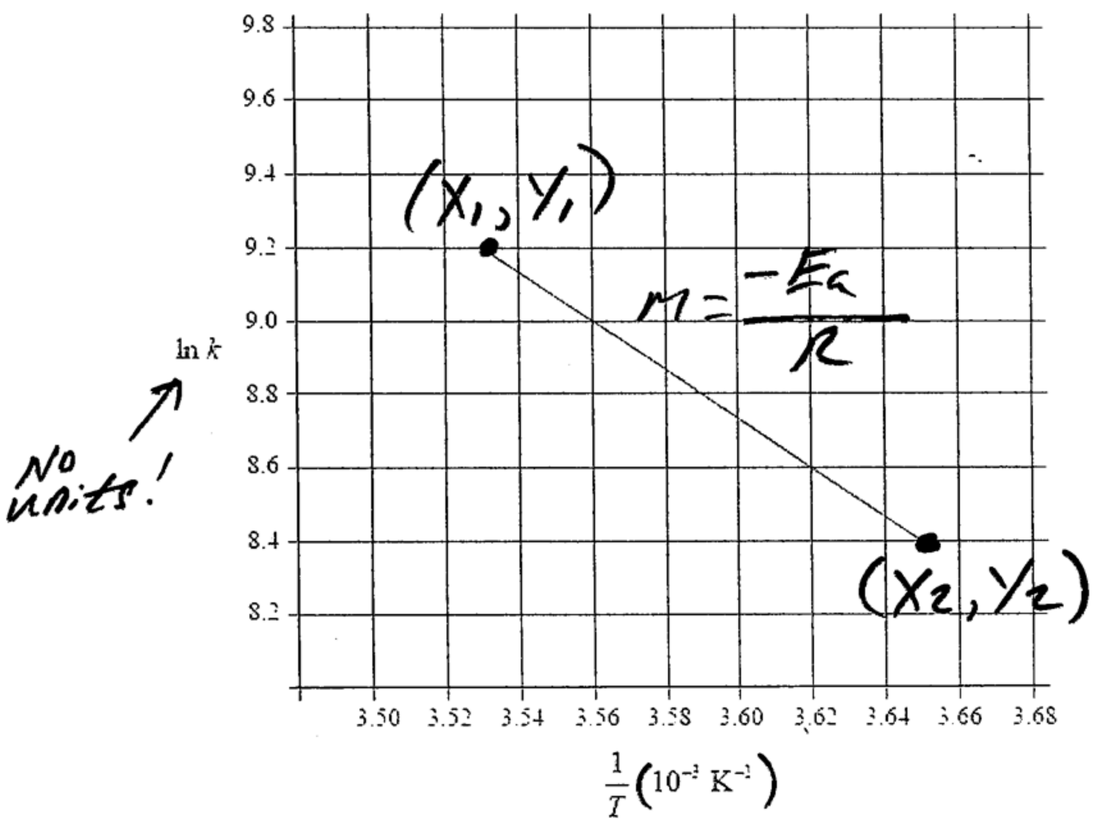
7. State **two** situations when the rate of a chemical reaction is equal to the rate constant.

1.) Zero order reaction

2.) All concentrations $= 1$ mol dm^{-3} (1M)

(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)

$$m = \frac{Y_2 - Y_1}{X_2 - X_1} = \frac{(8.4 - 9.2)}{(3.65 - 3.53) \times 10^{-3} K^{-1}} = -6{,}670 \, K$$

$$E_a = -mR = -(-6{,}670 \cancel{K})\left(8.31 \frac{J}{mol \cdot \cancel{K}}\right) =$$

$$55400 \, J/mol = \boxed{55.4 \, kJ/mol}$$