

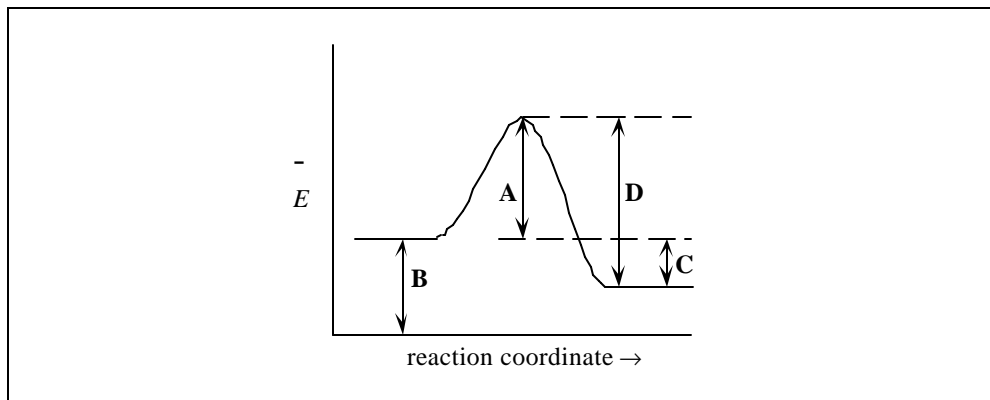
Worksheet 2 – Chapter 14 – Chemical Kinetics

1. The rate equation for a chemical reaction is determined by

- (A) theoretical calculations.
- (B) measuring reaction rate as a function of concentration of reacting species.
- (C) determining the equilibrium constant for the reaction.
- (D) measuring reaction rates as a function of temperature

2.

Which line in the diagram represents the activation energy for a forward reaction?



- (A) A (B) B (C) C (D) D

3. The rate law for the reaction $A + B \rightarrow C + D$

is first order in $[A]$ and second order in $[B]$. If $[A]$ is halved and $[B]$ is doubled, the rate of the reaction will

- (A) remain the same.
- (B) be increased by a factor of 2.
- (C) be increased by a factor of 4.
- (D) be increased by a factor of 8.

4. Bimolecular reactions (e.g. $A + B \Rightarrow C$) generally occur at much slower rates than one would predict from the rate of collision between molecules as calculated from the gas kinetic theory. The discrepancy can be explained in terms of

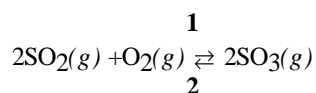
- (A) intermolecular repulsions.
- (B) nonspherical molecular shapes.
- (C) erroneous estimates of molecular size.
- (D) activation energies of reaction.
- (E) the uncertainty principle.

5. If a reaction proceeding by the mechanism $A + B \rightarrow C + D$

occurs at a rate x , and if the concentrations of A and B are both doubled, what will be the new rate of reaction?

- (A) x (B) $2x$ (C) $4x$ (D) $8x$ (E) $16x$

6. Given that



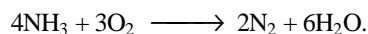
The forward reaction (1) is proceeding at a certain rate at some temperature and pressure; when the pressure is increased, we may expect for the forward reaction (1)

- (A) a greater rate of reaction and a greater yield of SO_3 at equilibrium.
- (B) a greater rate of reaction and the same yield of SO_3 at equilibrium.
- (C) a lesser rate of reaction and a lesser yield of SO_3 at equilibrium.
- (D) a lesser rate of reaction and a greater yield of SO_3 at equilibrium.
- (F) no change in rate or yield.

7. The addition of a catalyst in a chemical reaction

- (A) increases the concentration of products at equilibrium.
- (B) increases the fraction of reactant molecules with a given kinetic energy.
- (C) provides an alternate path with a different activation energy.
- (D) lowers the enthalpy change in the overall reaction.

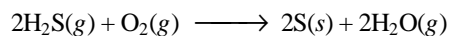
8. The oxidation of ammonia produces nitrogen and water via the reaction



If the rate of formation of N_2 is $2.0 \text{ mol}/(\text{L} \cdot \text{s})$, then the rate at which

- a. H_2O is being formed is $2.0 \text{ mol}/(\text{L} \cdot \text{s})$.
- b. NH_3 reacts is $4.0 \text{ mol}/(\text{L} \cdot \text{s})$.
- c. O_2 reacts is $1.5 \text{ mol}/(\text{L} \cdot \text{s})$.
- d. NH_3 reacts is $0.50 \text{ mol}/(\text{L} \cdot \text{s})$.
- e. H_2O is being formed is $0.67 \text{ mol}/(\text{L} \cdot \text{s})$.

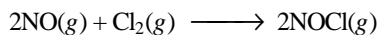
9. From a consideration of the following reaction system



we can conclude that

- a. the reaction is second order in H_2S and first order in O_2 .
- b. the reaction is first order in H_2S and second order in O_2 .
- c. $\text{rate} = k[\text{H}_2\text{S}]^2[\text{O}_2]$.
- d. $\text{rate} = k[\text{H}_2\text{S}][\text{O}_2]$.
- e. None of these conclusions are justified.

10. Nitrosyl chloride is produced from the reaction of nitrogen(II) oxide and chlorine.



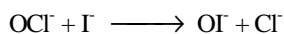
The following initial rates at a given temperature were obtained for the concentrations listed below:

Experiment	Rate $\frac{\text{mol}}{\text{L} \cdot \text{hr}}$	NO $\frac{\text{mol}}{\text{L}}$	Cl ₂ $\left(\frac{\text{mol}}{\text{L}}\right)$
1	2.21	0.25	0.25
2	8.83	0.50	0.25
3	17.5	0.50	0.50

The experimental rate law is

- rate = $k[\text{NO}]$.
- rate = $k[\text{NO}]^{1/2}[\text{Cl}_2]$.
- rate = $k[\text{NO}][\text{Cl}_2]^{1/2}$.
- rate = $k[\text{NO}]^2[\text{Cl}_2]$.
- rate = $k[\text{NO}][\text{Cl}_2]$.

11. The hypochlorite ion oxidizes the iodide ion in aqueous solution as represented by the equation



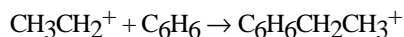
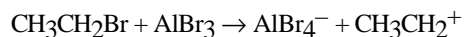
The rate of formation of OI⁻ is given by the rate law

$$\text{Rate} = k \frac{[\text{I}^-][\text{OCl}^-]}{[\text{OH}^-]}$$

The overall reaction order and the order with respect to OH⁻ are

- 0 and 1.
- 0 and -1.
- 1 and -1.
- 2 and 1.
- 2 and -1.

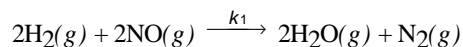
12. The following mechanism has been proposed for the formation of ethylbenzene:



Which substance serves as the catalyst?

- | | | | |
|-----|--------------------------------|-----|--|
| (A) | AlBr ₃ | (C) | CH ₃ CH ₂ ⁺ |
| (B) | AlBr ₄ ⁻ | (D) | C ₆ H ₆ CH ₂ CH ₃ ⁺ |

13. The table presents data for the reaction:



The temperature of the reaction is constant. The initial rate is in arbitrary units.

Exp.	Initial Concentration ($\text{mol}\cdot\text{L}^{-1}$)		Initial Rate
	$[\text{NO}] \times 10^{-3}$	$[\text{H}_2] \times 10^{-3}$	
I	6.0	1.0	18
II	6.0	2.0	36
III	1.0	6.0	3
IV	2.0	6.0	12

What is the rate law for this reaction?

- (A) rate = $k_1 [\text{H}_2] [\text{NO}]$ (C) rate = $k_1 [\text{H}_2]^2 [\text{NO}]$
 (B) rate = $k_1 [\text{H}_2]^2 [\text{NO}]^2$ (D) rate = $k_1 [\text{H}_2] [\text{NO}]^2$

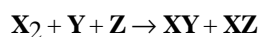
14. The reaction $2\text{A} + 2\text{B} \rightarrow \text{C} + \text{D}$ proceeds by this mechanism:



The rate equation for the reaction is

- (A) rate = $k[\text{A}] [\text{B}]$ (C) rate = $k \frac{[\text{A}]^2 [\text{B}]^2}{[\text{C}] [\text{D}]}$
 (B) rate = $k[\text{A}]^2 [\text{B}]^2$ (D) rate = $k[\text{A}]^2 [\text{B}]$

15. For the reaction



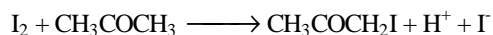
the mechanism was determined to be

- 1) $\text{X}_2 + \text{Y} \rightarrow \text{XY} + \text{X}$ (very slow)
 2) $\text{X} + \text{Z} \rightarrow \text{XZ}$ (very fast)

What is the rate law for this reaction?

- (A) rate = $k [\text{X}_2] [\text{Y}] [\text{Z}]$ (C) rate = $k [\text{X}] [\text{Z}]$
 (B) rate = $k [\text{X}_2]$ (D) rate = $k [\text{X}_2] [\text{Y}]$

16. In aqueous solution, iodine reacts with acetone as represented by the following stoichiometric equation:



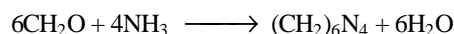
The experimental rate law is

$$\text{Rate} = k[\text{H}^+][\text{CH}_3\text{COCH}_3]$$

According to the information above, an increase in the hydrogen ion concentration has what effect on the reaction?

- It increases the value of the equilibrium constant.
- It decreases the value of the equilibrium constant.
- It increases the rate of the reaction.
- It decreases the rate of the reaction.
- It does not affect the rate of the reaction.

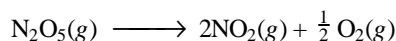
17. For the reaction



the rate is expressed as $\frac{1}{6} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t}$. An equivalent expression is

- $\frac{1}{2} \frac{\Delta[(\text{CH}_2)_6\text{N}_4]}{\Delta t}$.
- $6 \frac{\Delta[\text{CH}_2\text{O}]}{\Delta t}$.
- $-6 \frac{\Delta[\text{CH}_2\text{O}]}{\Delta t}$.
- $-\frac{1}{4} \frac{\Delta[\text{NH}_3]}{\Delta t}$.
- $-\frac{1}{6} \frac{\Delta[\text{H}_2\text{O}]}{\Delta t}$.

18. The gas-phase decomposition of N_2O_5 is a first-order process with a rate constant of $1.50 \times 10^{-3}/\text{s}$ at 55°C . The decomposition reaction is



Ten (10.0) g of N_2O_5 are placed in vessel 1 and 5.0 g of N_2O_5 in vessel 2. The vessels are at the same temperature and pressure. How much time is required for half of the N_2O_5 to decompose in each vessel?

- Vessel 1 requires twice as much time as vessel 2.
- Vessel 2 requires twice as much time as vessel 1.
- Vessel 1 requires three times as much time as vessel 2.
- Vessel 1 requires four times as much time as vessel 2.
- Vessel 1 requires the same amount of time as vessel 2.

- Exam 1 – Wednesday, September 18 – 12:30 PM – Renfrew 111
- Chapters 11 – 12 – 14
- Friday’s lecture will begin Chapter 15, bring lecture notes to class.

1.B 2.A 3.B 4.D 5.C

6. A 7.D 8.b 9.e

10.d 11.c 12.A

13.D 14.D 15.D

16.c 17.d 18.e