**AP Chem Chapter 14 Practice Problems**

**Multiple Choice**

*Identify the choice that best completes the statement or answers the question. You may refer to your green reference sheets, but calculator use is not permitted for multiple choice. Finish all the problems on one page before checking your answers.*

\_\_\_\_ 1. Which of the following experimental conditions is most likely to decrease the rate of a chemical reaction?

|  |  |
| --- | --- |
| a. | Increasing the surface area of any solid reactants present |
| b. | Decreasing the volume of the container in which the reaction takes place |
| c. | Decreasing the temperature at which the reaction is performed |
| d. | Increasing the concentration of reactants |

\_\_\_\_ 2. Which of the following is a correct statement about reaction order?

|  |  |
| --- | --- |
| a. | Reaction order can only be a whole number. |
| b. | Reactio Reaction order can be determined only from the coefficients of the balanced equation for the reaction. |
| c. | Reaction order can be determined only by experiment. |
| d. | A second-order reaction must involve at least two different compounds as reactants. |

\_\_\_\_ 3. The rate of a certain chemical reaction between substances M and N obeys the rate law shown below.

Rate = *k*[M][N]2

The reaction is first studied with [M] and [N] each 1 × 10–3 molar. If a new experiment is conducted with [M] and [N] each 2 × 10–3 molar, the reaction rate will increase by a factor of

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 2 | c. | 6 |
| b. | 4 | d. | 8 |

\_\_\_\_ 4. A reaction was observed for 20 days and the percentage of the reactant remaining after each day

was recorded in the table above. Which of the following best describes the order and the

half-life of the reaction?

|  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Time(days) | 0 | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 10 | 20 |
| % Reactant remaining | 100 | 79 | 63 | 50 | 40 | 31 | 25 | 20 | 10 | 1 |

Reaction Order Half-life (days)

|  |  |
| --- | --- |
| a. | First 3 |
| b. | First 10 |
| c. | Second 3 |
| d. | Second 6 |

\_\_\_\_ 5. If 87.5 percent of a sample of pure iodine-131 decays in 24 days, what is the half-life of iodine-131?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | 6 days | c. | 12 days |
| b. | 8 days | d. | 14 days |

\_\_\_\_ 6. Consider the following reaction, which occurs in the atmosphere.

2NO*(g) +* O2*(g)* → 2 NO2*(g)*

A possible mechanism for the overall reaction represented above is the following.

(1) NO*(g)* + NO*(g)* → N2O2*(g)* slow

(2) N2O2*(g)* + O2*(g)* → 2 NO2*(g)* fast

Which of the following rate laws agrees best with this possible mechanism?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | Rate = k[NO]2 | c. | Rate = |
| b. | Rate = | d. | Rate = k[NO]2[O2] |

\_\_\_\_ 7.

|  |  |  |  |
| --- | --- | --- | --- |
| Experiment | Initial [NO] (mol L-1) | Initial [O2] (mol L-1) | Initial Rate of Formation of NO2  (mol L-1s-1) |
| 1 | 0.10 | 0.10 | 2.5x10–4 |
| 2 | 0.20 | 0.10 | 5.0x10–4 |
| 3 | 0.20 | 0.40 | 8.0x10–3 |

The initial-rate data in the table above were obtained for the reaction represented below. What is the experimental rate law for the reaction?

2 NO*(g)* + O2*(g)* → NO2*(g)*

|  |  |  |  |
| --- | --- | --- | --- |
| a. | Rate = *k*[NO][O2] | c. | Rate = *k*[NO]2[O2] |
| b. | Rate = *k*[NO][O2]2 | d. | Rate = *k*[NO]2[O2]2 |

\_\_\_\_ 8. (CH3)3CCl*(aq)* + OH– → (CH3)3COH*(aq)* + Cl–

For the reaction represented above, the experimental rate law is given as follows.

Rate = *k*[(CH3)3CCl]

If some solid sodium hydroxide is added to a solution that is 0.010–molar in (CH3)3CCl and 0.10–molar in NaOH, which of the following is true? (Assume the temperature and volume remain constant.)

|  |  |
| --- | --- |
| a. | Both the reaction rate and *k* increase. |
| b. | Both the reaction rate and *k* decrease. |
| c. | Both Both the reaction rate and *k* remain the same. |
| d. | The reaction rate increases but *k* remains the same. |

\_\_\_\_ 9. rate = *k*[X]

For the reaction whose rate law is given above, a plot of which of the following is a straight line?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | [X] *versus* time | c. | 1/[X] *versus* time |
| b. | ln [X] *versus* time | d. | [X] *versus* 1/time |

\_\_\_\_ 10. Step 1. N2H2O2  N2HO2– + H+ (fast equilibrium)

Step 2. N2HO2– → N2O + OH– (slow)

Step 3. H+ + OH– → H2O (fast)

Nitramide, N2H2O2, decomposes slowly in aqueous solution. This decomposition is believed to occur according to the reaction mechanism above. The rate law for the decomposition of nitramide that is consistent with this mechanism is given by which of the following?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | Rate = *k* [N2H2O2] | c. |  |
| b. | Rate = *k* [N2H2O2] [H+] | d. |  |

\_\_\_\_ 11. 

The energy diagram for the reaction X + Y→ Z is shown above. The addition of a catalyst to this reaction would cause a change in which of the indicated energy differences?

|  |  |  |  |
| --- | --- | --- | --- |
| a. | I only | c. | III only |
| b. | II only | d. | I and II only |

\_\_\_\_ 12. The proposed steps for a catalyzed reaction between Ce4+ and Tl+ are represented below.

Step 1: Ce4+ + Mn2+ → Ce3+ + Mn3+

Step 2: Ce4+ + Mn3+ → Ce3+ + Mn4+

Step 3: Mn4+ + Tl+ → Tl3+ + Mn2+

Which of the following statements are CORRECT?

I. The overall balanced equation is 2 Ce4+ + Tl+ → Tl3+ + 2 Ce3+

II. Mn2+ is a catalyst for the reaction.

III. Mn4+ is a reaction intermediate.

|  |  |  |  |
| --- | --- | --- | --- |
| a. | I, only | d. | I and III |
| b. | I and II | e. | I, II and III |
| c. | II and III |  |  |

**Problem**

You will need your calculator for the free response questions. Keep your periodic table and formula sheet handy. Show all your work!

13. The half-life of a first-order decomposition reaction is 2.36 hours. If the initial concentration of reactant is 0.52 M, what is the concentration of reactant after 752 seconds? (2 pts)

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14. What is the half-life of a first-order reaction with a rate constant of 0.457 s-1? (1 pt)

Answer the following questions related to the kinetics of chemical reactions. (2005)

15.

I– *(aq)* + ClO– *(aq)* → IO–*(aq)* + Cl– *(aq)*

Iodide ion, I–, is oxidized to hypoiodite ion, IO–, by hypochlorite, ClO–, in basic solution according to the equation above. Three initial-rate experiments were conducted; the results shown in the following table.

|  |  |  |  |
| --- | --- | --- | --- |
| Experiment | [I–]  (mol L–1) | [ClO–]  (mol L–1) | Initial Rate of Formation of IO– (mol L–1 s–1) |
| 1 | 0.017 | 0.015 | 0.156 |
| 2 | 0.052 | 0.015 | 0.476 |
| 3 | 0.016 | 0.061 | 0.596 |

(a) Determine the order of the reaction with respect to each reactant listed below. Justify your answer. (2 pts)

(i) I–*(aq)*

(ii) ClO–*(aq)*

(b) For the reaction,

(i) write the rate law that is consistent with your answers for part (a); (1 pt)

(ii) calculate the value of the specific rate constant, *k*, and specify units. (2 pts)

16. An environmental concern is the depletion of O3 in Earth's upper atmosphere, where O3 is normally in equilibrium with O2 and O. A proposed mechanism for the depletion of O3 in the upper atmosphere is shown below.

Step I O3 + Cl → O2 + ClO

Step II ClO + O → Cl + O2

(a) Write a balanced equation for the overall reaction represented by Step I and Step II above. (1 pt)

(b) Clearly identify the catalyst in the mechanism above. Explain your answer. (1 pt)

(c) Clearly identify the intermediate in the mechanism above. Explain your answer. (1 pt)

(d) If the rate law for the overall reaction is found to be rate = *k*[O3][Cl], determine the following.

(i) The overall order of the reaction (1 pt)

(ii) Appropriate units for the rate constant, *k* ( 1 pt)

(iii) The rate-determining step of the reaction. Explain your answer. (1 pt)

17. 1998 D (modified)

Answer the following questions regarding the kinetics of chemical reactions.

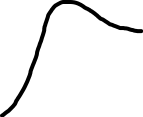
The diagram below shows the energy pathway for the hypothetical reaction

2A + B2 → 2 AB



2AB

potential energy



2A + B2

reaction progress

(a) Is this reaction endothermic or exothermic? Explain your answer. (1 pt)

(b) Clearly label the following directly on the diagram.

(i) The activation energy *(Ea)* for the forward reaction (1 pt)

(ii) The activation energy for the reverse reaction

(iii) The enthalpy change (*H*) for the reaction (1 pt)

(iv) How the diagram would be different in the presence of a catalyst (1 pt)

(c) The reaction 2 N2O5 → 4 NO2 + O2 is first order with respect to N2O5.

(i) Using the axes below, complete the graph that represents the change in [N2O5] over time as the reaction proceeds.

[N2O5]

time

(ii) Describe how the graph in (i) could be used to find the reaction rate at a given time, *t*. (1 pt)

(iii) Write the rate law for the reaction. (1 pt)

(iv) If more N2O5 were added to the reaction mixture at constant temperature, what would be the effect on the rate constant, *k* ? Explain. (1 pt)

(d) Data for the chemical reaction 2A → B + C were collected by measuring the concentration of A at 10-minute intervals for 80 minutes. The following graphs were generated from analysis of the data.



Use the information in the graphs above to answer the following.

(i) Write the rate-law expression for the reaction. Justify your answer. (1 pt)

(ii) Describe how to determine the value of the rate constant for the reaction. (1 pt)

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**AP Chem Chapter 14 Practice Problems**

**Answer Section**

**MULTIPLE CHOICE**

1. ANS: C PTS: 1

2. ANS: C PTS: 1

3. ANS: D PTS: 1

4. ANS: A PTS: 1

5. ANS: B PTS: 1

6. ANS: A PTS: 1

7. ANS: B PTS: 1

8. ANS: C PTS: 1

9. ANS: B PTS: 1

10. ANS: C PTS: 1

11. ANS: D PTS: 1

12. ANS: E PTS: 1

**PROBLEM**

13. ANS:

0.49M

Use the first order integrated rate law

Must show the setup for 1 pt

Correct answer for 1 pt

PTS: 1

14. ANS:

1.52 s

Use the t1/2 = ln2/k equation

PTS: 1

15. ANS:

(a) (i) comparing expt. 1 to expt. 2, while the hypochlorite concentration remains constant, the iodide concentration is essentially tripled { = } and the initial rate is essentially tripled { = }. This indicates a first order with respect to the iodide ion.

(ii) comparing expt. 1 to expt. 3, while the iodide concentration remains essentially constant (a 2.7% drop), the hypochlorite concentration is essentially quadrupled { = } and the initial rate is essentially quadrupled { = }. This indicates a first order with respect to the hypochlorite ion.

OR

(i) from experiments 1 & 2

=

=

3.05 ≈ 3.1*m*, where *m* = 1

PTS: 1

16. ANS:

(a) O3 + O → 2 O2

(b) Cl; used in step I and regenerated in step II, the amount at the end is the same as the beginning

(c) ClO; product of step I and used in step II, an intermediate is a material the is produced by a step and consumed later, it does not show as either a product or reactant in the overall equation.

(d) (i) second order overall

(ii) *k* unit is *M*-1 time-1

(iii) step 1. The rate law applies to the concentration of the materials in the slowest step, the rate determining step.

PTS: 1

17. ANS:

Answer

a) The reaction is exothermic...the products are lower in energy than the reactants.

b)

(b)

(c) i

ii the rate at time, *t*, is the slope of the tangent to the curve at time *t*

iii rate = k[N2O5]

iv *k* would remain unchanged, it is temperature dependent, not concentration dependent.

(d)i since the graph of ln[A] is a straight line, this indicates that it its 1st order with respect to A, ∴, rate = k [A]

ii *k* = - slope of the straight line of the ln[A] *vs.* time graph

PTS: 1