Chapter 17
Study Guide - Answers

Name: ______________________

Date:_______________ Per:________

DIRECTIONS: Answer the following in the space provided.

Define the following:

1. Collision Theory: *theory that relates the nature and frequencies of collisions in a chemical system to the reaction process – particles must collide in order to react*
2. Activation Energy: *the energy required to reach the transition state and form an activated complex*
3. Activated Complex: *a high energy assemblage of atoms shifting through various configurations on the way from reactants to products at the peak energy point of a chemical reaction*
4. Transition State: *the point in a chemical reaction where the reactant particles have collided and are about to form new products*
5. Reaction Mechanism: *series of elementary steps that occur in a chemical reaction*
6. Intermediate (Product): *product formed in one elementary step and consumed as a reactant in another elementary step*
   a. Elementary steps = individual steps within an overall net chemical reaction
7. Multistep Mechanism: *a reaction pathway consisting of two or more elementary steps*
8. Reaction Rate: rate at which reactants are turned into products. This is expressed by using the concentration (molarity) of reactants or products changed in a given amount of time (s). Units of reaction rate = *M*/s.
9. Rate Law: *an equation that relates reaction rate and concentrations of reactants - must be determined through experimentation*
10. Rate-Determining Step: *the slowest elementary step in a reaction mechanism (also called a rate-limiting step*
11. Reactant Order: *the value of the exponent applied to the reactant concentration in a rate law*
12. Reaction Order: *sum of the reactant order (exponents) in a rate law*
13. Unimolecular: *an elementary reaction in which a single compound/molecule acts as the only reactant*
14. Bimolecular: *an elementary reaction in which two compounds/molecules collide as reactants*
15. [ ]: *symbol used to represent concentration in moles/liter – may be applied to a value as in [0.10] = 0.10M or a formula as in [NaCl] = concentration of NaCl in moles/liter*
16. Describe collision theory: *The nature and frequency of collisions dictates whether and how quickly chemical reactions occur. Anything that increases the number of collisions will react the reaction rate.*
17. List the two things that make a collision "effective".
    a. *Appropriate orientation (angle of impact)*
    b. *Sufficient energy (kinetic) to cause electrons to shift*

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18. There are 5 factors that affect the rate of reaction. List each and describe why/how they affect the rate.

Factor	***Why/How it Affects Reaction Rate***
a. Nature of eactants	*Unstable bonds break easily – stable bonds require more energy* • *Ionic Reactants – usually fast reacting at room temp.* • *Covalent reactants – usually slow reacting at room temp.*
b. Temperature	*Faster moving particles = more frequent collisions.* *Faster particles = more energetic collisions*
c. Concentration	*Higher concentration = more frequent collisions* *Smaller distances between particles*
d. Surface Area	*More surface area = more frequent collisions*
e. Use of Catalysts	*Substances that speed up a chemical reaction without being permanently altered.* *Catalysts may:* • *lower the activation energy of one or more elementary steps in a reaction mechanism reaction.* • *allow for a different reaction pathway to occur*

DIRECTIONS: Calculate the average rate of reaction for the following.

19. A reaction occurs in which the concentration of Reactant A is reduced from 6.0 *M* to 3.2 *M* in 35.0 seconds.

$$\text{Average rate of reaction} = \frac{\Delta[\text{reactant or product}]}{\Delta\text{time}}$$

$$\text{Rate of consumption of A} = \frac{3.2M - 6.0M}{35s} = 0.080M/s$$

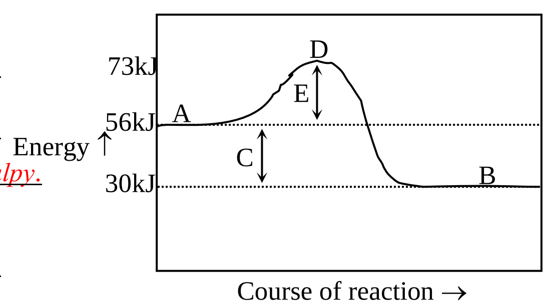
You can ignore the "–", because the rate is described as one of "consumption".

20. The following reaction occurs: A + B → C. Reactant B was made by dissolving 1.2 moles of solute to make 2.3L of aqueous solution. At the end of the reaction, there are only .3 moles remaining in the solution. The reaction required 30 seconds to complete.

Step 1 - Initial concentration of B = 1.2 mol / 2.3 L = 0.52M
Step 2 - Final concentration of B = 0.3 mol / 2.3 L = 0.13M
Step 3 - Δ concentration of B = 0.13M – 0.52M = -0.39M
Step 4 - Rate of consumption = 0.39M/30s = 0.13M/s*

You can ignore the "–", because the rate is described as one of "consumption".

DIRECTIONS: Use the diagram to answer the following.



21. The activated complex of the reaction occurs at what point? *D*

22. What does *C* represent? *The change in energy stored in the substances, (ΔH), or change in enthalpy.*

23. *E* represents the *Activation energy*

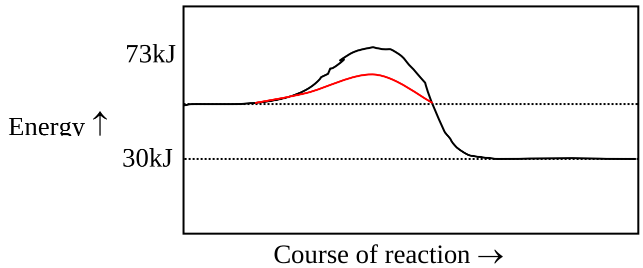
24. The activation energy is 17kJ = (73kJ – 56kJ)

25. The energy change (ΔH) in the system is -26kJ = ($\Delta H = H_{products} - H_{reactants}$) = (30kJ – 56kJ)

26. What occurs as the products are formed? *Energy is released as atoms form bonds and new compounds are made.*

27. Why must activation energy be present for a chemical reaction to occur? *Activation energy supplies the energy in collisions needed to break existing bonds.*

28. How would the use of a catalyst appear on this graph? ____*A catalyst would lower the activation energy of the reaction for one or more elementary steps*



DIRECTIONS: Answer the following in the space provided.

29. The initial rate of the following reaction in aqueous solution is monitored by measuring the number of moles of Hg_2Cl_2 that precipitate per liter per minute. The data obtained are listed in the table.

$2\ HgCl_2(aq) + C_2O_4^{2-}(aq) \rightarrow 2\ Cl^-(aq) + 2\ CO_2(g) + Hg_2Cl_2(s)$

Trial	$[HgCl_2]_0$	$[C_2O_4^{2-}]_0$	Initial rate [M/(L·min)]
1	0.105	0.15	1.8×10^{-5}
2	0.105	0.15	1.8×10^{-5} 2x
3	0.052 2x	0.30 2x	7.1×10^{-5} 8x
4	0.052	0.15	8.9×10^{-6}

a. Determine the rate law for the reaction using the initial rates method.

Rate = $k[HgCl_2][C_2O_4^{2-}]^3$

b. What is the value of the rate constant *k*?

Select one trial and substitute the rate & concentrations into the rate law. Using Trial 1:

$$k = \frac{Rate}{[HgCl_2]\,[C_2O_4^{2-}]^3} = \frac{1.8\times10^{-5}}{[0.105][0.15]^3} = 0.05079$$

0.051

c. Are all four trials necessary to answer parts a. & b.? Explain. *No. The first two trials are identical.*

30. One method for the destruction of ozone in the upper atmosphere is:

$O_3 + \sout{NO} \rightarrow \sout{NO_2} + O_2$ (slow)

$\sout{NO_2} + O \rightarrow \sout{NO} + O_2$ (fast)

$O_3 + O \rightarrow 2O_2$

a. Write the equation for the overall reaction:

b. Which species is an intermediate? NO_2

c. Which species is a catalyst? NO

d. Which is the rate-determining step? *First*

e. Write the rate law for the reaction: *rate = $k[O_3][NO]$*

31. A reaction has the rate law,

$$\text{rate} = k[A]^3[B][C]^2$$

a. What happens to the reaction rate if the concentration of A is doubled? *Increased by a factor of 8x. (2^3)*

b. What happens to the reaction rate if the concentration of B is doubled? *Increased by a factor of 2x. (2^1)*

c. What happens to the reaction rate if the concentration of C is doubled? *Increased by a factor of 4x. (2^2)*

d. What happens to the reaction rate if the concentration of A is halved? *Reduced by a factor of 1/8x. [($\frac{1}{2}$)^3]*

e. What happens to the reaction rate if the concentration of C is tripled? *Increased by a factor of 9x. (3^2)*

f. What happens to the reaction rate if the concentration of C is halved? *Reduced by a factor of 1/4x. [($\frac{1}{2}$)^2]*