Name Key (note typo on #15) Date

Unit 8.2 Study Guide and ANYTHING FROM PRACTICE PROBLEMS (2/7)

Reaction Rates

1. Write a balanced equation for the reaction between nitrogen gas and hydrogen to yield ammonia gas.

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

2. Write the rate expression for the reaction in number 1.

$$\text{Rate} = -\frac{1}{1}\frac{\Delta[N_2]}{\Delta t} = -\frac{1}{3}\frac{\Delta[H_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$$

3. According to the following **unbalanced** reaction, if the rate of appearance of oxygen gas is 4.00×10^{-2} M/s, what is the rate of disappearance of $KClO_3$ (g)?

2 $KClO_3(g) \rightarrow$ 2 $KCl(g)$ + 3 $O_2(g)$

a. 2.67×10^{-2} M/s
b. 6.00×10^{-2} M/s
c. 2.00×10^{-2} M/s
d. 3.00×10^{-2} M/s
e. 5.33×10^{-2} M/s

$$-\frac{1}{2}\frac{\Delta[KClO_3]}{\Delta t} = \frac{1}{3}\frac{\Delta[O_2]}{\Delta t}$$

$$-\frac{\Delta[KClO_3]}{\Delta t} = \frac{2}{3}(4.00E\text{-}2\ M/s) = -0.0267\ M/s$$

Rate laws

4. What is the general rate law for this reaction? $H_3BO_3 + 3\ HCl \rightarrow BCl_3 + 3\ H_2O$

$$\text{rate} = k[H_3BO_3]^x[HCl]^y$$

5. What is the general rate law for a zero order overall reaction? rate = k

6. Name at least one factor that can increase the success of reactant collisions. Explain **how** it increases the reaction rate.

Temp: increase temp, reactants move faster, more collisions, faster rate
Conc: increase conc., more reactants, more collisions, faster rate

7. The rate constant for the reaction $HNO_3 + NH_3 \rightarrow NH_4NO_3$ is 14.5 1/M·s If the concentration of nitric acid is 0.050 M and the concentration of ammonia is 0.10 M and both reactants are first order, what will the rate of this reaction be? $\text{Rate} = k[HNO_3][NH_3]$

$$\text{Rate} = (14.5\ M^{-1}s^{-1})(0.05)(0.10)$$
$$= 0.073\ M/s$$

8. The rate law for the following reaction is rate = $k[H_2][F_2]$. If the rate is 3.15×10^{-4} M/s when $[H_2] = 0.084$ M and $[F_2] = 0.25$ M, calculate the rate when $[H_2] = 0.039$ M and $[F_2] = 0.099$ M. (Hint: find the rate constant k)

$$3.15E\text{-}4\ M/s = k(0.084)(0.25) \qquad k = 0.015\ M^{-1}s^{-1}$$

$$\text{rate} = 0.015\ M^{-1}s^{-1}(0.039)(0.099) = 5.8E\text{-}5\ M/s$$

Initial Rates / Rate constant (k)

9. If the concentration of a reactant is doubled and the corresponding rate quadruples, what is the order with respect to that reactant? 2nd

10. What are the units for the rate constant if the rate law is: rate = k[A][B]²[C]?
    a. 1/s
    b. 1/M*s
    c. 1/M²*s
    d. M/s
    e. 1/M³s

11. When two compounds, A and B, are mixed together, they form compound C, by a reaction that's not well understood. Fortunately, the following rate information was experimentally determined, as shown below:

Experiment	[A] (mol/L)	[B] (mol/L)	Rate (mol/L·sec)
1	0.050	0.050	4.0×10^{-3}
2	0.10	0.050	8.0×10^{-3}
3	0.050	0.10	1.6×10^{-2}

a) Determine the rate law for this reaction.

Ex 1 / Ex 2 $\frac{.05}{.10} = \frac{1}{2}$ $\frac{4.0E-3}{8E-3} = \frac{1}{2}$ Ex 1 / Ex 3 $\frac{0.05}{0.10} = \frac{1}{2}$ $\frac{4.0E-3}{1.6E-2} = \frac{1}{4}$ Rate = k[A][B]²

b) Determine the rate constant for this reaction.

$4.0E^{-3}\ M/s = k\,[0.05]\,[0.05]^2$

$= k\,[0.000125]$

$k = 32\ \frac{1}{M^2 s}$

Integrated Rate Laws

12. To graphically find the rate constant for a first order reaction you should plot ln[A] versus time.
    a. For a second order reaction? $\frac{1}{[A]}$
    b. For a zero order reaction? [A]

13. For a chemical reaction A + 2 B → C, a plot of $1/[A]_t$ versus time *t* is found to give a straight line with a positive slope.
    a) What is the order of the reaction? 2nd
    b) How could you determine the rate constant *k* for the reaction from your graph?
    Calculate the slope $\left(\frac{y-y}{x-x}\right)$

Half life

14. What is the half-life for a first order reaction if the initial concentration of reactant is 1.25M and after 69.2 seconds the concentration has dropped to 0.955M?

$\ln(.955/1.25) = -k(69.2s)$ $\quad k = 3.89E^{-3}\ s^{-1}$

$t_{1/2} = \frac{0.693}{3.89E^{-3}s^{-1}} = 178s$

15. What is the reactant concentration after 78.9 seconds for a ~~first~~ second order reaction with a half-life of 3.10 minutes if the initial concentration was 0.555M?

$3.1\ \text{min} \times \frac{60\text{s}}{1\ \text{min}} = 186\text{s} = \frac{1}{k[0.555]}$ $\quad t_{1/2} = \frac{1}{k[A_0]}$

$k = 9.687\text{E}{-3}\ M^{-1}s^{-1}$ $\quad \frac{1}{A} = 9.687\text{E}{-3}\ M^{-1}s^{-1}(78.9) + \frac{1}{0.555}$

$A = 0.390\ M$

16. A researcher at GCC is running a new chemical reaction that obeys first order kinetics and discovers that after 24 hours only ½ of the reactants are turned into products. How long will it take in hours for 90% of the reactants to be reacted? (Hint: how much is remaining when 90% is reacted?)

$t_{1/2} = 24\text{ h} = \frac{0.693}{k}$ $\quad k = 0.02888\ h^{-1}$

$\ln(10/100) = -0.02888\ h^{-1} t$, $\quad t = 8.0\text{E}1$ hours

17. A pesticide decomposes following *first-order* kinetics.

(a) If the half-life of the pesticide is 12 years, what is the rate constant *k* for the decomposition reaction? $12 = \frac{0.693}{k}$ $\quad k = 0.0578\ yr^{-1}$

(b) What fraction of the pesticide will be left after 36 years?

$\frac{36}{12} = 3$ half-lives $\quad \left(\frac{1}{2}\right)^3 = 0.125 = \frac{1}{8} = 12.5\%$

(c) What fraction of the pesticide will be left after 100 years?

$\frac{100}{12} = 8.33$ half lives $\quad \left(\frac{1}{2}\right)^{8.33} = 0.0031 = .31\%$

(d) How many years will it take for 99.9% of the pesticide to decompose?

$100 - 99.9\% = 0.1\%$ left $= 0.001$ $\quad \frac{1}{2}^x = 0.001$ $\quad \frac{\ln 0.001}{\ln 0.5} = 9.97$ half lives $\times 12 =$ **120 yrs**

Mechanisms

18. The slow step in a mechanism is also called the rate determining step.

19. If an elementary step has two reactants, what is its molecularity? bimolecular

20. For this reaction: $2\ NO(g) + Cl_2(g) \rightarrow 2\ NOCl(g)$ the **real** rate law is rate = *k* [NO] $[Cl_2]$
If the following steps are the mechanism, which one must be the rate determining step? 1

a. What is the intermediate in this mechanism? $NOCl_2$

b. What is the overall order? 2

$NO(g) + Cl_2(g) \rightarrow NOCl_2(g)$ step one
$NO(g) + NOCl_2(g) \rightarrow 2NOCl(g)$ step two

21. For the following reaction, the rate law is found to be Rate = $k[Ce^{4+}][Mn^{2+}]$.

$2\ Ce^{4+} + Tl^{+} \rightarrow 2\ Ce^{3+} + Tl^{3+}$

One mechanism for this reaction, containing the following elementary steps, is shown below:

1. $Ce^{4+} + Mn^{2+} \text{---> } Ce^{3+} + Mn^{3+}$
2. $Ce^{4+} + Mn^{3+} \text{---> } Ce^{3+} + Mn^{4+}$
3. $Tl^{+} + Mn^{4+} \text{---> } Tl^{3+} + Mn^{2+}$

a. There is no catalyst and no intermediate.
b. There is no catalyst, but Mn^{2+} is the intermediate.
c. Mn^{3+} is the catalyst and there is no intermediate.
d. Mn^{2+} is the catalyst and Mn^{3+} and Mn^{4+} are the intermediates.
e. Mn^{4+} is the catalyst and Mn^{2+} and Mn^{3+} are the intermediates.

22. Which of the following equations represents the rate law for the following elementary process:

$A + B \rightarrow C + D$?

a. Rate = k[C][D]
b. Rate = k[A]
c. Rate = $k[A][B]^2$
d. Rate = k[A][B]
e. Rate = k[B]

23. Below is a possible mechanism for the oxidation of bromide ions by hydrogen peroxide in aqueous acid solution. What is the overall reaction equation for this process?

$H^{+} + H_2O_2 \rightarrow H_2O^{+} + OH^{-}$ (rapid equilibrium)

$H_2O^{+} + OH^{-} + 2Br^{-} \rightarrow HOBr + H_2O$ (slow)

$HOBr + H^{+} \rightarrow Br_2 + H_2O$ (fast)

a. $2\ H^{+} + H_2O_2 + Br^{-} + HOBr \rightarrow H_2O^{+} + OH^{-} + Br_2 + H_2O$
b. $H_2O^{+} + OH^{-} + H^{+} + Br^{-} \rightarrow Br_2 + H_2O$
c. $2\ H^{+} + H_2O_2 + 2\ Br^{-} \rightarrow Br_2 + 2\ H_2O$
d. $2\ H_2O^{+} + OH^{-} + 2\ Br^{-} \rightarrow H_2O_2 + Br_2 + 2\ H_2O$

Catalysis

24. A catalyst increases the rate of a reaction by

a. increasing the enthalpy of the reaction
b. decreasing the enthalpy of the reaction
c. lowering the activation energy of the reaction
d. raising the activation energy of the reaction

25. A catalyst is effective because

a. it supplies energy to the reactant molecules, allowing more of them to achieve energies in excess of the activation energy for the reaction
b. it increases the temperature of the molecules in the reaction mixture
c. it increases the number of collisions between molecules
d. it lowers the activation energy of the reaction by providing a lower energy mechanism or pathway