****** ANSWER KEY ***** **Unit 12 (Chapter 17) Review Worksheet – Reaction Rates & Equilibrium**

Directions: *Complete the following questions. Show all work on a separate piece of paper!!!*

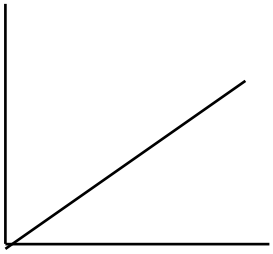
Target #1: Describe how to measure and calculate reaction rates.

1. How can rates of reaction be expressed? How would you expect the concentration of a reactant to change during a reaction? How would you expect the concentration of a product to change during a reaction?

Rxn. Rate = amt. of product formed or amt. of reactant used / Time interval

Concentration of reactants **decrease because it gets used up.**
Concentration of products **increase because it is formed.**

2. Sketch a graph of a reaction:



a. At a constant rate

b. Rate that is getting faster

c. Rate that is slowing down and eventually stopping

3. Describe how you could measure the rate of the following reaction:
$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$

Collect $H_2(g)$ produced, measure mL of gas per unit time (interval).
Collect $H_2(g)$ produced, measure pressure of gas per unit time (interval).
Zn(s), mass changed during rxn per unit time.

4. For a reaction, the initial concentration of a reactant is 5.0M. After 1 minute the concentration is 3.0M and after 2 minutes, the concentration of the reactant is 2.0M. What is the average rate of the reaction for the first minute? Second minute? Overall?

Time (min)	Concentration (M)
0	5.0M
1	3.0M
2	2.0M

1st min: $\frac{5.0M - 3.0M}{1\ min. - 0\ min.}$ = 2 M/min.
2nd min: $\frac{3.0M - 2.0M}{2\ min. - 1 min.}$ = 1 M/min.
Overall: $\frac{5.0M - 2.0M}{2\ min. - 0\ min.}$ = 1.5 M/min.

5. a. Using the graph below, calculate the rate of the reaction between the second and the fifth minute.

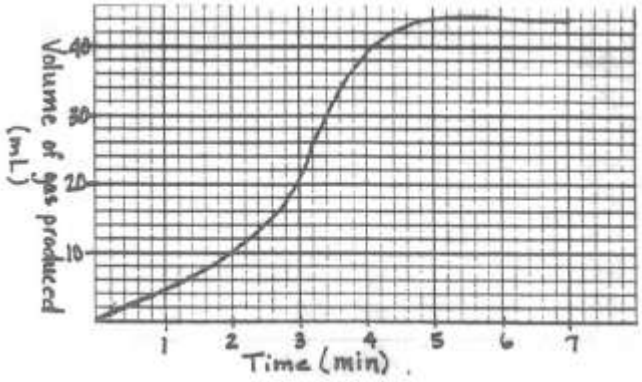
Rate = slope = $\frac{44mL - 10mL}{5min - 2min}$ = 11.3 mL/min.

b. When is the rate of the reaction the greatest?

Slope was steepest = 3-4 min. time interval

c. When does the reaction stop?

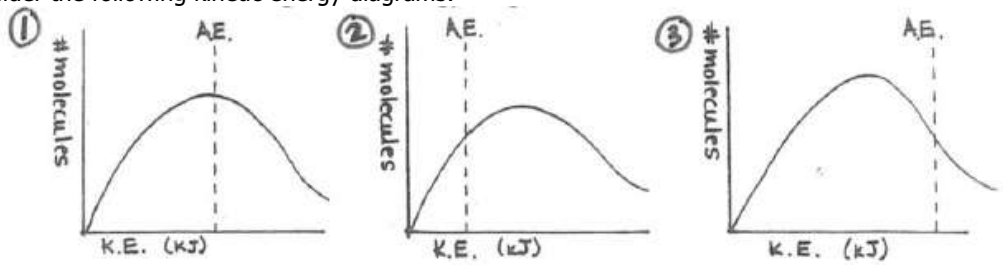
When slope = 0, rate = 0 = reaction is over. 5 min.



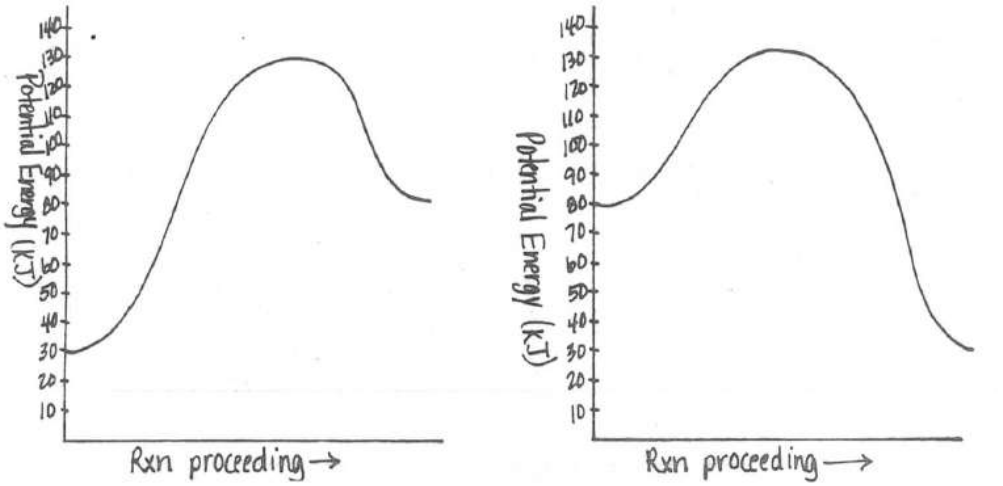
Target #2: Understand activation energy and interpret energy graphs.

6. Define the following terms:
   a. Activation energy - **(E_a) minimum amount of energy required for a reaction to occur.**
   b. Activated complex - **high energy arrangement of reactant molecules where the molecules have enough energy to continue the reaction.**
   c. Catalyst - **substance that speeds up a chemical reaction without itself being consumed.**
7. $N_2 + 3H_2 \rightarrow 2NH_3 + 92kJ$
   a. Is this reaction exothermic or endothermic? **exothermic**
   b. Sketch a potential energy diagram.

   c. How might the graph look if a catalyst is added to this reaction? **Dashed line – a catalyst lowers E_a needed**
8. Consider the following kinetic energy diagrams:



   a. Which reaction would be the fastest at room temperature? **2, more molecules have more KE past E_a**
   b. Which reaction is least likely to occur at room temperature? **3, few molecules have KE required for reaction to occur (E_a)**
   c. Which reaction will be the most affected by an increase in temperature? **3, ↑T = ↑KE for molecules = ↑ # molecules having E_a (E_a does NOT change)**
   d. Draw a new line(T2) on each graph representing the reaction at a higher temperature.
   e. Draw a new line(T3) on each graph representing the reaction at a lower temperature.
   f. How will a catalyst affect each of the above? Draw a dotted line to represent the effect of a catalyst**. (Lowers E_a)**
9. For each potential energy graph below:



   d. Label: reactants, products, activated complex, activation energy, change in enthalpy (ΔH)
   e. Find the activation energy for the forward and the reverse reaction.
      #1 = E_a = 130kJ- 30 kJ = 100kJ E_a reverse = 130-80kJ = 50kJ
      #2 = E_a = 130kJ – 80kJ = 50kJ E_a reverse = 130kJ – 30kJ = 100kJ
   f. Find the ΔH for the forward and the reverse reaction
      ΔH = H products – H reactants
      #1: ΔH = 80kJ – 30kJ = 50kJ forward ΔH = 30-80kJ = -50kJ reverse
      #2: ΔH = 30kJ – 80kJ = -50kJ forward ΔH = 80-30kJ = 50 kJ reverse
   g. Label each as exothermic or endothermic.
   h. Show the effects of a catalyst.
   i. How are these two graphs related? **Reverse reactions**

Target #3: Explain factors that change reaction rates.

10. What factors affect the rate of a chemical reaction? Explain each factor on the molecular level
    1. **Surface area. More surface area = more places where molecules are able to make contact (collisions) = increased rate.**
    2. **Concentration. Increased concentration = more molecules in solution = more molecules available for collisions.**
    3. **Nature of reactants. Chemical make-up of reactants (reactivity trends or # of bonds to break and form.)**
    4. **Temperature. Increased temperature = increase in the # of molecules in solution = increase in collisions = increase in rate**
    5. **Catalyst. Increase rate by lowering activation energy so more molecules have enough energy to react.**

11. How could you increase the rate of the reaction in #4? Explain your answer on the molecular level.
    - **Grind Zn(s) into a powder which increases the surface area which increases the chance of collisions, which increases the rate.**
    - **Increase the concentration of H_2SO_4(aq) which increases the # of molecules in solution, which increases the collisions which increases the rate.**
    - **Increase the temperature, which increases the Kinetic energy of the molecules, which increases the velocity, which increases the collisions, which increases the rate.**
    - **Add a catalyst, which lowers the activation energy so that more molecules have energy to react.**

12. Why is an explosion possible when a large amount of dry, powdered, combustible material is distributed as dust in the air?
    There is a very high surface area = more likely molecules will come in contact with $O_2$9g) and any spark will set off a rapid combustion reaction.

13. Consider the following separate trials for reacting aluminum and hydrochloric acid. Which reaction would most likely be the most rapid? Why?
    a. 1 gram of powdered aluminum reacts with 20mL of 0.5M HCl
    b. 1 gram chunk of aluminum reacts with 20mL of 0.5M HCl
    c. 1 gram of powdered aluminum reacts with 20mL of 1.5M HCl
    d. 1 gram chunk of aluminum reacts with 20mL of 1.5M HCl

    c. Al(s) is powdered = increased surface area and higher concentration of HCl(aq)

Target #4: Identify rate determining step from a reaction mechanism.

14. Define the following terms:
    a. Reaction mechanism - **sequence of steps in a chemical reaction.**
    b. Rate-determining step - **slowest reaction (step) in a reaction mechanism, which determines the rate of the overall reaction.**

15. Consider the following reaction mechanism:

$A + C \rightarrow AC$	slow
$AC + B \rightarrow AB + C$	fast
$AB + B \rightarrow AB_2$	fast

    a. Which reaction is the rate-determining step? **1st step, A + C → AC because it is the slowest step**
    b. How will doubling the concentration of A affect the rate of the overall reaction? Why?
       ↑rate, since A is a reactant in the rate determining step (slowest) = more molecules = ↑ chance of collisions = ↑rate
    c. How will doubling the concentration of B affect the rate of the overall reaction? Why?
       No change, since B is a reactant in the fast steps, speeding them up does not change the rate of the overall reaction.

16. In general, how could you increase the rate of a reaction mechanism? **Speed up the rate determining (slow) step by changing one of the factors listed in question #10.**

Target #5: Write the equilibrium expression & utilize it to calculate the equilibrium constant for a chemical rxn.

17. What is chemical equilibrium? The rate at which reactants forming products is equal to the rate at which products are forming reactants. **A closed system – no observable changes.**

18. What is equal in equilibrium? **Rate of the forward reaction = rate of reverse reaction.**
19. Why is a solid not included as part of the equilibrium expression?
    Solid concentrations are constant (does not change) because their densities do not change.

20. What is indicated by a large value for the equilibrium constant? A small value?
    Large K = products are favored (more products than reactants at ⇆)
    Small K = reactants are favored (more reactants than products at ⇆)
21. Would you expect the value for the equilibrium constant for the dissolving reaction for salt in water to be large or small? Explain.
    K would be large, because NaCl dissolves easily in H_2Om so many ions (products) are formed.

22. In the following reaction, which is at equilibrium, $[N_2] = 0.045M$, $[O_2] = 0.075M$, and $[N_2O_5] = 0.0028M$. Calculate K_{eq}.
    $2N_2\ (g) + 5O_2\ (g) \leftrightarrows 2N_2O_5\ (g) +$ energy
    $$K = \frac{[N_2O_5]^2}{[N_2]^2[O_2]^5} = \frac{[.0028]^2}{[.045]^2[.075]^5} = 1600$$

23. At equilibrium, at 25°C, $[H^+] = 1.0 \times 10^{-5}$ and $[CH_3COOH] = 1.0M$. If the equilibrium constant at 25°C is equal to 1.8×10^{-5}, find $[CH_3COO^{-1}]$. $CH_3COOH\ (aq) \leftrightarrows H^+\ (aq) + CH_3COO^{-1}\ (aq)$

    $$K = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} \qquad 1.8 \times 10^{-5} = \frac{[1.0 \times 10^{-5}][CH_3COO^-]}{[1.0]} \qquad [CH_3COO^-] = 1.8M$$

24. Consider the following equilibrium system:
    $Ag^+(aq) + 2NH_3(aq) \leftrightarrows Ag(NH_3)_2^+(aq) +$ energy
    a. What is the equilibrium constant expression for this reaction?
    $$K_{eq} = \frac{[Ag(NH_3)_2^+]}{[Ag^+][NH_3]^2}$$
    b. If the equilibrium concentrations are as follows, what is the value of K_{eq}?
    $[Ag^+] = 2.0M$, $[NH_3] = 6.0\ M$, $[Ag(NH_3)_2^+] = 8.0M$
    $$K_{eq} = \frac{[8.0]}{[2.0][6.0]^2} = .11$$
    c. According to LeChatelier's Principle, what is the result of adding more $NH_3(aq)$ to this system? (Assume all other factors are present.)
    NH_3 is a reactant, so adding more of a reactant will shift the ↔ to make more products , $Ag(NH_3)_2^+$.
    d. According to LeChatelier's Principle, what is the result of lowering the temperature of this system?
    Lowering temperature, the reaction will produce more energy to counteract taking energy away, so ⇆ will shift toward the energy (exo); in this case, shift to make more products.
    e. How will the change in part d affect the value of K_{eq}?
    K_{eq} will get larger $\frac{[Prod]}{[React]}$ ← top # getting larger = larger value
    f. What happens if a catalyst is added to this system?
    No effect to ⇆ (rate forward and rate reverse both speed up)
    g. What is the equilibrium constant for the reverse reaction if the conditions are identical to those in part b above?
    $$K_{eq} = \frac{[Ag^+][NH_3]^2}{[Ag(NH_3)_2^+]} = \frac{[2.0][6.0]^2}{[8.0]} = 9.0$$

Target #6: Predict the changes when a system at equilibrium is disturbed (Le Chatelier's principle).

25. List the factors that affect equilibrium. Give an example of each and explain HOW they affect equilibrium.
    1. Concentration (gases and aqueous solutions): **increase [reactant] = ⇆ shift to products**
       increase [product] = ⇆ shift to reactants
       decrease [reactant] = ⇆ shift to reactants
       decrease [product] = ⇆ shift to products
    2. Temperature: **increase Temp = ⇆ shift away from the energy**
       decrease Temp = ⇆ shift toward the energy
    3. Pressure and Volume (gases): **Increase P or decrease V = ⇆ shift to the side that takes up the least space (if all gases, least # of moles)**
       Decrease P or increase V = ⇆ shift to the side that takes up the most space (if all gases, most # of moles)

Consider the following reactions for questions 28 – 31.

a. $4PH_3\ (g) \leftrightarrows P_4\ (g) + 6H_2\ (g)$ $\Delta H = 1184$ kJ
b. $2Ag\ (s) + Cu^{+2}\ (aq) \leftrightarrows Cu\ (s) + 2Ag^{+}\ (aq)$
c. $FeO\ (s) + CO\ (g) \leftrightarrows Fe\ (s) + CO_2\ (g)$
d. $CN^{-}\ (aq) + H_2O\ (l) \leftrightarrows HCN\ (aq) + OH^{-}\ (aq)$
e. $4Fe\ (s) + 3O_2\ (g) \leftrightarrows 2Fe_2O_3\ (s)$ $\Delta H = -1674$ kJ

26. Write equilibrium expressions for each of the above reactions.
    a. $\mathbf{K_{eq} = \frac{[P_4][H_2]^6}{[PH_3]^4}}$
    b. $\mathbf{K_{eq} = \frac{[Ag^{+}]^2}{[Cu^{+2}]}}$
    c. $\mathbf{K_{eq} = \frac{[CO_2]}{[CO]}}$
    d. $\mathbf{K_{eq} = \frac{[HCN][\ OH^{-}]}{[CN^{-}]}}$
    e. $\mathbf{K_{eq} = \frac{1}{[O_2]^3}}$

27. How will an increase in pressure affect each of the above equilibrium systems?
    Increase P = shift to side that takes up least amt. of space
    a. ← b. no effect c. no effect d. no effect e. →
28. How will an increase in temperature affect reactions a and e? How will K_{eq} be affected?
    Increase T = reaction favors endothermic reaction or opposite energy, to use the excess E
    a. →, K increases ([prod]↑)
    b. ←, K decreases (because [react]↑)
29. Describe the affects of each of the following changes.
    a. Adding more Ag^{+} to system b. ← **make more reactants since a product was added (use up product)**
    b. Adding more PH_3 to system a. → **make more products since a reactant was added (use up reactant)**
    c. Decreasing the temperature for system e. → **since energy was taken away, the rxn produces more energy (exothermic), shifts toward energy.**
    d. Adding more water to system d. **no effect- pure liquid's concentration does not change**
    e. Adding more FeO to system c. **no effect – pure solid's concentration does not change.**
    f. Removing CO_2 as it forms from system c. → **take away a product, the rxn will make more**
    g. Adding a solution of sodium chloride to system b. **$Cl^{-}(aq) + Ag^{+}(aq) \rightarrow AgCl(s)$ a ppt sp $[Ag^{+}]$ decreases so rxn shifts → to make more Ag^{+} to replace that was taken away.**

30. How would the equilibrium $[N_2]$ be affected by the following changes.
    Use I for increase, D for decrease, and R for remains the same.
    $2N_2\ (g) + 5O_2\ (g) \leftrightarrows 2N_2O_5\ (g) +$ energy
    a. increase volume **(I)**
    b. increase $[N_2O_5]$ **(I)**
    c. decrease $[N_2O_5]$ **(D)**
    d. decrease temperature **(D)**
    e. add a catalyst **(R)**
    f. add O_2 **(D)**