Accel. Unit 10- Rates and Equilibrium Name: ________________________ Period: ______

Unit 10 Test Date: _________

- **All homework is due the day of the unit test.**
- **All quizzes and quiz retakes are due the day of the unit test.**
- **All labs are due the day of the unit test.**

Additional Resources Available at:

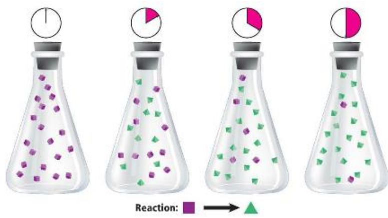
- **www.blendedaccelchem.weebly.com OR www.accelwarriorchem.weebly.com**

Guided Notes: Reaction Kinetics and Collision Theory

Expressing Reaction Rates

- Some chemical reactions are ________________ and others are __________________, but chemists need to be more ____________________.
- What is a rate?
- How do we use rates in everyday life?

- How would we measure the rate of a reaction?
- Equation for rate



- What happens to the amount of reactants over time? __________________________
- What happens to the amount of products over time? ___________________________
- Do you think you would observe the same changes in reactants and products for every reaction? Explain.

Reaction Rate

- Reaction rate for chemistry is defined as: ____
  ____
- Concentration: ____
  - solute: ______________________________________
  - solvent: ______________________________________
  - ex: salt in water, salt is the _______________________, water is the ________________________
  - unit typically used for concentration in chemistry: _________________,which means: ______________
- Reaction rates are determined _________________________ by measuring the __________________________ of the reactants and/or products in a ________________ ___________________.
- Reaction rates CANNOT by calculated from a ____.
- Reaction rates must always be _______________________________.

Collision Theory

- In order for a reaction to occur:
  - reactants must _________________
  - collisions must be in the ______________________________________
  - collisions must have a ____ for bonds to break
- Activated Complex: a temporary, unstable arrangement of atoms in which ______________________________ and ____.
- ___________________________ is another name for activated complex.
- Collisions with the correct orientation must also have a sufficient amount of _____________________.
- This amount of energy is called the ____.
- Symbol: _______
- How would a high vs. a low activation energy affect the speed of a reaction?

Activation Energy

- Reaction #1:

- Reaction #2:

- Which graph is exothermic? ______ How do you know?
- Which graph is endothermic? ______ How do you know?
- Which graph has a higher activation energy? ______
- Which reaction in the graphs will be faster? Explain.

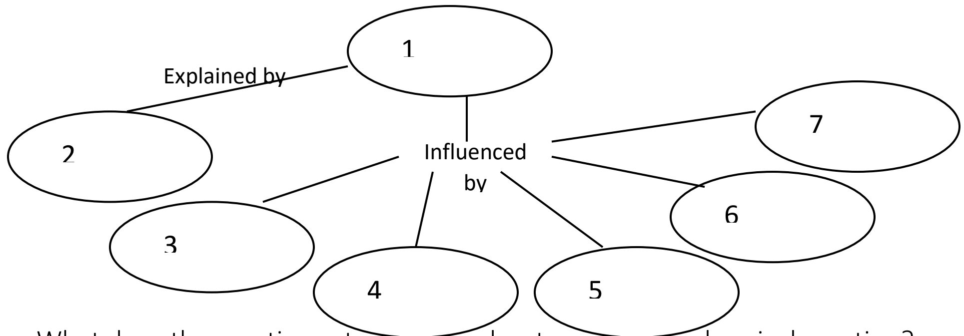
Factor Affecting Reaction Rate:

- ________________________________
- ________________________________
- _________________________________
- _________________________________
- _________________________________
    - substance that _______________________ the rate of reaction ________________________________.
    - creates a lower _________________________________
    - ____
    - Sketch:

Kinetics (Reaction Rate) Practice

1. Complete the following concept map using the following terms:

Surface area	*Temperature*	*Concentration*	*Catalyst*
Collision theory	*Reaction rates*	*Reactivity*	



2. Define reaction rate. What does the reaction rate indicate about a particular chemical reaction?

3. In addition to colliding, what else must happen in order for a reaction to occur?

4. Use the collision theory to discuss how the following factors affect the rate of a chemical reaction:
   a. Temperature

   b. Concentration

   c. Surface area

5. What role does the reactivity of the reactants play in determining the rate of a chemical reaction?

6. *Answer the following questions about catalysts:*
   a. What is the difference between a homogeneous and a heterogeneous catalyst?

   b. How does a catalyst affect the activation energy for a chemical reaction?

   c. What is the result of adding a catalyst to a reaction?

7. Would the changes listed below increase or decrease the rate of the following reaction:

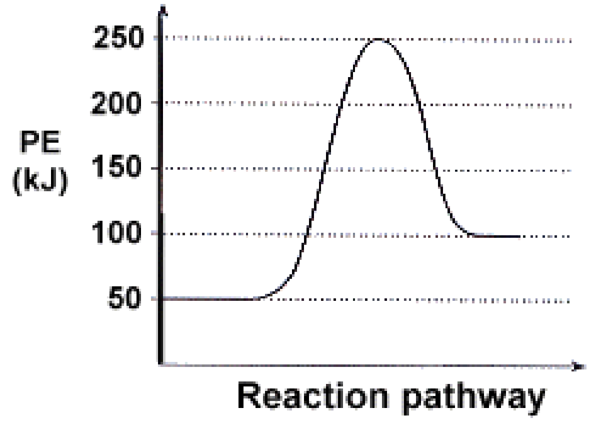
$$I_2\,(s) + Cl_2\,(g) \rightarrow 2ICl\,(g)$$

   a. decreasing temperature ________________________ c. crushing I_2 ____________________________

   b. Increasing $[Cl_2]$ ________________________ d. adding a catalyst ____________________________

Activation Energy Diagrams

Use the graph below to answer questions 1-7: Include labels on any numerical values.



1. Label the position of the ***reactants*** on the graph.
2. Label the position of the ***products*** on the graph.
3. Lable the position of the ***activated complex*** on the graph.
4. How much energy do the reactants have at the start of the reaction? ______________________________
5. What is the activation energy for this reaction? ____ Label this on the graph.
6. How much energy do the products have at the end of the reaction? ____________________________________
7. Is this reaction exothermic or endothermic? Explain your answer using evidence from the graph.

8. Draw an energy diagram on the axes below using the given information. Be sure to include labels and units on both the x-axis and y-axis.

   Potential energy of reactants = 350 kJ/mole
   Activation energy = 100 kJ/mole
   Potential energy of products = 250 kJ/mole

9. Is this reaction exothermic or endothermic? Explain your answer using evidence from the graph.

10. You add a catalyst to the reaction you graphed in question 8, which lowers the activation energy of the reaction from 100 kJ/mole to 50 kJ/mole. Draw the energy diagram of the catalyzed reaction on the same set of axes above (use a dashed line or a different color and label the reaction with the catalyst).

Guided Notes: Rate Laws

Review: Molarity

- measures the ___________________________
- solute is measured in _______________
- solution is measured in _______________
- abbreviated with a capital ______

Practice:

1. What is the molarity of a solution that has 10 grams of sodium sulfate in 100 mL of solution?

Rate Laws:

- increased concentration of a ______________________ usually ______________________ the rate of a reaction
- however, _________________ concentration might actually have little effect on the rate of __________________

Rate Order and Rate Laws:

- For the reaction A + B → C + D
- General form of Rate Law:

$$rate = k[A]^x[B]^y$$

- rate laws are found ______________________
    - change the concentration of ___________________________ at a time to see how the rates are affected
- Rate units: M/s (change in molarity per second)

Rate Law Example #1:

Reaction: A + B → C

Trial	[A]	[B]	Rate (M/sec)
1	1.0	2.0	0.50
2	2.0	2.0	1.00
3	2.0	6.0	3.00

1. What happens to the rate when [A] doubles?
2. What is the rate order of reactant A?
3. What happens to the rate when B triples?
4. What is the rate order of reactant B?
5. What is the rate law for this reaction?

Rate Law Example #2:

Reaction: A → B + C

Trial	[A]	Rate (M/sec)
1	2.5	1.00
2	5.0	4.00
3	7.5	16.00

1. What happens to the rate when [A] doubles?

2. What is the rate order of reactant A?

3. What is the rate law for this reaction?

Rate Law Example #3:

Reaction: A + B → C

Trial	[A]	[B]	Rate (M/sec)
1	2.0	4.0	3.0
2	6.0	2.0	1.5
3	6.0	4.0	3.0

1. What happens to the rate when [A] triples?

2. What is the rate order of reactant A?

3. What happens to the rate when [B] doubles?

4. What is the rate order of reactant B?

5. What is the rate law for this reaction?

Rate Laws Practice

General form of a rate law is **rate = $k[A]^x[B]^y$** for a reaction with two reactants

1. Use the data table below to answer questions about the reaction **$A_2 + B_2 \rightarrow 2\ AB$**

Trial	$[A_2]$	$[B_2]$	Rate (M/s)
1	0.01	0.05	0.01
2	0.01	0.10	0.02
3	0.02	0.10	0.04

a. What trials do you use to determine the effect of $[A_2]$ on the reaction rate?

b. What is the rate order (the exponent) with respect to $[A_2]$?

c. What trials do you use to determine the effect of $[B_2]$ on the reaction rate?

d. What is the rate order (the exponent) with respect to $[B_2]$?

e. What is the rate law for this reaction?

2. Use the data table below to answer questions about the reaction **$C + D \rightarrow E$**

Trial	[C]	[D]	Rate (M/s)
1	0.1	0.01	0.02
2	0.1	0.02	0.04
3	0.2	0.02	0.16

a. What trials do you use to determine the effect of [C] on the reaction rate?

b. What is the rate order (the exponent) with respect to [C]?

c. What trials do you use to determine the effect of [D] on the reaction rate?

d. What is the rate order (the exponent) with respect to [D]?

e. What is the rate law for this reaction?

3. Use the data table below to answer questions about the reaction **C + D → E**

Trial	[C]	[D]	Rate (M/s)
1	0.1	0.01	0.02
2	0.1	0.02	0.08
3	0.1	0.03	0.18
4	0.1	0.04	0.32
5	0.2	0.04	1.28
6	0.3	0.04	2.88

a. What trials do you use to determine the effect of [C] on the reaction rate?
b. What is the rate order (the exponent) with respect to [C]?

c. What trials do you use to determine the effect of [D] on the reaction rate?
d. What is the rate order (the exponent) with respect to [D]?

e. What is the rate law for this reaction?

4. Use the data table below to answer questions about the reaction **F + G → H**

Trial	[F]	[G]	Rate (M/s)
1	0.01	0.4	0.02
2	0.02	0.4	0.16
3	0.03	0.4	0.54
4	0.1	0.2	5
5	0.1	0.4	20
6	0.1	0.6	45

a. What trials do you use to determine the effect of [F] on the reaction rate?
b. What is the rate order (the exponent) with respect to [F]?

c. What trials do you use to determine the effect of [G] on the reaction rate?

d. What is the rate order (the exponent) with respect to [G]?

e. What is the rate law for this reaction?

Reaction Order and Rate Law Expression Worksheet #2

1. Reaction: C + D $\rightarrow$ E

Exp #	[C]	[D]	Rate (mole dm^{-3} s^{-1})
1	0.1	0.01	0.02
2	0.1	0.02	0.04
3	0.1	0.03	0.06
4	0.1	0.04	0.08
5	0.2	0.04	0.08
6	0.3	0.04	0.08

   a. What is the rate order of reactant C?

   b. What is the rate order of reactant D?

   c. What is the rate law for the reaction?

2. Reaction: F + G $\rightarrow$ H

Exp #	[F]	[G]	Rate (mole dm^{-3} s^{-1})
1	0.01	0.4	0.02
2	0.02	0.4	0.04
3	0.03	0.4	0.06
4	0.1	0.2	5
5	0.1	0.4	10
6	0.1	0.6	15

   a. What is the rate order of reactant F?

   b. What is the rate order of reactant G?

   c. What is the rate law for the reaction?

3. Reaction: C + D $\rightarrow$ E

   a. What is the rate order of reactant C?

   b. What is the rate order of reactant D?

   c. What is the rate law for the reaction?

Exp #	[C]	[D]	Rate (mole dm^{-3} s^{-1})
1	0.1	0.01	0.02
2	0.1	0.02	0.08
3	0.1	0.03	0.18
4	0.1	0.04	0.32
5	0.2	0.04	1.28
6	0.3	0.04	2.88

4. Reaction: F + G $\rightarrow$ H

Exp #	[F]	[G]	Rate (mole dm^{-3} s^{-1})
1	0.01	0.4	0.02
2	0.02	0.4	0.16
3	0.03	0.4	0.54
4	0.1	0.2	5
5	0.1	0.4	20
6	0.1	0.6	45

a. What is the rate order of reactant F?

b. What is the rate order of reactant G?

c. What is the rate law for the reaction?

5. Reaction: A_2 + B_2 $\rightarrow$ 2 AB

Exp #	[A_2]	[B_2]	Rate (mole dm^{-3} s^{-1})
1	0.001	0.001	0.01
2	0.001	0.002	0.02
3	0.001	0.003	0.03
4	0.001	0.004	0.04
5	0.002	0.004	0.16
6	0.003	0.004	0.36

a. What is the rate order of reactant A_2?

b. What is the rate order of reactant B_2?

c. What is the rate law for the reaction?

Guided Notes: Keq

2 Types of Reactions:

1. Completion Reactions:
   a. Results in a complete ______________________ of _________________________ to ____________________
      i. Example:
   b. 2 indicators of a completion reaction are formation of a _____________ or formation of a ______
   c. most reactions ____ ______ go to completion
   d. have a _________________ sided arrow in its equation
2. Reversible Reactions:
   a. Can occur in both the ___________ and ___________ directions
   b. Example:
      i. equations have a _________ sided arrow
   c. forward arrow indicates ____
      i.
   d. reverse arrow indicates ____
      i.
   e. both reactions will ____

Chemical Equilibrium:

A state in which the _____________________ and _____________________reactions take place at _____________ rates.

___________________________________ = ____

- The amount of the __________________________ and _________________________ are _____________ at equilibrium
- Equilibrium is _____________ - reactions are still occurring, even though we may not be able to see it
  - Sketch the graph below and describe what is being shown about the concentrations of the substances.

- What happens to the forward rate as it approaches equilibrium?
- What happens to the reverse rate as it approaches equilibrium?
- What is true about the forward and the reverse rate at equilibrium?

The Law of Chemical Equilibrium:

- At a given _______________________, a chemical system may reach a state in which a particular _______________ of __________________ and _________________ concentrations has a constant value.
- General example reaction:
  - What do the lower case letters represent?
  - What do the capital letters represent?
- Write the general equilibrium constant expression:
  - [ ] = ____

- Keq:
  - ____
  - ____
  - ____
  - ____
- If Keq > 1
  - ____
  - ____
- if Keq < 1
  - ____
  - ____
  - Which is better for business?

2 <u>Types of Equilibrium:</u>

1. <u>Homogenous equilibrium</u>: ____
2. Heterogeneous equilibrium: ____
   a. if any of the substances in the reactions are ______________ or ___________, leave them **out** of Keq
   b. only use ____________________ and __________________ solutions in the expression for Keq

Example 1: Write the equilibrium expression for the following equation:

$N_2O_{4\,(g)} \leftrightarrow 2NO_{2\,(g)}$

Keq =

Example 2: Write the equilibrium expression for the following equation:

$C_{(s)} + H_2O_{(g)} \leftrightarrow CO_{(g)} + H_{2\,(g)}$

Keq =

Example 3: Calculate Keq for the reaction below when $[SO_3]$=0.0160M, $[SO_2]$=0.00560M, and $[O_2]$=0.0210M. Are the products or the reactants favored?

$2SO_{3(g)} \leftrightarrow 2SO_{2(g)} + O_{2(g)}$

<u>Check for Understanding</u>**:** Determine the value of Keq at 400K for the decomposition of phosphorous pentachloride if:

$[PCl_5]$ = 0.135M , $[PCl_3]$ = 0.550M, and $[Cl_2]$=0.550 M.

$PCl_5(g) \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$

Equilibrium Constant (K_{eq}) - Chemistry

1. **Write the equilibrium constant (K_{eq}) expressions for the following homogeneous equilibria.**
   a. $C_2H_4O\ (g) \leftrightarrow CH_4\ (g) + CO\ (g)$
   b. $3O_2\ (g) \leftrightarrow 2O_3\ (g)$
   c. $2N_2O\ (g) + O_2\ (g) \leftrightarrow 4NO\ (g)$
   d. $4NH_3\ (g) + 3O_2\ (g) \leftrightarrow 2N_2\ (g) + 6H_2O\ (g)$
   e. $4HCl\ (g) + O_2\ (g) \leftrightarrow 2Cl_2\ (g) + 2H_2O\ (g)$
   f. $PCl_5\ (g) \leftrightarrow PCl_3\ (g) + Cl_2\ (g)$

2. **Write the equilibrium constant (K_{eq}) expressions for the following heterogeneous equilibria.**
   a. $C_4H_{10}\ (l) \leftrightarrow C_4H_{10}\ (g)$
   b. $NH_4HS\ (s) \leftrightarrow NH_3\ (g) + H_2S\ (g)$
   c. $CO\ (g) + Fe_3O_4\ (s) \leftrightarrow CO_2\ (g) + 3FeO\ (s)$
   d. $(NH_4)_2CO_3(s) \leftrightarrow 2NH_3(g) + CO_2(g) + H_2O(g)$

For the following problems, show _all_ of your work including set-up (with K_{eq} expression) and answer with units if needed.

3. At 773 K, the reaction $\mathbf{2NO\ (g) + O_2\ (g) \leftrightarrow 2NO_2\ (g)}$ produces the following concentrations: $[NO] = 3.49 \times 10^{-4}$ M; $[O_2] = 0.80$ M; $[NO_2] = 0.25$ M. Calculate the equilibrium constant (K_{eq}) for this reaction.

4. The chemical equation for the decomposition of formamide is: $\mathbf{HCONH_2\ (g) \leftrightarrow NH_3\ (g) + CO\ (g)}$ Calculate K_{eq} using the following equilibrium data: $[HCONH_2] = 0.0637$ M, $[NH_3] = 0.518$ M and $[CO] = 0.518$ M.

5. Calculate K_{eq} for the reaction for iron and water if the equilibrium concentrations are as follows: $[H_2O] = 1.00$ M & $[H_2] = 4.50$ M. **$2Fe\ (s) + 3H_2O\ (g) \leftrightarrow Fe_2O_3\ (s) + 3H_2\ (g)$**

6. At 793 K, the equilibrium constant for the reaction **$NCl_3\ (g) + Cl_2\ (g) \leftrightarrow NCl_5\ (g)$** is 39.3.
   a. Do the products or the reactants dominate in this equilibrium? Explain your answer in complete sentences.

   b. If the equilibrium constant for this reaction were less than 1, would the reactants or products be dominant? Explain your answer in complete sentences.

7. The equilibrium constant is 9.36 for the following reaction: **$A\ (g) + 3B\ (g) \leftrightarrow 2C\ (g)$.** The table below provides concentration data for two different reaction mixtures of these gases.

	A (mol/L)	**B (mol/L)**	**C (mol/L)**
Mixture 1	0.716	0.208	0.425
Mixture 2	0.562	0.491	0.789

   a. Calculate the K_{eq} for each mixture. Use the back of the sheet to show your work.

   b. Are both reactions at equilibrium? Explain your answer in complete sentences.

Guided Notes: Le'Chatelier's Principle

Background Knowledge:

1. What happens if you are running on a treadmill and someone increases the speed?

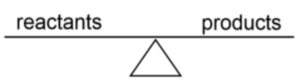
2. What happens if you are riding your bike and the wind picks up?

-- These are ________________________________ being put on you.
--Chemists put ________________________________ on chemical reactions.
Why do chemists want to put stresses on chemical reactions?
--Chemists put stresses on chemical reactions to produce more ________________________________.
--________________________ chemists use this.

Le'Chatelier's Principle: If a ____________________ is applied to a system at __________________________, the system shifts in the direction that relieves the ______________________.

Changes in Concentration:
Adding Reactants

1. What will happen to the balance if you add more reactants?



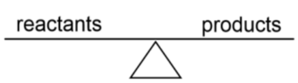
2. What happens if I add more CO?

$$CO(g) + 3H_2(g) \leftrightarrow CH_4 + H_2O$$

3. The reaction will shift to the __________________.

Removing Products

1. What will happen to the balance if you remove products?



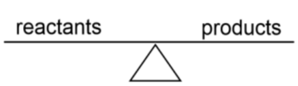
2. What happens if I remove H_2O?

$$CO(g) + 3H_2(g) \leftrightarrow CH_4 + H_2O$$

3. The reaction will shift to the ___________________.

Adding Products

1. What will happen to the balance if you add products?



2. What happens if I add H_2O?

$$CO(g) + 3H_2(g) \leftrightarrow CH_4 + H_2O$$

3. The reaction will shift to the _____________________.

Changes in Volume and Pressure:

Decreasing the Volume

1. What happens to the pressure when volume is decreasing? _________________
2. What happens to the number of collisions? _________________
3. To determine if the reaction will shift, we need to look at the number of __________________ of the reactants and products.

$$\mathbf{CO(g) + 3H_2(g) \leftrightarrow CH_4 + H_2O}$$

4. Which side of the reaction contains more moles? _______________
5. Volume only has an effect on the reaction if the ____ of reactants differs from the number of products.
6. This reaction has more moles of ______________________________, so the reaction will shift to the ________.

Changes in Temperature

1. Alters both the ____ and the ______________________________.
2. Think of heat as either a ____________________________ or ___________________________.
3. Is this an exothermic or an endothermic reaction? _________________________________
4. Is heat considered a product or reactant in the reaction below? ________________________

$$\mathbf{CO(g) + 3H_2(g) \leftrightarrow CH_4 + H_2O + heat}$$

5. In this reaction, adding more heat would shift the reaction to the ___________________________.

Addition of a Catalyst

1. _________________________ up a reaction, but does so in both ways.
2. __________________________ is just reached _____________________.

Summary: **Le'Chatelier's Principle: Changes in ________________________, ____________________________, and _______________________________ make a difference in the amount of product formed in a reaction.**

Practice:

For the reaction below, which change will cause the reaction to shift to the right?

$$\mathbf{CH_4(g) + 2H_2S(g) + heat <\!\!-\!\!-\!\!-\!\!> CS_2(g) + 4H_2(g)}$$

a. decrease the concentration of dihydrogen sulfide

b. increase the pressure on the system

c. increase the temperature on the system

d. increase the concentration of carbon disulfide

e. decrease the concentration of methane

Guided Notes: Ksp

Solubility Product Constant – K_{sp}

- K_{sp}:
  - General Equation:
- Since the reactant is ALWAYS a _______________, K_{sp} = ____
  - b and c are the _____________ on the ions
- The ***smaller*** K_{sp} is the ______ soluble salt
- K_{sp} can be used to calculate the ________________________ of ________________.

Practice - K_{sp}

1. Write the K_{sp} expression for the solvation of Ag_2SO_4.

First, determine the ions that will be formed:

Put the ions in the K_{sp} expression (must include charges!):

Use the coefficients to determine how many moles of each ion will be formed. Put those numbers in for b & c (as exponents):

(if the exponent is ____________________, it is not used in the expression)

2. Write the K_{sp} expression for the solvation of magnesium hydroxide. Formula: _______________________

3. Write the Ksp expression for the solvation of calcium phosphate. Formula: ______________________

4. Calculate the solubility for AgCl at 298K. (Ksp=1.8×10^{-10}).

    a. Write the Ksp expression for the solvation of AgCl.

    b. Simply the expression.

    c. Fill in your Ksp value and solve.

5. Calculate the solubility for $CaCO_3$ at 298K. (Ksp=3.4×10^{-9}).

6. Calculate the Ksp for $PbCl_2$ with a solubility of 5.0×10^{-3}mol/L.

    a. Write the Ksp expression for the solvation of $PbCl_2$.

    b. Simplify the expression.

    c. Fill in your “x” (molar concentration) and solve.

7. Calculate the Ksp for $BaCrO_4$ with a solubility of 1.5×10^{-5}mol/L.

Ksp Practice Problems

Use your notes or read the portion of "Using solubility product constants" (pg.614-617). Pay attention to any and all examples!

1. What is the solubility product constant and when is it used?

2. How can you calculate ion concentration using the solubility product constant?

3. Write the Ksp expression for the following compounds:
   a. PbF_2
   b. $Zn(OH)_2$
   c. $MgCO_3$

4. Use the Ksp values from the table to calculate the following: (Show all of your work)
   a. The solubility in mol/L of $PbCrO_4$.
   b. The solubility in mol/L of Ag_2SO_4.
   c. $[F^-]$ in a saturated solution of CaF_2 at equilibrium.

Compound	**Ksp at 298 K**
$PbCrO_4$	2.3×10^{-13}
Ag_2SO_4	1.2×10^{-5}
CaF_2	3.5×10^{-11}

Unit 10 Review – Accelerated Chemistry

1. What is a reaction rate and what units are used with reaction rates?

2. What is the collision theory?

3. List the factors that affect the rate of a reaction. Explain how each factor affects the rate.

A

B

C

D

E

4. Draw a reaction diagram for an exothermic reaction and label the following: reactants, products, activation energy, activated complex.

5. For the reaction 3 ClO^- (aq) → ClO_3^- (aq) + 2 Cl^-(aq) doubling the concentration of ClO^- quadruples the initial rate of formation of ClO_3^- . What is the rate expression for the reaction?

6. The reaction $C_6H_5N_2Cl$ (aq) + H_2O (l) → C_6H_5OH (aq) + N_2 (g) + HCl (aq) is first order in [$C_6H_5N_2Cl$] and zero order in [H_2O]. What is the rate expression?

7. For the reaction A + B →AB , the following data was obtained:
   a. Write the rate expression for the reaction.

Trial	Initial [A]	Initial [B]	Initial Rate mol/L*min
1	0.480 M	0.190 M	0.350
2	0.480 M	0.380 M	0.350
3	0.240 M	0.190 M	0.087

8. What 2 factors will drive a reaction to completion?

a) ________________________________

b) ________________________________

9. Describe a reversible reaction. Give an example.

10. Describe dynamic equilibrium. Give an example.

11. At equilibrium how do the forward and reverse reaction rates compare? The forward rate __________ the reverse rate.

12. State Le Chatelier's Principle.

13. What are the 3 possible stresses we can apply to a system at equilibrium?

a) ___________________________ b) __________________________ c) _______________________________

14. Use the reaction ($2SO_{2(g)} + O_{2(g)} \leftrightarrow 2SO_{3(g)}$ + heat) to determine what will happen (shift left/right, no change) if the following stresses are applied:

a. SO_2 is added _______________ b. Volume is increased ____________ c. Heat is added ________________

15. What is the general formula for the equilibrium constant, K_{eq}?

16. What does the value of K_{eq} tell a chemist about a reaction if....
If the value of K_{eq} is greater than 1 ......

If the value of K_{eq} is less than 1 ....

17. Write the equilibrium constants for these reversible reactions – ALL CHEMICALS ARE GASES:

a. $2A + B \leftrightarrow C + 3D$ b. $NO + O_2 \leftrightarrow NO_3$ c. $CO_2 + H_2 \leftrightarrow CO + H_2O$

18. Calculate K_{eq} for reaction **17a** if the equilibrium concentrations are: [A]=0.100M, [B]=0.230M, [C]=1.17M, & [D]=2.19M.

19. The equilibrium constant in **17b** is .025. If [NO] = .36 M and [O_2] = .21 M, what is the equilibrium concentration of NO_3?

20. If K_{eq} in **17c** is 6.37×10^{-3}, [CO_2] = 0.037M, [H_2] = 0.28M, and [CO] = 0.084M, calculate [H_2O].

21. Describe K_{sp}.

22. What is the generic formula for K_{sp}? ______________________________

23. Write the expression for K_{sp} for the following sparingly soluble salts:
$PbBr_2$ __________________________________ $Ca_3(PO_4)_2$ _______________________________

24. Calculate the Ksp of $CaSO_4$ if a saturated solution has a concentration of 1.58×10^{-4}.

25. The solubility product constant of BaCO3 is 2.6×10^{-9}. Calculate the solubility (in mol/L) of $BaCO_3$.

26. The solubility product constant of Ag2CrO4 is 1.1×10^{-12}. Calculate the [Ag+] in a solution of Ag_2CrO_4 at equilibrium.

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