

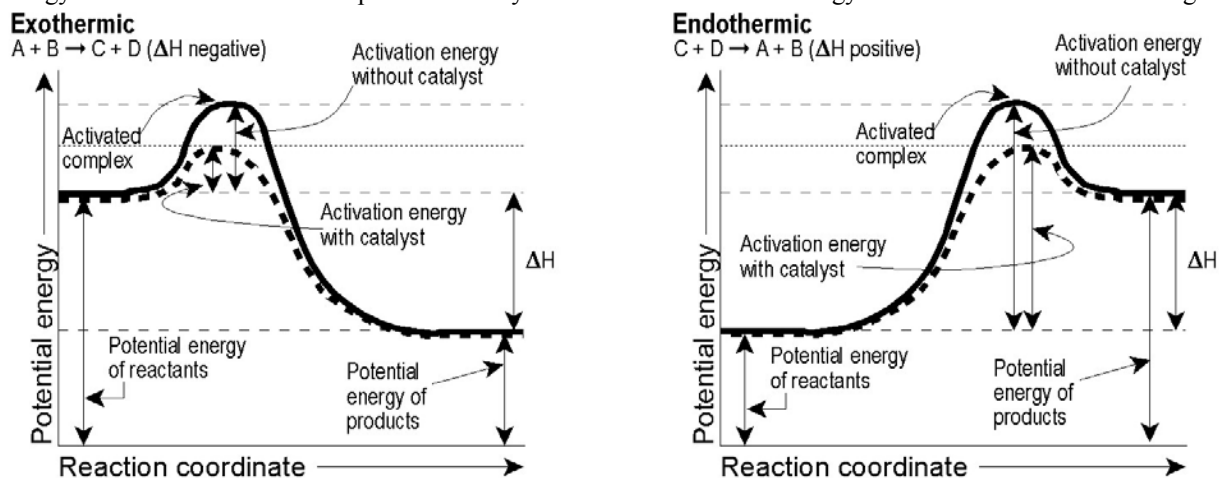
## Test Review № 9

**Collision theory.** In order for a reaction to occur, particles of the reactant must collide. Not all collisions cause reactions. An effective collision is one in which the colliding particles approach each other at the proper angle and with the proper amount of energy to cause a reaction. The greater the rate of effective collisions is, the greater the reaction rate is.

**Reaction mechanisms.** Effective collisions between more than two particles at a time are rare. If all the particles shown on the reactant side of a balanced equation had to collide for a reaction to occur, the reaction would not take place. Chemical reactions occur by a series of intermediate steps between the initial reactants and final products. Each step probably involves a collision of only two particles. The series of steps that lead from reactants to products is called a **reaction mechanism**. The slowest step of the reaction mechanism is called the **rate determining step**. Increasing the concentration of the reactant(s) that enter the rate determining step increases the reaction rate increasing the concentration of only reactants not involved in the rate determining step has little effect on the reaction rate

**Transition state theory.** During a chemical reaction, intermediate products form that exist for only brief periods of time while the atoms rearrange themselves. These intermediate products have high energy because they are formed by high energy collisions. High energy intermediate products are unstable and break apart to form the final product(s). A high energy intermediate product is called an **activated complex** or a **transition state complex**. The energy needed to form the activated complex is the **activation energy**.

**Role of energy in reactions.** In order for a reaction to begin, energy is needed. The energy needed to begin the reaction is the activation energy. Activation energy comes from effective collisions. During a chemical reaction, heat may be released or absorbed. Heat released or absorbed during a chemical reaction is called heat of reaction or enthalpy difference ( $\Delta H$ ). The enthalpy difference is the difference between the potential energy of the products and the reactants. [ $\Delta H = H_{\text{products}} - H_{\text{reactants}}$ ]. Exothermic reactions are reactions in which energy is released. This means the potential energy of the products is lower than the potential energy of the reactants and  $\Delta H$  is negative. Endothermic reactions are reactions in which energy is absorbed. This means the potential energy of the products is higher than the potential energy of the reactants and  $\Delta H$  is positive. Catalysts reduce the activation energy but have no effect on the change in enthalpy.



**Entropy.** Entropy is randomness or disorder. In nature, high entropy is favored, yet many things that occur naturally are quite organized. This is because nature also favors low potential energy or **enthalpy**. Entropy and enthalpy sometimes conflict with each other. In order to determine if a reaction will occur when only one of the two factors is favored by nature, it is necessary to determine the free energy.

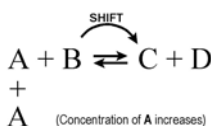
**Gibbs Free Energy Change.** Whether a reaction proceeds spontaneously or not depends on the balance between the drive toward greater stability (reduced potential energy); and the drive toward less organization (increased entropy). The Gibbs free energy change ( $\Delta G$ ) predicts whether or not a reaction is spontaneous. It is the difference between the energy change ( $\Delta H$ ) and the product of the absolute or Kelvin temperature ( $T$ ) and the entropy change ( $\Delta S$ ).

$$\Delta G = \Delta H - T\Delta S$$

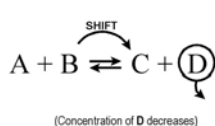
**Reaction rates.** The speed of chemical reactions depends on several factors: [1] *Nature of reactants.* Chemical reactions occur by breaking and rearranging existing bonds. The fewer electrons that need to be rearranged, the faster the reaction is. As a result, reactions between ionic substances in aqueous solution, such as double replacement reactions, are rapid. Reactions in which covalent bonds are broken, such as decomposition of hydrogen peroxide, occur slowly at room temperature. [2] *Concentration of reactants.* An increase in concentration results in an increase in the frequency of collisions. Usually as the concentration increases, the reaction rate increases, however, if the concentration of only the reactants that are NOT involved in the rate determining step is increased, the number of collisions are increased without effecting the reaction rate. Increasing pressure increases the concentration only for gases. [3] *Surface area.* Increasing the surface area of reactants increases the reaction rate by increasing the opportunity for collisions. [4] *Temperature.* As temperature increases so does the reaction rate by increasing both the frequency and effectiveness of collisions. [5] *Catalysts.* Catalysts speed up reactions without being permanently altered by changing the reaction mechanism so less activation energy is required.

**Equilibrium.** Some reactions are reversible. When the forward and reverse reaction rates are equal in reversible systems the system is at equilibrium. It is a dynamic equilibrium because the concentration of the reactants and the products remain constant despite the continuation of both the forward and reverse reactions. The concentrations of reactants and products at equilibrium can be quite different. Equilibrium can be attained from either the forward or the reverse reaction. There are several types of equilibrium: [1] *Phase Equilibrium.* Since phase changes are reversible, in closed systems a state of equilibrium between phases can be reached. When solids or liquids are confined in a container, equilibrium will be reached when there are enough particles of gas (vapor) to cause the rate of return to the original phase to be equal to the rate of escape. This causes the characteristic vapor pressure of a substance at a given temperature. [2] *Solution equilibrium.* Equilibrium can be reached between the gas dissolved in a liquid and the undissolved gas above it. Low temperature and high pressure favor solution of gases. Solid added to a saturated solution will fall to the bottom instead of dissolving. Equilibrium is reached between the dissolved solute and the undissolved solute when the rate of dissolving equals the rate of crystallization. [3] *Chemical equilibrium.* Equilibrium occurs when forward and reverse chemical reactions occur at equal rates.

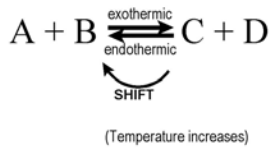
**Le Chatelier's principle.** When stress is applied to a system in equilibrium, the reaction will shift in a direction that relieves the stress and a new equilibrium will be established. Applied stresses include changes in concentration, pressure, or temperature.



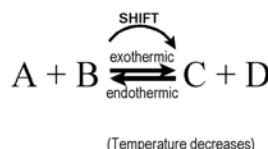
Shift due to increase  
in concentration



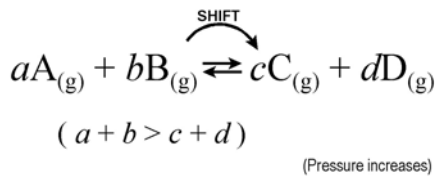
Shift due to decrease  
in concentration



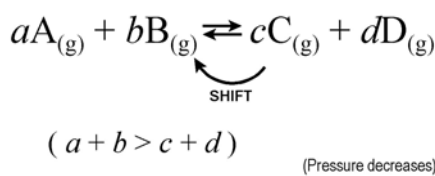
Shift due to increase in  
temperature



Shift due to decrease in  
temperature

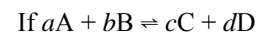


Shift due to increase in pressure



Shift due to decrease in pressure

**Law of Chemical Equilibrium.** The law of chemical equilibrium is shown in the box to the right. There are several things that can be done to interpret equilibrium using the law. First, since the numerator of the fraction is based on the concentration of the products, an equilibrium constant greater than one ( $k_{eq} > 1$ ) refers to a reaction that favors the formation of product, while an equilibrium constant less than one ( $k_{eq} < 1$ ) refers to a reaction that favors the formation of the reactants. Second, appropriate equilibrium expressions need to be written base on the balanced equation. And third, mathematical problems can be done substituting values for the concentration into the equilibrium expression, either to determine if equilibrium has been reached, or to determine the equilibrium concentration of one of the reactants or products.



$$k_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

### Sample Problems

The reaction  $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$  is at equilibrium at 1 atm and 298 K. The equilibrium constant for the reaction at 1 atm and 298 K is as follows:  $K_{eq} = 3.5 \times 10^{-1}$

(a) Which is favored, the forward or the reverse reaction?

Answer: The reverse reaction is favored ( $K_{eq} < 1$ )

(b) What is the equilibrium expression?

Answer:

$$3.5 \times 10^{-1} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

(c) If the concentration of hydrogen iodide is  $2.65 \times 10^{-2} \text{ M}$ , the concentration of hydrogen is  $5.0 \times 10^{-3} \text{ M}$ , and the concentration of iodine is  $4.0 \times 10^{-1} \text{ M}$ , is the reaction at equilibrium?

Answer:

$$K_{eq} = \frac{(2.65 \times 10^{-2})^2}{(5.0 \times 10^{-3})(4.0 \times 10^{-1})} = 3.5 \times 10^{-1}$$

Yes!

(d) If the reaction is at equilibrium, what is the concentration of hydrogen iodide if the concentration of hydrogen is  $3.0 \times 10^{-3} \text{ M}$ , and the concentration of iodine is  $2.5 \times 10^{-2} \text{ M}$ ?

Answer:

$$3.5 \times 10^{-1} = \frac{[\text{HI}]^2}{(3.0 \times 10^{-3})(2.5 \times 10^{-2})}$$

$$[\text{HI}] = \sqrt{(3.5 \times 10^{-1})(3.0 \times 10^{-3})(2.5 \times 10^{-2})}$$

$$[\text{HI}] = \sqrt{2.625 \times 10^{-5}} = 5.1 \times 10^{-3} \text{ M}$$

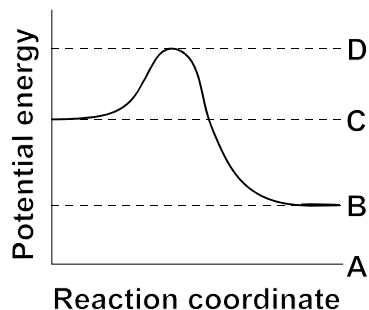
Answer the questions below by circling the number of the correct response

- An increase in temperature increases the rate of chemical reactions. This is primarily because the (1) concentration of the reactants increases, (2) number of effective collisions increases, (3) activation energy increases, (4) average kinetic energy decreases.
- An increase in temperature increases the rate of a chemical reaction because the (1) activation energy increases, (2) activation energy decreases, (3) number of molecular collisions increases, (4) number of molecular collisions decreases.
- The difference between the heat content of the products and the heat content of the reactants is (1) entropy of reaction, (2) free energy, (3) heat of reaction, (4) activation energy.
- The purpose of the catalyst in a reaction is to (1) change the activation energy required of the reaction, (2) provide the energy necessary to start the reaction, (3) increase the amount of product formed, (4) decrease the amount of reactants used
- Given the reaction  $\text{A} + \text{B} \rightleftharpoons \text{AB} + 210 \text{ kJ}$ . If an activation energy of 21 kJ is required, the activation energy of the reverse reaction is (1) 21 kJ, (2) 189 kJ, (3) 210 kJ, (4) 231 kJ.
- The difference between the potential energy of the reactants and the potential energy of the products is (1)  $\Delta G$ , (2)  $\Delta S$ , (3)  $\Delta H$ , (4)  $\Delta T$ .
- When a catalyst is added to a system at equilibrium, there is a decrease in the activation energy of (1) the forward reaction, only, (2) the reverse reaction, only, (3) both the forward and reverse reaction, (4) neither the forward nor the reverse reactions.
- The net effect of a catalyst is to change the (1) potential energy of the reactants, (2) potential energy of the products, (3) heat of reaction, (4) rates of both forward and reverse reactions

## TEST 9 REVIEW

9. Heat of reaction,  $\Delta H$ , is equal to (1)  $H_{\text{products}} + H_{\text{reactants}}$ , (2)  $H_{\text{products}} \times H_{\text{reactants}}$ , (3)  $H_{\text{products}} - H_{\text{reactants}}$ , (4)  $H_{\text{products}}/H_{\text{reactants}}$ .
10. According to transition theory, activation energy is used to (1) make bonds, (2) form an activated complex, (3) break bonds, (4) make the reaction exothermic.
11. For an endothermic reaction, the energy of the activated complex is (1) higher than the energy of the reactants, only, (2) higher than the energy of the products, only, (3) higher than the energy of both the reactants and the products, (4) lower than the energy of both the reactants and the products
12. For a given chemical reaction, the potential energy of the reactants is less than the potential energy of the products. This reaction is (1) endothermic and energy is absorbed, (2) endothermic and energy is given off, (3) exothermic and energy is absorbed, (4) exothermic and energy is given off
13. As a catalyst increases the rate of a reaction, the activation energy of the reaction (1) decreases, (2) increases, (3) remains the same
14. In a chemical reaction, the products have a lower potential energy than the reactants. This reaction must have a negative (1)  $\Delta G$ , (2)  $\Delta H$ , (3)  $\Delta S$ , (4)  $\Delta X$ .

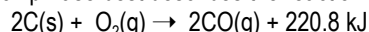
Answer questions 15-19 by referring to the diagram below.



15. The energy of activation for the given reaction is represented by the interval (1) BC (2) AB (3) AD (4) DC
16. The heat of the reactants for the forward reaction is represented by the interval (1) AB (2) AC (3) DC (4) BD
17. The energy of activation for the reverse reaction is represented by the interval (1) AB (2) BC (3) CD (4) BD
18. The energy of activated complex for the forward reaction is represented by (1) AB (2) BC (3) AD (4) BD
19. The heat of the reaction,  $\Delta H$ , is represented by the interval (1) AB (2) BC (3) CD (4) BD
20. A decrease in the rate of the reaction in a gaseous system could be caused by an increase in (1) temperature (2) concentrations (3) pressure (4) volume
21. The function of a catalyst in a reaction is to lower the (1)  $\Delta H$  of the reaction (2) energy of the reactants (3) activation energy required (4) energy of the products
22. If a catalyst is introduced into a chemical system, the heat of reaction ( $\Delta H$ ) for the reaction (1) increases (2) decreases (3) remains the same
23. As the rate of a given reaction increases due to an increase in the concentration of the reactants, the activation energy for that reaction (1) decreases (2) increases (3) remains the same
24. When a catalyst lowers the activation energy, the rate of a reaction (1) decreases (2) increases (3) remains the same
25. The difference between the heat content of the products and the heat content of the reactants is (1) entropy of reaction (2) free energy (3) heat of reaction (4) activation energy
26. Given the reaction  $a + b = ab + 50 \text{ kJ}$ . If an activation energy of 5 kJ is required, the activation energy of the reverse reaction is (1) 5 kJ (2) 45 kJ (3) 50 kJ (4) 55 kJ
27. As a gas is liquefied, the entropy of the gas-liquid system (1) decreases (2) increases (3) remains the same
28. As iodine molecules change from the solid phase to the vapor phase, the entropy of the system (1) decreases (2) increases (3) remains the same
29. When calcium carbonate decomposes according to the equation  $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$ , the entropy of the system (1) decreases (2) increases (3) remains the same
30. In the reaction:  $\text{A}(\text{s}) + \text{B}(\text{s}) \rightarrow \text{C}(\text{g}) + \text{D}(\text{g}) + \text{heat}$ , the entropy of the system (1) increases and the reaction is exothermic (2) increases and the reaction is endothermic (3) decreases and the reaction is exothermic (4) decreases and the reaction is endothermic
31. Which change is accompanied by a decrease in entropy?  
 (1)  $\text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(\text{s})$  (2)  $\text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(\text{g})$   
 (3)  $\text{H}_2\text{O}(\text{s}) \rightarrow \text{H}_2\text{O}(\text{g})$  (4)  $\text{H}_2\text{O}(\text{s}) \rightarrow \text{H}_2\text{O}(\ell)$
32. Which compound has a higher potential energy than the elements from which it is formed? (see standard energies of formation chart)  
 (1) aluminum oxide (s) (2) nitrogen monoxide (g) (3) hydrogen oxide (l) (4) carbon dioxide (g)
33. The least amount of energy would be evolved by the formation of one mole of (see standard energies of formation chart) (1)  $\text{H}_2\text{O}(\text{g})$  (2)  $\text{SO}_2(\text{g})$  (3)  $\text{CO}_2(\text{g})$  (4)  $\text{CO}(\text{g})$
34. As the reactants are converted to product in the reaction  $\text{A}(\text{g}) + \text{B}(\text{g}) \rightarrow \text{C}(\text{s})$ , the entropy of the system (1) decreases (2) increases (3) remains the same

## TEST 9 REVIEW

35. Which phrase best describes the reaction below?



- (1) exothermic with an increase in entropy (2) exothermic with a decrease in entropy (3) endothermic with an increase in entropy (4) endothermic with a decrease in entropy

Based on the information below for questions 36 - 40, are the reactions described spontaneous? (1) *Yes*, (2) *No*, (3) *Likely*, or (4) *Unlikely*

36. An exothermic reaction at high temperature with decreasing entropy

37. An endothermic reaction at high temperature with decreasing entropy

38. An exothermic reaction at low temperature with decreasing entropy

39. An endothermic reaction at low temperature with decreasing entropy

40. An exothermic reaction at high temperature with increasing entropy

41. Which reactants will combine at the slowest rate at room temperature? (all concentrations are 1 M.) (1)  $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq})$  (2)  $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$  (3)  $\text{Fe}^{+2}(\text{aq}) + 2\text{OH}^-(\text{aq})$  (4)  $\text{CH}_3\text{OH}(\text{aq}) + \text{HCOOH}(\text{aq})$

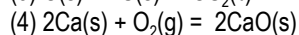
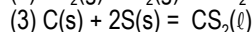
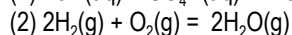
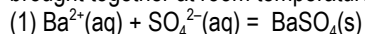
42. A decrease in the rate of the reaction in a gaseous system could be caused by an increase in (1) temperature (2) concentrations (3) pressure (4) volume

43. An increase in the rate of all chemical reactions results from (1) an increase in pressure (2) an increase in temperature (3) a decrease in pressure (4) a decrease in temperature

44. For a given system at equilibrium, lowering the temperature will (1) increase the rate of reaction (2) increase the concentration of products (3) favor the exothermic reaction (4) favor the endothermic reaction

45. The effect of an increase in temperature on the rates of chemical reactions is to increase the rate of (1) endothermic reactions, only (2) exothermic reactions, only (3) both endothermic and exothermic reactions (4) neither endothermic nor exothermic reactions

46. Which reaction proceeds most rapidly when the reactants are brought together at room temperature?



47. Given the system:  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2(\text{g})$ . A catalyst is added to the system while the pressure and temperature are kept constant. Compared to its original concentration, the new equilibrium concentration of the  $\text{NO}_2$  has (1) decreased (2) increased (3) remained the same

48. Given the system:  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2(\text{g})$ . Additional oxygen is added to the system while the temperature and pressure are kept constant. Compared to the original concentration, the new equilibrium concentration of the  $\text{N}_2$  has (1) decreased (2) increased (3) remained the same

49. Given the system:  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2$ . Part of the nitrogen is removed from the system while the pressure and temperature are kept constant. Compared to the original concentration, the new equilibrium concentration of the  $\text{O}_2$  has (1) decreased (2) increased (3) remained the same

50. Given the system:  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2(\text{g})$ . The pressure on the system is decreased while the temperature is kept constant. At the new equilibrium, the number of moles of  $\text{NO}_2$  has (1) decreased (2) increased (3) remained the same

51. Given the system  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2$ . The temperature is decreased and the pressure is kept constant. Compared to its original concentration, the new equilibrium concentration of the  $\text{O}_2$  has (1) decreased (2) increased (3) remained the same

52. Given the system  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2$ . Additional nitrogen is added to the system while the temperature and pressure are kept constant. Compared to the original concentration, the new equilibrium concentration of the  $\text{O}_2$  has (1) decreased (2) increased (3) remained the same

53. Given the system  $\text{N}_2(\text{g}) + 2\text{O}_2(\text{g}) + 67 \text{ kJ} = 2\text{NO}_2$ . Part of the oxygen is removed from the system while the pressure and temperature are kept constant. Compared to the original concentration, the new equilibrium concentration of the  $\text{N}_2$  has (1) decreased (2) increased (3) remained the same

54. In all reversible chemical reactions, equilibrium has been reached when (1) the opposing reactions cease (2) the molar concentrations of the reactants and the products become equal (3) any one substance leaves the field of reaction (4) the rates of the opposing reactions are equal

55. The equation for a saturated solution of an ionic solid is  $\text{AB}(\text{s}) = \text{A}^+(\text{aq}) + \text{B}^-(\text{aq})$ . If the temperature is kept constant and the concentration of the  $\text{A}^+(\text{aq})$  is increased, the concentration of the  $\text{B}^-(\text{aq})$  (1) decreases (2) increases (3) remains the same

56. In the reaction  $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) = 2\text{HCl}(\text{g})$ , if the temperature remains constant and the pressure increases, the amount of  $\text{HCl}$  formed (1) decreases (2) increases (3) remains the same

57. Given the equilibrium reaction:  $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) = 2\text{SO}_3(\text{g}) + \text{heat}$ . If the pressure remains constant and the temperature increases, the amount of  $\text{SO}_3(\text{g})$  (1) decreases (2) increases (3) remains the same

58. Given:  $\text{CH}_3\text{COOH}(\text{aq}) = \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$ . If the pressure on the equilibrium system were increased, then the equilibrium point would (1) shift to the right (2) shift to left (3) remain the same

59. Given the equilibrium system:  $N_2(g) + 3H_2(g) = 2NH_3(g) + 86 \text{ kJ}$ . As the pressure on the system increases, temperature remaining constant, the number of moles of  $NH_3(g)$  produced (1) decreases (2) increases (3) remains the same
60. Given the equilibrium system:  $N_2(g) + 3H_2(g) = 2NH_3(g) + 86 \text{ kJ}$ . As the temperature is increased, pressure remaining constant, the concentration of  $N_2(g)$  (1) decreases (2) increases (3) remains the same
61. In the reaction  $A + B = 2C + \text{heat}$ , the concentration of the product could be increased by (1) decreasing the concentration of B (2) increasing the concentration of A (3) adding more heat to the system (4) adding a catalyst
62. Given the reaction at equilibrium:  $N_2(g) + 3H_2(g) = 2NH_3(g)$ . If the volume of the reaction vessel is decreased at constant temperature, there will be an increase in the number of moles of (1)  $N_2$ , only (2)  $NH_3$ , only (3)  $H_2$ , only (4)  $N_2$ ,  $H_2$ , and  $NH_3$
63. Consider any reaction at equilibrium. If the temperature is decreased and the pressure is kept constant, the rate of the endothermic reaction will (1) decrease (2) increase (3) remain the same
64. When the system  $S + O_2 = SO_2$  is at equilibrium, the concentration of  $SO_2$  will (1) decrease (2) increase (3) remain the same
65. The system  $2SO_2(g) + O_2(g) = 2SO_3(g)$  is in equilibrium. At constant temperature, the equilibrium point will shift to the right if there is (1) a decrease in pressure (2) an increase in pressure (3) a decrease in the concentration of  $O_2(g)$  (4) an increase in the concentration of  $SO_3(g)$
66. When a catalyst is added to a system in equilibrium, the equilibrium concentration of the products (1) increases (2) decreases (3) remains the same
67. For the reaction  $2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g)$  at a constant temperature, K is equal to  
 (1)  $\frac{[H_2]^4[O_2]^2}{[H_2O]}$  (2)  $\frac{[H_2O]^2}{[H_2]^2[O_2]}$  (3)  $\frac{[H_2]^2[O_2]}{[H_2O]^2}$  (4)  $\frac{[H_2O]}{[H_2]^4[O_2]^2}$
68. If  $K_{eq} < 1$ , the reaction (1) favors the formation of the products, (2) favors the formation of the reactants, (3) favors neither the formation of the reactants nor the products, (4) is NOT reversible.

14	2	28	4	56	3
13	1	27	1	55	1
12	1	26	4	54	4
11	3	25	3	53	2
10	2	24	2	52	1
9	3	23	3	51	2
8	4	22	3	50	1
7	3	21	3	49	2
6	3	20	4	48	1
5	4	19	2	47	3
4	1	18	3	46	1
3	3	17	4	45	3
2	3	16	2	44	3
1	2	15	4	43	2

**Answers**

STANDARD ENERGIES OF FORMATION OF COMPOUNDS AT 1 atm AND 298 K		
Compound	Heat (Enthalpy) of Formation * kJ/mol ( $\Delta H_f^\circ$ )	Free Energy of Formation kJ/mol ( $\Delta G_f^\circ$ )
Aluminum oxide $Al_2O_3(s)$	-1674.1	-1580.9
Ammonia $NH_3(g)$	-46.0	-16.3
Barium sulfate $BaSO_4(s)$	-1471.8	-1361.0
Calcium hydroxide $Ca(OH)_2(s)$	-985.2	-897.9
Carbon dioxide $CO_2(g)$	-393.3	-394.2
Carbon monoxide $CO(g)$	-110.4	-137.1
Copper (II) sulfate $CuSO_4(s)$	-770.8	-661.3
Ethane $C_2H_6(g)$	-84.4	-33.0
Ethene (ethylene) $C_2H_4(g)$	52.3	68.1
Ethyne (acetylene) $C_2H_2(g)$	226.6	209.0
Hydrogen fluoride $HF(g)$	-270.9	-273.0
Hydrogen iodide $HI(g)$	26.3	1.7
Iodine chloride $ICl(g)$	18.0	-5.4
Lead (II) oxide $PbO(s)$	-215.3	-188.1
Magnesium oxide $MgO(s)$	-601.1	-568.9
Nitrogen monoxide $NO(g)$	90.3	86.5
Nitrogen dioxide $NO_2(g)$	33.0	51.4
Potassium chloride $KCl(s)$	-436.4	-408.8
Sodium chloride $NaCl(s)$	-410.9	-383.7
Sulfur dioxide $SO_2(g)$	-296.4	-299.7
Water $H_2O(g)$	-241.6	-228.2
Water $H_2O(l)$	-285.5	-237.0

\* Minus sign indicates an exothermic reaction.  
 Sample equations:  
 $2Al(s) + 3/2 O_2(g) \rightarrow Al_2O_3(s) + 1674.1 \text{ kJ}$   
 $2Al(s) + 3/2 O_2(g) \rightarrow Al_2O_3(s) \quad \Delta H = -1674.1 \text{ kJ/mol}$