Review for Unit Test 5: Introduction to Equilibrium Systems (Chapter 7)

Objectives:

1. Define or explain: enthalpy, entropy, reversible reactions, Gibb's Free Energy, spontaneous reactions, equilibrium, steady state, closed system, homogeneous system, heterogeneous system, the equilibrium expression (K_{eq}), and Le Chatelier's Principle.

2. What are the two driving forces in chemical reactions? If you are given a chemical reaction and a value for ΔH, be able to identify whether entropy and enthalpy are increasing or decreasing. Be able to predict reactions which will be non-spontaneous at any temperature, spontaneous at any temperature or are equilibrium reactions.

3. Use Gibb's Free Energy equation to predict whether a reaction is spontaneous at a given temperature. Be able to determine the temperature at which a spontaneous reaction will become non-spontaneous (the "turning point").

4. What criteria must be met in order for a reaction (system) to be at equilibrium?

5. Be able to write K_{eq} expressions for homogeneous and heterogeneous equilibrium systems. Know the relationship between the $K_{forward}$ and $K_{reverse}$ for a reaction.

6. Be able to interpret information about equilibrium from a concentration vs. time graph.

7. Be able to calculate ΔG, K_{eq}, Q and concentrations of products and reactants at equilibrium. Be able to solve equations using ICE tables, perfect squares, the quadratic equation and the "500" rule when K_{eq} is very small.

8. Be able to apply Le Chatelier's principle to homogeneous equilibrium systems to predict the direction that a system will shift in response to various stresses. Go through your notes carefully and be sure that you know the "special cases or rules". Be able to predict the effects of:

- changing the pressure or volume of gaseous systems
- changing temperature
- changing concentration of reactants or products
- adding a catalyst
- adding an inert gas

9. Be able to explain why the above stresses affect (or do not affect) equilibrium systems by referring to the relative rates of the forward and reverse reactions (eg. by increasing the concentration of a reactant, you increase the rate of the forward reaction by increasing the number of collisions. Eq'm is re-established when more product has formed and the rate of the forward and reverse reactions is again equal.)

10. Given concentrations of products and reactants and the value of K_{eq} for a reaction, use Le Chatelier's principle or Q to predict the direction that a system will shift to achieve equilibrium. Use ICE or E'ICE tables to find the concentrations of all species at equilibrium.

11. Be able to <u>**interpret**</u> the meaning of different values of ΔG, K_{eq}, and Q.

Practice Multiple Choices for Equilibrium:

1. Identify the **incorrect** statement below regarding chemical equilibrium:
a) equilibrium is achieved when the forward reaction rate equals the reverse reaction rate
b) equilibrium is achieved when the concentration of species become constant
c) equilibrium is achieved when the reaction quotient Q equals the equilibrium constant
d) equilibrium is achieved when the reactant and product concentrations become equal

2. Which of the following are at chemical equilibrium?
   I) a sealed bottle of liquid bromine has orange bromine vapour above the liquid
   II) the rate of flow of a waterfall is constant
   III) an acetylene torch for welding has a constant flame
   IV) a sealed bottle of champagne is 125 years old

a) all of these are equilibrium systems
b) I and II are equilibrium systems
c) II and III are equilibrium systems
d) I and IV are equilibrium systems

3. Which of the following changes can affect the value of the equilibrium constant?
a) introducing a catalyst
b) changing the temperature
c) changing the concentrations of species
d) changing the pressure inside the reaction vessel

4. For a particular chemical reaction, $\Delta H = +60.0$ kJ and $\Delta S = +121$ J/K. At what temperature (in K) would this reaction become spontaneous?
a) 0.496 K
b) 496 K
c) 273.5 K
d) 2.02 K

5. For the reaction: $Ba(OH)_2(s) + 2\ NH_4NO_3(s) \leftrightarrow Ba(NO_3)_2(aq) + 2\ NH_3(g)$
   $\Delta H = 170.44$ kJ and $\Delta S = 657.4$ J/K
   Calculate ΔG for this reaction at 25ºC.
a) – 25.5 kJ
b) 154 kJ
c) -1.63×10^5 kJ
d) – 366 kJ

6. Hydrogen peroxide decomposes according to the following reaction: $H_2O_2(l) \leftrightarrow H_2O(l) + \frac{1}{2} O_2(g)$
   $\Delta H = -98.2$ kJ and $\Delta S = 70.1$ J/K. At 100ºC, which of the following statements is **true**?
a) ΔG is negative, so the reaction is spontaneous
b) ΔG is negative, so the reaction is non-spontaneous
c) ΔG is positive, so the reaction is spontaneous
d) ΔG is positive, so the reaction is non-spontaneous

7. The reaction that takes place between "Al's Rusty Balls" is: $Fe_2O_3(s) + 2\ Al(s) \leftrightarrow 2\ Fe(s) + Al_2O_3(s)$
   For this reaction, $\Delta H = -851.5$ kJ and $\Delta S = -38.58$ J/K. Which of the following statements is **TRUE** about the process?
a) both driving forces favour the products
b) both driving forces favour the reactants
c) enthalpy favours the products and entropy favours the reactants
d) enthalpy favours the reactants and entropy favours the products

8. Which of the following combinations will cause a reaction to be non-spontaneous at all temperatures?
a) $-\Delta H$ and $+\Delta S$
b) $-\Delta H$ and $-\Delta S$
c) $+\Delta H$ and $-\Delta S$
d) $+\Delta H$ and $+\Delta S$

9. Which of the following statements is **true** about the reaction:

$$Ca(s) + H_2O(l) \leftrightarrow Ca(OH)_2(aq) + H_2(g) + \text{heat}$$

a) this reaction is spontaneous at all temperatures
b) this reaction is non-spontaneous at all temperatures
c) this reaction is reversible
d) as this reaction proceeds to equilibrium in a closed system, the pressure will decrease

10. At equilibrium, which of the following statements is **true** about the system $A + B \leftrightarrow C + D$

I) the total concentration of A and B is equal the total concentration of C and D
II) the forward reaction has stopped
III) the reverse reaction has stopped

a) I and II only
b) I and III only
c) I, II and III
d) none of these statements is true

11. A puddle evapourates in the sunshine. This is evidence of:

I) the tendency to minimum enthalpy
II) the tendency to maximum entropy
III) an equilibrium system
IV) a high activation energy for the reaction

a) I only
b) II only
c) I, II and III only
d) I, II, III and IV

12. The boiling point of any liquid is the temperature at which the tendency to minimum enthalpy is offset by the tendency to maximum entropy for that substance. When water boils, the enthalpy change is $\Delta H = 40.7$ kJ and the entropy change is 109.1 J/K. Use this information to calculate the boiling point of water. (ie. you are calculating the "turning point")

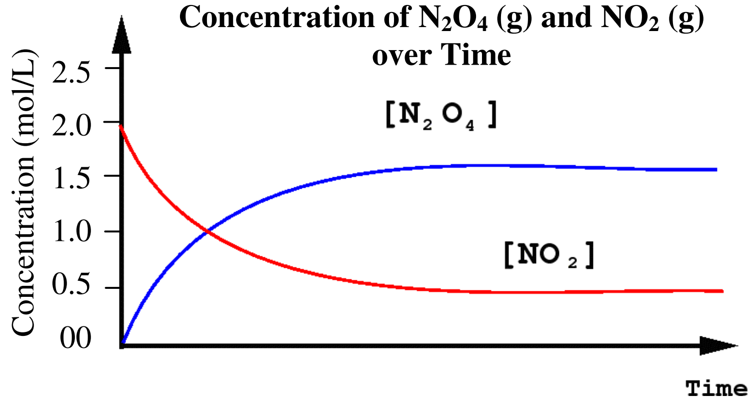
a) 100 K
b) 109 K
c) 373 K
d) 273 K

13. For a certain reaction, $\Delta G = -30$ kJ/mol. Which of the following statements is **true** for this reaction?

a) $K_{eq} = 0$
b) K_{eq} is negative
c) K_{eq} is > 1
d) K_{eq} is < 1

14. For the reaction $2\ NO_2(g) \leftrightarrow N_2O_4(g)$, referring to the graph to the right, the value for Keq at this temperature is approximately:

a) 6.0
b) 3.0
c) 0.33
d) 1.0



15. Identify the statement which is **incorrect** about a system at equilibrium:

a) a system at equilibrium has constant mass
b) the tendency to minimum entropy is balanced by the tendency to maximum enthalpy
c) to be at equilibrium the system must have constant temperature
d) at equilibrium, the macroscopic properties of the system are constant

16. A large value for the equilibrium constant indicates that:

a) the reaction has a small rate constant
b) a catalyst has been added to the system
c) the reaction favors the formation of reactants
d) the reaction favors the formation of products

17. In which of the following systems does the tendency to maximum entropy favour the products and the tendency to minimum enthalpy favour the reactants?

a) $Zn(OH)_2$ (s) + H_2CO_3 (aq) ↔ 2 H_2O (l) + $ZnCO_3$ (aq) + heat
b) H_2O (g) + CO_2 (g) ↔ H_2CO_3 (l) + heat
c) P_4(s) + 10 Cl_2(g) + heat ↔ 4 PCl_5(g)
d) PCl_5(g) + heat ↔ PCl_3(g) + Cl_2(g)

18. For the system: 2 SO_3(g) ↔ 2 SO_2(g) + O_2(g), the equilibrium constant expression (K_c) is:

a) $[SO_2]^2 / [SO_3]$
b) $[SO_3]^2 / [SO_3]^2[O_2]$
c) $[SO_2]^2[O_2] / [SO_3]^2$
d) $[SO_2][O_2]$

19. With reference to the following reaction taking place in a closed system, the term “dynamic equilibrium” means:

$$2\,NO\,(g) + O_2\,(g) \leftrightarrow 2\,NO_2(g)$$

a) all of the NO and O_2 have reacted to form NO_2
b) the number of molecules of NO_2 is equal to the number of molecules of NO
c) the total number of molecules of NO_2 is equal to the number of molecules of both NO and O_2
d) the products are forming at the same rate that the reactants are reforming

20. For the equilibrium system below, which of the following will increase the concentration of NO_2 (g)?

$$2\,NO\,(g) + O_2\,(g) \leftrightarrow 2\,NO_2(g) + \text{heat}$$

a) a decrease in the total pressure at constant temperature
b) a decrease in the concentration of O_2 (g) at constant temperature
c) a decrease in the temperature at constant pressure
d) an increase in the volume of the reaction vessel at constant temperature

21. What is the equilibrium constant expression for the reaction below?

$$2\,CaO(s) + 2\,SO_2(g) + O_2(g) \leftrightarrow 2\,CaSO_4(s)$$

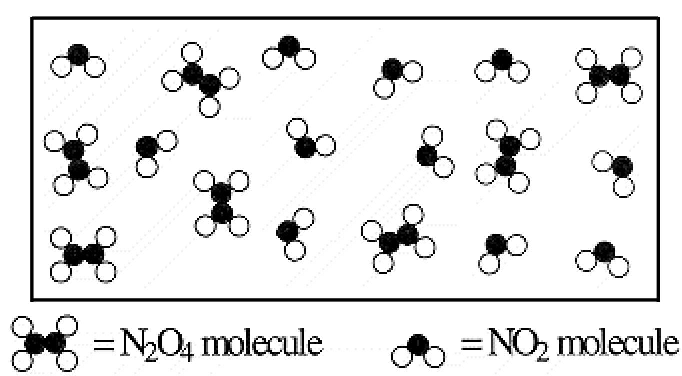
a) $K_c = [CaSO_4]^2 / [CaO]^2[SO_2]^2[O_2]$
b) $K_c = 1 / [SO_2]^2[O_2]$
c) $K_c = [CaO]^2[SO_2]^2[O_2] / [CaSO_4]^2$
d) $K_c = [SO_2][O_2]$

22. For the equilibrium system:

$$N_2O_4\,(g) \leftrightarrow 2\,NO_2\,(g)$$

calculate the value for K_{eq} for the system as shown in the diagram to the right:

a) 17.3
b) 4.45
c) 1.57
d) 0.636



23. Nitrogen reacts with hydrogen to form ammonia: $N_2(g) + 3 H_2(g) \leftrightarrow 2 NH_3(g)$
An equilibrium mixture at a given temperature is found to contain 0.31 mol/L N_2, 0.50 mol/L H_2, and 0.14 mol/L NH_3. Calculate the value of K_c at this temperature.

a) 1.97
b) 1.107
c) 0.903
d) 0.506

24. The equilibrium constant for the reaction: $2 HBr(g) \leftrightarrow H_2(g) + Br_2(g)$ is $K_c = 1.26 \times 10^{-12}$ at 500 K. This implies that:

a) at equilibrium, the total concentration of the product is much greater than that of the reactants
b) the reaction has a large negative ΔG
c) the rate of this reaction is very fast
d) the reactants have much lower enthalpy than the products

25. Consider the following reversible reaction: $N_2(g) + 3 H_2(g) \leftrightarrow 2 NH_3(g)$
In a 3.00 litre container at 400 °C at equilibrium, there are: 0.0420 mole N_2, 0.516 mole H_2 and 0.0357 mole NH_3. Evaluate K_c.

a) 0.202
b) 16.0
c) 1.99
d) 4.94

26. If the equilibrium constant for the reaction: $A + 2 B \leftrightarrow C + 5 D$ has a value of 4.0, what is the value of the equilibrium constant for the reaction: $C + 5 D \leftrightarrow A + 2 B$ at the same temperature?

a) 0.25
b) 2.0
c) 8.0
d) 16

27. Consider the following reaction. Which is the correct equilibrium constant expression?

$$4 Br_2(l) + CH_4(g) \leftrightarrow 4 HBr(g) + CBr_4(g)$$

a) $K_c = [CBr_4][HBr] / [Br_2][CH_4]$
b) $K_c = [CBr_4][HBr]^4 / [Br_2]^4[CH_4]$
c) $K_c = [HBr]^4[CBr_4] / [CH_4][Br_2]$
d) $K_c = [CBr_4][HBr]^4 / [CH_4]$

28. The equilibrium constant for the reaction: $Br_2(g) + F_2(g) \leftrightarrow 2 BrF(g)$ is 54.7. What is the equilibrium concentration of BrF if the initial concentrations of bromine and fluorine were both 0.250 mol/L?

a) [BrF] = 0.241 M
b) [BrF] = 0.199 M
c) [BrF] = 0.39 M
d) [BrF] = 0.25 M

29. For the system: $PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$, $K_c = 1.90$ at a certain temperature.
The system is in equilibrium with concentrations $[PCl_3] = 0.500$ M and $[Cl_2] = 0.500$ M. What is the PCl_5 concentration?

a) 0.950 M
b) 1.90 M
c) 0.500 M
d) 0.475 M

30. When $K_c >> 1$ for a chemical reaction:
I) the equilibrium would be achieved rapidly
II) the equilibrium would be achieved slowly
III) reactant concentrations would be much greater than product concentrations at equilibrium
IV) product concentrations would be much greater than reactant concentrations at equilibrium

a) I and II only
b) I and IV only
c) II and III only
d) IV only

31. For the reaction: cyclopropane ↔ propene, at a certain temperature, $K_c = 3.0$
If cyclopropane is placed in a closed flask at a concentration of 2.0 M, what will the concentration of propene be at equilibrium?

a) 1.0 M
b) 1.5 M
c) 0.66 M
d) 2.0 M

32. The reversible reaction: $2\ SO_2(g) + O_2\ (g) \leftrightarrow 2\ SO_3(g)$ is at equilibrium. Before the reaction, the concentrations of the gases were 0.060 mol/L of SO_2, 0.050 mol/L of O_2 and 0.00 mol/L of SO_3. At equilibrium, the concentration of SO_3 was 0.040 mol/L. Calculate K_c for this reaction.

a) 133
b) 8.88
c) 4.0
d) 13.3

33. For the system: $2\ HI(g) \leftrightarrow H_2(g) + I_2(g)$, at 445°C, the value for K_c is 0.020
A mixture of H_2, I_2, and HI in a vessel at 445°C has the following concentrations: [HI] = 2.0 M, $[H_2] = 0.50$ M and $[I_2] = 0.10$ M. Which one of the following statements about Q_c (the reaction quotient) is **TRUE** for the above system?

a) $Q_c = K_c$; the system is at equilibrium
b) Q_c is less than K_c; more HI will form
c) Q_c is less than K_c; so more H_2 & I_2 will form
d) Q_c is greater than K_c; more H_2 and I_2 will form

34. Consider the reaction: $N_{2(g)} + O_{2(g)} \leftrightarrow 2\ NO_{(g)}$
0.25 mol of nitrogen and 0.25 mol of oxygen are placed in a 5.00 litre reaction vessel and heated. At equilibrium 0.16 mol of NO are present. What is the value of K_{eq} for the reaction?

a) 0.41
b) 0.89
c) 28
d) 2.6

35. A quantity of HI was sealed in a tube and heated to 425°C. At equilibrium, the concentration of HI in the tube was 0.0706 mol/L. Calculate the equilibrium concentration of H_2, given:

$$H_2\ (g) + I_2\ (g) \leftrightarrow 2HI\ (g) \qquad K_c = 54.6 \text{ at } 425°C$$

a) 9.55×10^{-3} M
b) 1.85×10^{-4} M
c) 1.17×10^{-3} M
d) 4.78×10^{-3} M

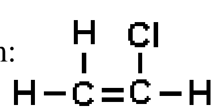
36. Consider the reaction: $N_2(g) + O_2(g) \leftrightarrow 2\ NO(g)$, $K_c = 0.10$ at 2000°C
Starting with initial concentrations of 0.040 mol/L of N_2 and 0.040 mol/L of O_2, calculate the equilibrium concentration of NO in mol/L:

a) 0.0055 mol/L
b) 0.011 mol/L
c) 0.0096 mol/L
d) 0.080 mol/L

37. Consider the following reaction: $H_2(g) + I_2(g) \leftrightarrow 2\ HI(g)$, Kc = 25.0 at 1100 K. If a flask is filled with HI (g) at a concentration of 4.00 M HI and sealed, what will the concentration of I_2 (g) be at equilibrium?

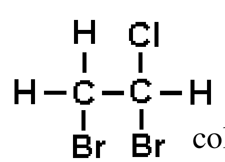
a) 2.00 M
b) 0.148 M
c) 0.571 M
d) 0.363 M

38. Consider the equilibrium:



+ Br_2 (l) ⟶

orange



colourless liquid

At higher temperatures, the colour of the system becomes less orange. The forward reaction is:

a) exothermic
b) non-spontaneous at all temperatures
c) endothermic
d) spontaneous at all temperatures

39. Nitrogen reacts with hydrogen to form ammonia: $N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g)$ $K_c = 2.0$
At equilibrium, the concentration of N_2 is 0.15 M and the concentration of H_2 is 0.30 M. What is the concentration of NH_3 in this mixture?

a) 2.7×10^{-2} M
b) 8.1×10^{-3} M
c) 0.16 M
d) 9.0×10^{-2} M

40. For the reaction: $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$, at 450°C, $K_c = 0.040$.
If a reaction is initiated with 0.40 mole of Cl_2 and 0.40 mole of PCl_3 in a 2.0 litre container, what is the equilibrium concentration of Cl_2 in the same system?

a) 0.072 M
b) 0.11 M
c) 0.16 M
d) 0.040 M

41. The equilibrium expression (K_c) for the system: $2\ ICl(s) \leftrightarrow I_2(s) + Cl_2(g)$ is:

a) $[I_2][Cl_2] / [ICl]^2$
b) $[Cl_2]$
c) $[I_2][Cl_2] / 2[ICl]$
d) $([I_2] + [Cl_2]) / 2[ICl]$

42. For the reversible reaction: $2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g)$. Concentrations of 0.060 mol/L of SO_2 and 0.050 mol/L of O_2 are allowed to react in a closed container. At equilibrium, the concentration of SO_3 is 0.040 mol/L. What is the equilibrium concentration of O_2?

a) 0.010 M
b) 0.030 M
c) 0.020 M
d) 0.040 M

43. For the equilibrium system: $2\ H_2O(g) \leftrightarrow 2\ H_2(g) + O_2(g)$
Given that the forward reaction is **endothermic**, which of the following changes will **decrease** the equilibrium amount of H_2O?

a) adding more oxygen
b) decreasing the volume of the container
c) adding a solid phase catalyst
d) increasing the temperature at constant pressure

44. Consider the following gas-phase reaction at equilibrium: $Cl_2\ (g) + 3\ F_2\ (g) \leftrightarrow 2\ ClF_3\ (g)$
If the concentration of F_2 (g) is suddenly doubled at constant pressure and volume, which of the following will be true when the system is again at equilibrium?

a) the concentrations of both F_2 (g) and ClF_3 (g) will be higher; Cl_2 (g) will be lower
b) the concentrations of both F_2 (g) and Cl_2 (g) will be lower; ClF_3 (g) will be higher
c) the concentration of ClF_3 (g) will be lower; Cl_2 (g) and F_2 (g) will both be higher
d) the concentrations of all three species will be unaffected

45. Given the following reaction, equilibrium constant, and molar concentrations of the three species,

$2\ NO\ (g)$ +	$O_2\ (g)$ $\leftrightarrow$	$2\ NO_2\ (g)$,	$K_c = 6.2 \times 10^5$
0.0513 M	0.112 M	0.000212 M	

we can accurately predict:

a) that the reaction is at equilibrium
b) that the reaction is not at equilibrium and must proceed from left to right to reach equilibrium
c) that the reaction is not at equilibrium and must proceed from right to left to reach equilibrium
d) that the reaction is non-spontaneous and is not an equilibrium system

46. For the reaction: $2\ NO\ (g) + O_2\ (g) \leftrightarrow 2\ NO_2\ (g)$, $K_c = 6.2 \times 10^5$ at 500K. What is the Kc for the reverse reaction at the same temperature?

a) -6.2×10^5
b) 6.2×10^{-5}
c) 1.6×10^{-6}
d) 1.6×10^6

47. Consider the equilibrium system: $2\ ICl(s) \leftrightarrow I_2(s) + Cl_2(g)$. Which of the following changes will increase the total amount of Cl_2 that can be produced?

I) removing some of the $I_2(s)$
II) adding more $ICl(s)$
III) removing the Cl_2 as it is formed
IV) decreasing the volume of the container

a) I and II
b) III and IV
c) II and III only
d) III only

48. What would be the effect of increasing the volume of the following system at equilibrium?

$$2\ CO(g) + O_2(g) \leftrightarrow 2\ CO_2(g)$$

a) the K_p value would get larger
b) the equilibrium would be perturbed and would show a net shift to the left
c) the equilibrium would be perturbed and would show a net shift to the right
d) there would be no effect; the system is at equilibrium

49. Which of the following equilibrium systems will **NOT** be affected by a change in pressure?

a) $H_2O\ (l) + CO_2\ (g) \leftrightarrow CH_2O\ (l) + O_2\ (g)$
b) $2\ NH_3(g) + 2\ N_2O_3(g) \leftrightarrow 3\ N_2(g) + 3\ H_2O_2(g)$
c) $2\ N_2H_4(l) \leftrightarrow 2\ N_2(g) + 4\ H_2(g)$
d) $2\ NH_3(g) + 3\ N_2O(g) \leftrightarrow 4\ N_2(g) + 3\ H_2O(g)$

50. For the following system at equilibrium: $2\ SO_3\ (g) + Cl_2\ (g) \leftrightarrow 2\ SO_2Cl_2\ (g) + O_2\ (g)$, which of the statements is/are **correct**?

a) addition of Cl_2 would cause the concentrations of both SO_2Cl_2 and O_2 to increase
b) addition of O_2 would cause the concentrations of both SO_2Cl_2 and SO_3 to increase
c) increasing the pressure of the system would cause the reaction to shift from left to right
d) both a) and c) are correct

51. At equilibrium, a 1.0 litre container was found to contain 0.20 moles of A, 0.20 moles of B, 0.40 moles of C and 0.40 mole of D. If 0.10 moles of A and 0.10 moles of B are added to this system, what will be the new equilibrium concentration of A?

$$A(g) + B(g) \leftrightarrow C(g) + D(g)$$

a) 0.37 mol/L
b) 0.87 mol/L
c) 0.47 mol/L
d) 0.23 mol/L

52. For the reaction: $2\ H_2(g) + O_2(g) \leftrightarrow 2\ H_2O(g)$ at a certain temperature the K_{eq} is 1.8×10^5. If the concentrations in a reaction vessel at a given time are $[H_2] = 3.0$ M, $[O_2] = 2.0$ M and $[H_2O] = 0.010$ M, which of the following statements is **true**?

a) the system is at equilibrium
b) the system is not at equilibrium, the reaction will proceed to the left
c) the system is not at equilibrium, the reaction will proceed to the right
d) the E_a for the forward reaction is very large

53. Which of the following statements is **true** about Keq, the equilibrium constant?

a) it always remains the same for a reaction regardless of the reaction conditions
b) it increases if the concentration of one of the products is increased
c) it changes with changes in the temperature
d) it may be changed by the addition of a catalyst

54. Consider the reversible reaction at equilibrium at 392°C: $2A(g) + B(g) \leftrightarrow C(g)$
At equilibrium, the partial pressures are found to be: A: 6.70 atm, B: 10.1 atm, C: 3.60 atm. Evaluate K_p for this reaction.

a) 7.94×10^{-3}
b) 0.0532
c) 0.146
d) 54.5

55. The reaction: $2\ NO_2(g) \leftrightarrow 2\ NO(g) + O_2(g)$ is endothermic. Which of the following will cause the concentration of NO_2 (g) to increase?

a) decreasing the concentration of NO (g)
b) increasing the volume of the system
c) increasing the pressure of the system
d) increasing the temperature of the system

56. The following system is in equilibrium in a closed vessel: $N_2(g) + 3\ Cl_2(g) \leftrightarrow 2\ NCl_3(g)$.
What will happen if one mole of He gas is injected into the system?

a) equilibrium is perturbed and more product will form as equilibrium is restored
b) equilibrium is perturbed and more reactants will form as equilibrium is restored
c) the equilibrium concentrations of the system will not change
d) the equilibrium constant will get larger

57. For which of the following reactions would the equilibrium concentrations **NOT** be affected by a change in the total pressure?

a) $N_2(g) + 3\ H_2(g) \leftrightarrow 2\ NH_3(g)$
b) $2\ NO_2(g) \leftrightarrow N_2(g) + 2\ O_2(g)$
c) $PCl_3(g) + Cl_2(g) \leftrightarrow PCl_5(g)$
d) $CO(g) + H_2O(g) \leftrightarrow CO_2(g) + H_2(g)$

58. Ammonia (NH_3) is produced using the Haber process: $N_2\ (g) + 3\ H_2\ (g) \leftrightarrow 2\ NH_3\ (g)$.
This reaction is exothermic. Which of the following would **increase** the amount of NH_3 obtained?

I) decrease the pressure
II) increase the temperature
III) increase the concentration of N_2
IV) decrease the volume

a) I and II only
b) I and III only
c) III and IV only
d) II and IV only

59. Laughing gas, N_2O, can be prepared (ha, ha!) from H_2 and NO:

$H_2\ (g) + 2\ NO\ (g) \leftrightarrow N_2O\ (g) + H_2O\ (g)$, $K_c = 2.3 \times 10^6$ at a certain temperature

If the closed system (ha, ha!) contains 0.050 M H_2 (g), 0.020 M NO (g), 5.4 M N_2O (g) and 8.7 M H_2O (g), then we can accurately predict (ha, ha!) that:

a) this system is far from equilibrium and will shift to the left to achieve equilibrium
b) this system is far from equilibrium and will shift to the right to achieve equilibrium
c) the reaction is at equilibrium
d) these questions are laughable

60. If the reaction quotient for a reaction is larger than the equilibrium constant (i.e., $Q_c > K_c$) then the reaction:

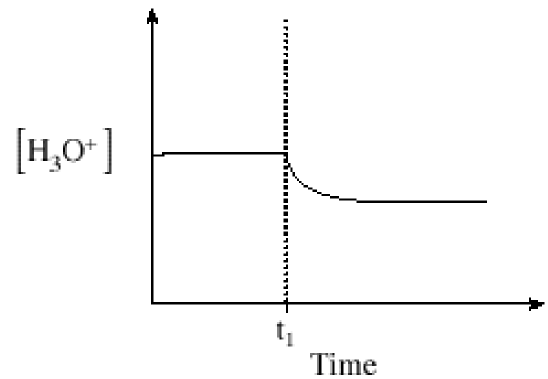
a) will always proceed to equilibrium very rapidly
b) must shift from right to left to reach equilibrium
c) must shift from left to right to reach equilibrium
d) is at equilibrium

61. Consider the following system at equilibrium:

$$CH_3COOH\ (l) + H_2O\ (l) \leftrightarrow CH_3COO^{-1}\ (aq) + H_3O^+\ (aq) + heat$$

A stress was imposed on the system at t = 1 and it caused the response shown in the rate graph below. Which of the following stresses would produce this response?

a) increasing the volume of the container
b) adding H_3O^+
c) decreasing the temperature
d) the addition of CH_3COO^-



62. For the equilibrium system below, when the temperature is increased, the solution turns a dark blue. Based on this observation:

$$\underset{(pink)}{Co(H_2O)_6{}^{2+}{}_{(aq)}} + 4Cl^-{}_{(aq)} \rightleftarrows \underset{(blue)}{CoCl_4{}^{2-}{}_{(aq)}} + 6H_2O_{(\ell)}$$

a) the forward reaction is endothermic and K_{eq} has increased
b) the forward reaction is exothermic and K_{eq} has increased
c) the forward reaction is endothermic and K_{eq} has decreased
d) the forward reaction is exothermic and K_{eq} has decreased

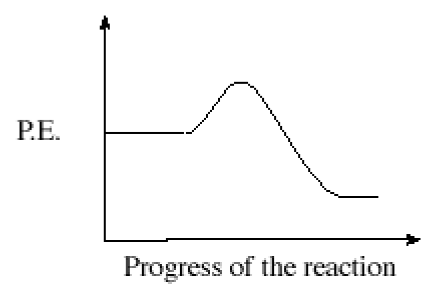
63. Consider the following reaction: $C\ (s) + 2\ H_2\ (g) \leftrightarrow CH_4\ (g)$ $\Delta H = -74.8$ kJ

Which of the following will cause an increase in the value of K_{eq} ?

a) decreasing the temperature
b) increasing $[H_2]$
c) finely powdering the C(s)
d) decreasing the volume

64. Consider the potential energy diagram for an equilibrium system shown to the right. When the temperature of the system is increased, the equilibrium shifts to the:

a) right and K_{eq} increases
b) right and K_{eq} decreases
c) left and K_{eq} increases
d) left and K_{eq} decreases



65. Consider this equilibrium: $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$ $K_{eq} = 50.0$

What is the value K_{eq} for the reaction rewritten as:

$$2HI_{(g)} \rightleftarrows H_{2(g)} + I_{2(g)} \qquad K_{eq} = ?$$

a) 25.0
b) –50.0
c) 50.0
d) 0.0200

Provide full written answers to following types of questions. Include a complete solution including an equation, substitution, final answer, correct number of sig digs and all necessary units:

1. Consider the following equilibrium: $H_2 (g) + I_2 (g) \leftrightarrow 2 HI (g)$

   At equilibrium $[H_2]$=0.00220mol/L, $[I_2]$=0.00220mol/L, and [HI]=0.0156mol/L. Calculate K_{eq} .

2. Consider the reaction: $2 NOBr (g) \leftrightarrow 2 NO (g) + Br_2 (g)$, $K_{eq} = 0.064$

   At equilibrium, a 1.00L flask contains 0.030 mol NOBr and 0.030 mol NO. How many moles of Br_2 are present in the flask at equilibrium?

3. Consider the following reaction: $2 NO (g) \leftrightarrow N_2 (g) + O_2 (g)$ Keq = 87.9

   If 1.50 mol of N_2 (g), 1.50 mole O_2 (g) and 0.100 mol NO (g) are placed in a 10.0 L reaction vessel:

a) Is this system at equilibrium?
b) If it is not at equilibrium, in which direction will the system shift to achieve equilibrium?
c) Calculate the concentrations of all species at equilibrium.

4. Consider the equilibrium system: $SO_2(g) + NO_2(g) \leftrightarrow SO_3(g) + NO(g)$

   At a given temperature, analysis of an equilibrium mixture finds the following concentrations:

   $[SO_2]$ = 4.0 M, $[NO_2]$ = 0.50 M, $[SO_3]$ = 3.0 M, [NO] = 2.0 M. Calculate K_{eq} for the reaction at this temperature.

5. Gas X_2 reacts with gas Y_2 according to the equation: $X_2 (g) + Y_2 (g) \leftrightarrow 2 XY (g)$

   0.50 mole each of X_2 and Y_2 are placed in a 1.0 litre vessel and allowed to reach equilibrium at a given temperature. The equilibrium concentration of XY is found to be 0.025 mol/L. What is the equilibrium constant for this reaction?

6. At 462°C, for the reaction: $2 NOCl(g) + heat \leftrightarrow 2 NO(g) + Cl_2(g)$ $K_{eq} = 8.0 \times 10^{-5}$

   If 2.5 moles of NOCl (g) are placed in a 2.00 L flask and allowed to react, what will the concentration of NO (g) be at equilibrium?

7. Consider the following reaction: $H_2(g) + I_2(g) \leftrightarrow 2 HI(g)$ K_c is 25 at 1100 K

   HI (g) is placed in a sealed flask at an initial concentration of 4.00 M HI(g). When the system is at equilibrium, what is the concentration of $I_2(g)$?

8. Consider the reaction: $2 HF (g) \leftrightarrow H_2(g) + F_2(g)$ Kc = 4.0 at a certain temperature

   3.0 mol of HF (g), 2.0 mol of H_2 (g) and 4.0 mol of F_2 are placed in a 5.0 L reaction vessel.

a) In which direction will the system shift to achieve equilibrium (show your work)
b) Calculate the concentrations of all species at equilibrium

9. Consider the reaction: $SO_2 (g) + NO_2 (g) \leftrightarrow SO_3 (g) + NO (g)$.

   The system is at equilibrium in a 1.0 L vessel and contains 0.50 moles SO_3 (g), 0.50 mol NO (g), 0.60 moles SO_2 (g) and 0.10 moles NO_2 (g). 0.5 moles of NO_2 (g) is added to the system. What is the concentration of NO (g) when equilibrium is re-established?

Answers to Multiple Choice

1. d	14. a	27. d	40. a	53. c
2. d	15. b	28. c	41. b	54. a
3. b	16. d	29. d	42. b	55. c
4. b	17. d	30. d	43. d	56. c
5. a	18. c	31. b	44. a	57. d
6. a	19. d	32. a	45. b	58. c
7. c	20. c	33. c	46. c	59. c
8. c	21. b	34. b	47. d	60. b
9. a	22. a	35. a	48. b	61. d
10. d	23. d	36. b	49. a	62. a
11. b	24. d	37. c	50. a	63. a
12. c	25. c	38. c	51. d	64. d
13. c	26. a	39. d	52. c	65. d

Answers to Full Calculations: (full written solutions are in the "answer book)

1. Keq = 50.3
2. $[Br_2]$ = 0.064 M

3a) Q = 225, Q ≠ Keq so the system is not at equilibrium

3b) Q > Keq so the system will shift to the left (⟵) to produce more reactants

3c) at eq'm, $[N_2]$ = $[O_2]$ = 0.147 M and [NO] = 0.0157 M

4. Keq = 3.0
5. Keq = 2.6×10^{-3}
6. use the 500 rule. [NO] at eq'm is 0.063M
7. $[I_2]$ at eq'm is 0.57 M

8a) Q = 0.889, this is less than Keq so the reaction will shift to the right to make more product

8b) use quadratic equation, x = 0.126 so at eq'm, [HF] = 0.35 M, $[H_2]$ = 0.53 M, $[F_2]$ = 0.93 M

9. use initial eq'm concentrations to calculate Keq = 4.17, then use E' ICE table to find [NO] = 0.74 M