* ***This is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.***
* ***For MC, DO NOT USE A CALCULATOR. You may use ONLY a periodic table.***
* ***Try to work at a pace of about 1.3 min per MC question. Time yourself. You know how important it is that you practice working for speed. Use 15 minutes per FR.***
* ***Then when time is up, continue working without a calculator.***
* ***After you have completed as many as possible without the calculator. Finish with a calculator.***
* ***Please email me if you think you find any issues or typos.***

1. When the substances in the equation below are at equilibrium at pressure P and temperature T, the equilibrium can be shifted to favor the products by

* MgO(s) + H2(g)  🡪 Mg(s) + H2O(g) ∆H = −14 kJ
  1. increasing the pressure in the reaction vessel while keeping the temperature constant
  2. increasing the pressure by adding an inert gas such as argon
  3. decreasing the temperature
  4. allowing some hydrogen gas to escape at constant P and T
  5. adding a catalyst

2 A(g) + B(g) 🡪 2C(g)

1. When 0.60 mole of A and 0.75 mole of B are placed in an evacuated 1.00 L flask, the reaction represented above occurs. After the reactants and the product reach equilibrium and the initial temperature is restored, the flask is found to contain 0.30 mole of product C. Based on these results, the equilibrium constant, *Kc* for the
   1. 0.60
   2. 0.90
   3. 1.7
   4. 3.4
   5. 6.0
2. What is the balanced net ionic equation that occurs when an excess of ammonia gas is bubbled through a solution saturated with silver chloride?
   1. Ag+ + Cl− + NH4+ + OH− → NH4Cl + AgOH
   2. Ag+ + 2H2O + 2NH3 → 2NH4+ + Ag(OH)2−
   3. Ag+ + H2O + NH3 → NH4+ + AgOH
   4. AgCl + 2NH3 → Cl− + Ag(NH3)2+
   5. 2Ag+ + H2O + 2NH3 → 2NH4+ + Ag2O
3. Which of the following systems would NOT experience a change in the number of moles of the substances present at equilibrium when the volume of the system is changed at constant temperature?
   1. SO(g) + NO(g) ⇄ SO2(g) + ½N2(g)
   2. O2(g) + 2H2(g) ⇄ 2H2O(g)
   3. N2(g) + 2O2(g) ⇄ 2NO2(g)
   4. N2O4(g) ⇄ 2NO2(g)
   5. CH4(g) + 2O2(g) ⇄ CO2(g) + 2H2O(g)
4. For the reaction 2W(g) ⇄ 2X(g) + Y(g), the equilibrium constant, *Kp*, is 8 x 103 at 298 K. A mixture of the three gases at 298 is placed in a rigid metal cylinder, and the initial pressures are Px = 1 atm, Py = 0.8 atm, and Pw = 2 atm. At the instant of mixing, which of the following is true?
   1. More product will form.
   2. More reactant will form.
   3. ∆S = 0
   4. ∆Gº = 0
   5. ∆Gº > 0

CaCO3(s) ⇄ CaO(s)  + CO2(g)

1. After the equilibrium represented above is established, some pure CO2(g) is added to the reaction vessel at constant temperature. After equilibrium is reestablished, all of the following will happen EXCEPT
   1. *Kp* for the reaction will increase.
   2. *Kc* for the reaction will remain the same.
   3. The concentration of CaCO3 in the reaction vessel remains constant.
   4. The concentration of CaO in the reaction vessel remains constant.
   5. ∆G will become zero.

N2(g) + 3 H2(g) ⇄ 2 NH3(g) ∆H = −92 KJ

1. Which of the following changes alone would cause a decrease in the value of *Keq* for the reaction represented above?
   1. Removing NH3
   2. Increasing the temperature
   3. Adding an inert gas such as argon
   4. Adding more hydrogen gas at constant temperature
   5. Adding a catalyst

2SO2(g) + O2(g) ⇄ 2SO3(g) ∆H = −198 kJ

1. Consider the equilibrium above. Which of the following changes will increase the concentration of SO2(g)?
   1. Increasing the concentration of O2(g)
   2. Increasing the pressure in the reactant vessel at constant temperature
   3. Increasing the mass of SO3 present
   4. Adding a catalyst
   5. Decreasing the temperature
2. Consider the reaction system,

* NiO(s) + H2(g)  ⇋ Ni(s) + H2O(g)
* The equilibrium constant expression is
  1. Keq = (Ni)(H2O)/(NiO)(H2)
  2. Keq = (Ni)/(NiO)(H2)
  3. Keq = 1/(NiO)(H2)
  4. Keq = 1/(H2)
  5. Keq = (H2O)/(H2)

1. Given the equilibrium,

2SO2(g) + O2(g) ⇋ 2SO3(g)

* If this equilibrium is established by beginning with equal number of moles of SO2 and O2 in an empty 1.0 L bulb, then the following must be true at equilibrium:
  1. [SO2] = [SO3]
  2. 2[SO2] = 2[SO3]
  3. [SO2] = [O2]
  4. [SO2] < [O2]
  5. [SO2] > [O2]

1. Consider: N2O4(g) ⇋ 2NO2(g)

* At 25°C, in a 1.0 L container 0.11 mole of N2O4 reacts to form 0.10 mol of N2O4 and 0.02 mole of NO2 at equilibrium.
* At 90°C, 0.11 mole of N2O4 results in 0.050 mole of N2O4 and 0.12 mole of NO2 at equilibrium. From these data we can conclude
  1. Equilibrium constants, in general, are always larger at higher temperatures.
  2. N2O4 molecules react by a first order rate law.
  3. the reaction is endothermic.
  4. NO2 molecules react twice as fast as N2O4 molecules at 90°C.
  5. the equilibrium constant for the reaction above decreases with an increase in temperature.

*The next 2 questions refer to the following:*

At a given temperature, 0.300 mole NO, 0.160 mol Cl2 and 0.500 mol ClNO were placed in a 10.0 Liter container. The following equilibrium is established:

2ClNO(g) ⇋ 2NO(g) + Cl2(g)

At equilibrium, 0.600 mol of ClNO was present

1. The number of moles of Cl2 present at equilibrium is
   1. 0.090
   2. 0.050
   3. 0.060
   4. 0.200
   5. 0.110
2. The equilibrium constant, *Kc*, is closest to:
   1. 10-6
   2. 10-3
   3. 1
   4. 103
   5. 106
3. At 985°C, the equilibrium constant for the reaction,

H2(g) + CO2(g) ⇋ H2O(g) + CO(g)

* is 1.5. What is the equilibrium constant for the reaction below?
* H2O(g) + CO(g) ⇋ H2(g) + CO2(g)
  1. -1.5
  2. 0.75
  3. 3.0
  4. 0.67
  5. There is no way of knowing what it would be.

1. What is the relationship between *Kp* and *Kc* for the reaction, 2ICl(g) ⇋ I2(g) + Cl2(g)
   1. Kp = Kc (RT)-1
   2. Kp = Kc  (RT)
   3. Kp = Kc (RT)2
   4. Kp  = Kc
   5. Kp  = Kc (2RT)
2. What is the relationship between *Kp* and *Kc* for the reaction, 3A(g) ⇋ 2B(g) + C(s)
   1. Kp = Kc (RT)-1
   2. Kp = Kc  (RT)
   3. Kp = Kc (RT)2
   4. Kp  = Kc
   5. Kp  = Kc (2RT)
3. For the reaction system,

N2(g) + 3 H2(g)  ⇋ 2NH3(g) + heat

* the conditions that would favor maximum conversion of the reactants to products would be
  1. high temperature and high pressure
  2. high temperature, pressure unimportant
  3. high temperature and low pressure
  4. low temperature and high pressure
  5. low temperature and low pressure

1. Solid HgO, liquid Hg, and gaseous O2 are placed in a glass bulb and are allowed to reach equilibrium at a given temperature.

2HgO(s) ⇋ 2Hg(L) + O2(g) ΔH = +43.4 kcal

* The mass of HgO in the bulb could be increased by
  1. adding more Hg.
  2. removing some O2.
  3. reducing the volume of the bulb.
  4. increasing the temperature.
  5. removing some Hg.

1. For the equilibrium system

* H2O(g) + CO(g)  ⇋ H2(g) + CO2(g) ΔH = -42 kJ/mol
* *Kc* equals 0.62 at 1260 K. If 0.10 mole each of H2O, CO, H2 and CO2 (each at 1260 K) were placed in a 1.0-Liter flask at 1260 K, when the system came to equilibrium.

|  |  |  |
| --- | --- | --- |
|  | The temperature would | The mass of CO would |
| a. | decrease | Increase |
| b. | decrease | decrease |
| c. | remain constant | increase |
| d. | increase | decrease |
| e. | increase | increase |

1. Consider the reaction

* SO2(g) + ½O2(g) ⇋ SO3(g)  *Kc* = 49 at 1000 K
* What is the value of *Kc* for the reaction below?
* 2SO3(g) ⇋ 2SO2(g) + O2(g)
  1. 
  2. 
  3. 
  4. 7
  5. (49)2

1. For which reaction is Kc equal to Kp?
   1. H2(g) + S(s) ⇄ H2S(g)
   2. 2H2O(g) ⇄ 2H2(g) + O2(g)
   3. 3H2(g) + N2(g) ⇄ 2NH3(g)
   4. H2(g) + Br2(L) ⇄ 2HBr(g)
   5. 2NO2(g) ⇄ N2O4(g)

* heat + N2O4(g) ⇋ 2NO2(g)
* For the next 3 questions, consider an equilibrium system based on the reaction above. This equilibrium mixture is contained in a piston.

1. Which occurs when the volume of the system is increased at constant temperature?

**#molecules total # of molecules**

**of NO2 of all gases *Kp***

* 1. increases increases remains same
  2. increases decreases remains same
  3. increases remains same remains same
  4. remains same decreases decreases
  5. remains same increases decreases

1. Which occurs when a catalyst is added?

**partial pressure total pressure**

**of N2O4 of all gases *Kp***

* 1. decreases increases remains same
  2. increases decreases remains same
  3. stays the same stays the same remains same
  4. increases decreases decreases
  5. increases increases decreases

1. Which occurs when the temperature is increased at constant volume?

**#molecules total # of molecules**

**of N2O4 of all gases *Kp***

* 1. increases decreases remains same
  2. decreases increases remains same
  3. increases decreases increases
  4. decreases decreases increases
  5. decreases increases increases

1. 4NH3(*g*) + 3O2(*g*) ↔ 2N2(*g*) + 6H2O(*g*) + energy

* Which of the following changes to the system at equilibrium shown above would cause the concentration of H2O to increase?
  1. The volume of the system was decreased at constant temperature.
  2. The temperature of the system was increased at constant volume.
  3. NH3 was removed from the system.
  4. N2 was removed from the system.
  5. O2 was removed from the system.

4H2(g) + CS2(g) ⇋ CH4(g) + 2H2S(g)

A mixture of 2.50 mol H2(g), 1.50 mol CS2(g), 1.5 mol CH4(g) and 2.00 mol H2S(g) is placed in a 5.0 L rigid reaction vessel, and the system reaches equilibrium according to the equation above. When the equilibrium is achieved, the concentration of CH4(g) has become 0.25 mol L−1.

*The next 5 questions refer to this equilibrium system.*

1. Changes in concentration occur as this system approaches equilibrium. Which expression gives the best comparison of the changes in those concentrations shown in the ratio to the right?
   1. 
   2. 
   3. 
   4. 
   5. 
2. What is the change in the number of moles of H2S(g) present as the system moves from its original state to the equilibrium described?
   1. -2.50
   2. 1.25
   3. -0.50
   4. -0.25
   5. -0.10
3. What is the number of moles of CS2(g) at equilibrium?
   1. 0.25
   2. 0.35
   3. 0.75
   4. 1.25
   5. 1.75
4. What is the concentration in moles per liter of H2(g) at equilibrium?
   1. 0.50
   2. 0.70
   3. 1.0
   4. 2.0
   5. 3.5
5. Which correctly describes the values for ∆G, the free energy change, and Q, the reaction quotient, when the mixture was prepared.
   1. ∆G = 0, Q = Keq
   2. ∆G > 0, Q < Keq
   3. ∆G > 0, Q > Keq
   4. ∆G < 0, Q > Keq
   5. ∆G < 0, Q < Keq

N2(g) + 3H2(g) ⇋ 2NH3(g) + 92 kJ

* The next 4 questions apply to an equilibrium system based on the reversible reaction given above.

1. When the temperature of such an equilibrium system is increased at constant volume, which property is least affected?
   1. density
   2. pressure
   3. concentration of NH3(g)
   4. concentration of N2(g)
   5. average kinetic energy
2. How are the rates of the opposing reactions and the value of the equilibrium constant, Kp, affected when heat is added to this system?
   1. The reaction rates remain unchanged and Kp decreases.
   2. The forward reaction will be favored and Kp decreases.
   3. The forward reaction will be favored and Kp increases.
   4. The reverse reaction will be favored and Kp decreases.
   5. The reverse reaction will be favored and Kp increases.
3. Which observation confirms the fact that equilibrium has been reached in such a system confined in a closed, rigid container?
   1. The density remains constant.
   2. The odor of ammonia can first be detected.
   3. The pressure is decreasing at a constant rate.
   4. The partial pressure of hydrogen remains constant.
   5. The mass of the system has decreased to a constant value.
4. Which occurs when such an equilibrium system is subjected to a stress by the addition of H2(g) and the system proceeds to a new equilibrium?
   * 1. Heat will be released.
     2. Some of the added H2(g) will be consumed.
     3. Some of the N2(g) present originally will be consumed.
   1. ii only
   2. iii only
   3. i and iii only
   4. ii and iii only
   5. i, ii, and iii

***Free Response - Use a calculator and your formula sheets. Timed for 30 min.***

1. The following reaction occurs at 35ºC

* 2 NOCl(g)  ⇄ 2 NO(g)  + Cl2(g)
  1. A 95.3 g sample of NOCl is placed in an evacuated 12.0 L container at 35ºC
     1. What is the initial pressure in atmospheres inside the container before any decomposition occurs?
     2. What is the concentration in moles per liter of NOCl in the container before any decomposition occurs?
  2. If the NOCl is 34.7 percent decomposed when equilibrium is established at 35ºC, calculate the value for the equilibrium constant, *Kc*, for this decomposition reaction.
  3. Calculate *Kp* at 35ºC for this reaction.
  4. The experiment is repeated with another 95.3 g sample of NOCl using the same container. This time the temperature of the reaction vessel is increased to 250ºC. After equilibrium is reestablished, the number of moles at equilibrium is 1.25 for NO and 0.625 for Cl2. Is this reaction endothermic or exothermic? Justify your answer.

1. In a saturated solution of Ag2S at 25ºC, the concentration of S2− is 3.42 x 10−17 molar. The equilibrium is represented by the equation below:
   1. Write the reaction expression for the solubility product constant, *Ksp*, and calculate its value at 25ºC
   2. Calculate the solubility in moles per liter of Ag2S in a 0.00250 M K2S solution.
   3. Predict whether a precipitate of Ag2S will form when 250.0 ml of a 9.0 x 10−3 molar AgNO3 solution is mixed with500.0 ml of a 7 x 10−2 molar K2S solution at 25ºC. Calculations to support your prediction must be shown.
   4. Consider the following system:

| AgCl(s) ⇄ Ag+(aq) + Cl−(aq) | *Ksp* = 1.0 x 10−10 |
| --- | --- |
| * 1. Ag+(aq) + NH3(aq) ⇋ Ag(NH3)+(aq) | *K1* = 2.1 x 10−3 |
| Ag(NH3)+(aq) + NH3(aq) ⇄ Ag(NH3)2+(aq) | *K2* = 8.2 x 103 |

* + 1. Write a balanced summary equation for the above system.
    2. Calculate K for the reaction in (i).
    3. Write the equilibrium expression for this system.

1. c The negative ∆H tells us this is an exothermic reaction, and decreasing the temperature will cause the reaction to shift to the product side to replace the heat. Pressure changes will not cause any shift since there are the same number of moles of gas on both sides of the equation. Adding inert gas would not have any effect as it would increase the total pressure but not change the partial pressure of any of the individual gases. Allowing H2 to escape will decrease a reactant and cause the equilibrium to shift to replace H2 by shifting left. Adding a catalyst will only make both the forward and reverse reactions occur faster, but will not shift the equilibrium position.
2. c It would be wise to make a RICE box for this problem. resulting in the equilibrium values shown in the E row. It is a 1 L container, so moles = molarity. Then plug these values into the equilibrium expression as shown. This will reduce to 0.6 in the denominator which is to say 1/(6/10) which is 10/6 which without a calculator, it would be best to scan the answers and see that c is the obvious choice.
3. d This is a reaction type question. You should realize that silver chloride is insoluble and this alone reduces your choices to only one good option, c.
4. e I will move this to the MC for *Ksp* practice. The reaction for this solubility is PbI2 ⇄ Pb2+ + 2I− this leads to an equilibrium expression of *Ksp* = [Pb2+] [I−]. Considering the values of lead ions and iodide ions at equilibrium, they will be x and 2x thus the equilibrium expression becomes 3.2 x 10−8 = (x)(2x)2 This reduces to 3.2 x 10−8 = 4x3 and this is0.8 x 10−8 = x3 . Look for easy math, this is 8 x 10−9 = x3 which becomes 2 x 10−3.
5. e Each of the reactions represent gaseous equilibria. When the volume of a system changes, at a constant equilibrium, the pressure changes inversely. LeChatelier’s principle will cause the system to shift to relieve the stress. When the number of gaseous moles are the same on both sides, there is no shift in the equilibrium position.
6. a insert the initial values into the equilibrium expression  and solve for Q = 0.2 thus Q < Kp and thus the reaction must proceed in the forward direction. It is likely that the ∆S value is positive since more particles form, and ∆G for this reaction will be negative since the forward reaction occurs.
7. a After equilibrium is established and more CO2 gas is added, the reaction will shift in favor of more reactants to use up the extra carbon dioxide. While the position changes, the *Kp* and *Kc* values do not change. The quantities of the solid constituents change, but their concentrations do not change. When equilibrium is established, ∆G always = 0 (not it’s ∆G = 0, not ∆Gº, which is not likely = 0).

| R | 2A + B ⇄ 2C | | |
| --- | --- | --- | --- |
| I | 0.6 | 0.75 | 0 |
| C | -0.3 | -0.15 | +.3 |
| E | 0.3 | 0.6 | 0.3 |

| R | 2ClNO ⇄ 2NO + Cl2 | | |
| --- | --- | --- | --- |
| I | 0.5 | 0.3 | 0.16 |
| C | -0.1 | -0.1 | -0.05 |
| E | 0.6 | 0.2 | 0.11 |

| R | 2ClNO ⇄ 2NO + Cl2 | | |
| --- | --- | --- | --- |
| I | 0.5 | 0.3 | 0.16 |
| C | -0.1 | -0.1 | -0.05 |
| E | 0.6 | 0.2 | 0.11 |

| R | 2ClNO ⇄ 2NO + Cl2 | | |
| --- | --- | --- | --- |
| I | 0.5 | 0.3 | 0.16 |
| C | -0.1 | -0.1 | -0.05 |
| E | 0.6 | 0.2 | 0.11 |

1. b The exothermic reaction as indicated by the negative ∆H value will shift towards products when the temperature is increased which results in a decrease in the *Keq* value. Decreasing the temperature would cause a shift to the right which would increase *Keq*. Adding an inert gas or a catalyst will cause no shift in the equilibrium position, and adding more hydrogen gas will cause a positional shift to the right, but no change in the Keq value.
2. c Increasing the mass of SO3 will cause a shift to the left. Adding a catalyst will cause no shift, and the other three options all cause shift to the right.
3. e Remember that solids are not included in the equilibrium expression.
4. d We have know information in this problem to tell us where the equilibrium position will land. But since there are only reactants started with, we know that the reaction must proceed in the forward direction, and stoichiometry will prevail, thus we can know that we will lose 2 moles of SO2 for every one mole of O2 which means SO2 will deplete faster resulting in less SO2 than O2 present at equilibrium, wherever it may land.
5. c The information in this problem tells us that at a hight temperature, the equilibrium position is further to the right, or in other words, more product than at cooler temps. This indicates that the reaction is endothermic as the “stress” of the energy will drive the reaction toward products in an “attempt” to use up some of the added heat.
6. e Work a RICE box, I think you will find that there will be 0.11 mole of Cl2 at equilibrium.
7. b First you must change to moles to molarity, then without a calculator, you should be able to come close to 1 x 10−3
8. d Since the reaction has been reversed, the equilibrium constant must be inverted. The inverse of 1.5 which is ³/₂ is equal to ⅔ or 0.67
9. d Since ∆n = 0, then the (RT)∆n factor = 1, thus *Kp* = *Kc*
10. a Remember that solids are not part of the equilibrium expression and should not be included when considering the change form Kc to Kp.
11. d All the substances are gases, thus low temp will favor the reaction toward the energy and increasing pressure will favor the product side with less molecules than the reactant side.
12. c Beware of the solids and liquids since as long as some is present, changing their quantities does not effect equilibrium. Thus a and e have no effect on the equilibrium position. Options b and d will cause the position to shift to the right.
13. a Calculate Q = 1 which is more than Kc, thus the reaction will proceed to the left to achieve equilibrium.
14. c Reversing the reaction requires the inverse of the K, and doubling the coefficients requires squaring of K, thus option c will take care of both.
15. a Looking for a reaction in which ∆n = 0 since *Kp* = *Kc*(RT)∆n and when ∆n is zero, (RT) will equal 1 leaving *Kp* = *Kc*. Beware of solids and liquids as they do not show up in the equilibrium expression and thus should not be counted in ∆n
16. a When the volume is increased, the reaction will shift to the side of more gas molecules.
17. c Addition of a catalyst does not upset equilibrium, it simply would get you to equilibrium faster, if if already there, it would simply speed the rate of the forward and reverse reactions equally , not affecting the equilibrium position.
18. e Remember that K is only constant at a particular temp and any change in temp will cause a shift that will affect the magnitude of K. Thus an increase in volume will cause a shift in the equilibrium shift toward more molecules which is to the right which decreases the N2O4 and will increase the value of K
19. a+d It is important to note that water is in gaseous form, thus it’s concentration can change (unlike when it is in liquid form). Removing N2 will cause the equilibrium position to shift to the right. Options a, b, c, and e will cause the equilibrium to shift to the left. Option a does cause the reaction to shift to the left, however, since the volume is decreased, the concentration has an immediate increase and as the reaction shifts to the right, the mole amount of water will decrease from the dramatic change due to the volume change, but it’s concentration will never get below what it originally started at before the volume decrease.
20. c since at equilibrium, CH4 is less than the starting quantity, thus the reaction must have shifted left, for a loss of H2S and a gain of CS2, and the stoichiometry is 2:1.
21. c in this question, it is important to pay attention to moles and molarity, and the fact that the reaction is done in a 5 L container. The equilibrium concentration for CH4 0.25 M \* 5 L is actually 1.25 mol, this means that there was a loss of 0.25 mole of to reach equilibrium. A Rice Box could help

| R | 4H2(g) + CS2(g) ⇋ CH4(g) + 2H2S(g) | | | |
| --- | --- | --- | --- | --- |
| I | 2.5 mol | 1.5 mol | 1.5 mol | 2 mol |
| C | +1 mol | +0.25 mol | -0.25 mol | -0.5 mol |
| E | 3.5 mol | 1.75 mol | 1.25 mol |  |

| R | 2 NOCl(g)  ⇄ 2 NO(g)  + Cl2(g) | | |
| --- | --- | --- | --- |
| I | 1.46 mol | 0 | 0 |
| C | -0.507 mol | +0.507 mol | +0.253 mol |
| E | 0.953 mol | 0.507 mol | 0.212 mol |



1. e Use the rice box above
2. b Use the rice box above, but don’t forget to change the moles to Molarity
3. c Since the reaction shifts to the left, the forward reaction is not spontaneous, thus + ∆G, and the shift left indicates that Q must be larger than K
4. a The change in temperature will cause a shift left because the reaction is exothermic. The density of the system will not change because the mass of the system will remain constant.
5. d Since the reaction shifted to the left, due to a temperature change, K will become smaller.
6. d When the pressure remains constant, there “shifting to achieve a new equilibrium position” will be completed.
7. e The stress of adding a reactant will cause a shift right which will increase the temp since the reaction is exothermic.
8. ai Change 93.3 g to 1.46 moles, then insert into PV = nRT to get P = 3.08 atm

* aii 1.46 mole/12 L = 0.122 M
* b 34.7% decomposed \* 1.46 moles = 0.507 mole reacted, see the rice box to the right. Change the equilibrium mole values to concentrations and plug into the equilibrium expression.  *Kc* = 6.1 x 10−3
* c Remember that *Kp* = *Kc*(RT)∆n thus *Kp* = (6.1 x 10−3)(0.0821\*308)3-2 thus *Kp* = 0.15
* d At this higher temp, there are more moles of product, indicating the forward reaction was favored, indicating it the endothermic reaction shifted the equilibrium to the right.

1. a For the reaction: Ag2S(s) ⇋ Ag++ S2− *Ksp* = [Ag+]2 [S2−]
   * [Ag+] will be 2x[S2−] , insert the values into the expression, thus *Ksp* = [6.84 x 10−17]2 [3.42 x 10−17] *Ksp* = 1.60 x 10−49

* b 1.60 x 10−49 = [Ag+]2[0.0025], [Ag+] = 8.0 x 10−24 \* 1Ag2S/2Ag+ = 4.0 x 10−24
* c Use MV = MV to solve for the diluted concentrations when the two solutions are mixed. [Ag+] = 0.30M, [S2−] = 0.047M
* solve for Q = 4.2 x 10−5 wich is greater than *Ksp* thus a ppt will form.
* di AgCl(s) + 2NH3 ⇄ Cl− + Ag(NH3)2+
* dii *Krxn = Ksp\*K1\*K2, thus*  = 1.72 x 10−9 diii