MC Practice F2 Solubility Equilibrium, K_{sp}

Name____________________

- ***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.***
- ***DO NOT USE A CALCULATOR. You may use ONLY the periodic table and formula sheets.***
- ***Try to finish at a rate of about 1.5 min per question. Time yourself. It is important that you practice working for speed.***
- ***Please mail me if you think you find any issues or typos.***

1. Temperature is often given with K_{sp} values because
   a. the solubility of solids always increases with increasing temperature.
   b. the solubility of solids varies with temperature changes.
   c. solubility changes with temperature but K_{sp} values do not.
   d. K_{sp} varies with temperature even though concentrations do not.
   e. the number of ions varies with the kind of salt that is dissolving.

2. The K_{sp} expression for silver phosphate is
   a. $K_{sp} = [Ag^+][PO_4^{3-}]$
   b. $K_{sp} = [Ag^+]^2[PO_4^{3-}]$
   c. $K_{sp} = [Ag^+]^3[PO_4^{3-}]$
   d. $K_{sp} = [Ag^+][PO_4^{3-}]^3$
   e. $K_{sp} = \dfrac{[Ag^+][PO_4^{3-}]}{[Ag_3PO_4]}$

3. The solubility of nickel(II) hydroxide, $K_{sp} = 1.6 \times 10^{-16}$, is about
   a. $\sqrt[3]{4.0\times10^{-17}}$
   b. $\sqrt[3]{\dfrac{1.6\times10^{-16}}{2}}$
   c. $\sqrt{1.6\times10^{-16}}$
   d. $\sqrt{(1.6\times10^{-16})(3)}$
   e. $\sqrt{(1.6\times10^{-16})(4)}$

4. What is the molar solubility in water of PbI_2?
   (The K_{sp} for PbI_2 is 3.2×10^{-8})
   a. 3.2×10^{-8} M
   b. 8.0×10^{-8} M
   c. $(1.6)^{1/2} \times 10^{-8}$ M
   d. $(1.6)^{1/3} \times 10^{-8}$ M
   e. 2×10^{-3} M

5. Determine the solubility of lead(II) fluoride, $K_{sp} = 4.0 \times 10^{-8}$, in a 0.0040 M lead(II) nitrate solution.
   a. 2.0×10^{-3} M
   b. 2.0×10^{-2} M
   c. $\sqrt{4.0\times10^{-5}}\,M$
   d. $\dfrac{\sqrt{1.0\times10^{-5}}}{2}\,M$
   e. $\dfrac{4.0\times10^{-8}}{4.0\times10^{-3}}\,M$

6. The solubility of silver sulfide is 8.0×10^{-17} M. Determine the K_{sp} of this salt.
   a. 64×10^{-51}
   b. $[16.0 \times 10^{-17}]\ [8.0 \times 10^{-17}]$
   c. $[16.0 \times 10^{-17})^2\ [8.0 \times 10^{-17}]$
   d. $[8.0 \times 10^{-17}]\ [4.0 \times 10^{-17}]$
   e. $[4.0 \times 10^{-17})^2\ [8.0 \times 10^{-17}]$

7. Equal volumes of 1.6×10^{-5} M KCl and 1.6×10^{-5} M $AgNO_3$ are mixed. the K_{sp} for silver chloride is 1.6×10^{-10}. As these two solutions are combined,
   a. a precipitate of AgCl forms.
   b. there is no precipitate formed.
   c. NaCl will precipitate.
   d. $AgNO_3$ will precipitate.
   e. the $[Na^+]$ will become 0.020 M

8. The solubility of metal(II) sulfide, $K_{sp} = 1.6 \times 10^{-11}$, is about
   a. 4.0×10^{-6}
   b. $\sqrt[3]{4}\times10^{-4}$
   c. $\sqrt[3]{1.6\times10^{-11}}$
   d. 8.0×10^{-6}
   e. 2.0×10^{-6}

9. Calculate the pH of a saturated metal hydroxide, XOH, whose $K_{sp} = 1 \times 10^{-8}$.
   a. 4.0
   b. 6.0
   c. 7.0
   d. 8.0
   e. 10.0

10. Calculate the K_{sp} of a saturated metal hydroxide, $X(OH)_2$, solution whose pH = 9.00.
   a. 2.0×10^{-28}
   b. 5.0×10^{-16}
   c. 2.0×10^{-15}
   d. 2.0×10^{-10}
   e. 5.0×10^{-10}

11. On the basis of the K_{sp} values listed below, what is the order of least soluble to most soluble for these compounds?

   AgBr, $K_{sp} = 5.4 \times 10^{-13}$

   Ag_2CO_3, $K_{sp} = 8.0 \times 10^{-12}$

   AgCl, $K_{sp} = 1.8 \times 10^{-10}$

   a. $AgBr < Ag_2CO_3 < AgCl$
   b. $AgBr < AgCl < Ag_2CO_3$
   c. $AgCl < Ag_2CO_3 < AgBr$
   d. $Ag_2CO_3 < AgBr < AgCl$

12. On the basis of the solubilities (in mol/L) listed below, what is the correct order of decreasing K_{sp} values (largest first).

   AgBr 7.3×10^{-7}

   AgCN 7.7×10^{-9}

   AgSCN 1.0×10^{-6}

   a. AgBr > AgCN > AgSCN
   b. AgCN > AgBr > AgSCN
   c. AgCN > AgSCN > AgBr
   d. AgSCN > AgBr > AgCN

13. Magnesium hydroxide, $Mg(OH)_2$, has a $K_{sp} = 6.3 \times 10^{-10}$. The solubility of $Mg(OH)_2$ will be the lowest in 1.0 L of which of the following?
   a. 0.10 M HCl
   b. 0.10 M NaOH
   c. 0.10 M $MgCl_2$
   d. pure H_2O

Use the following information to answer the next 3 questions

> 150 ml of saturated SrF_2 solution is present in a 250 ml beaker at room temperature. The molar solubility of SrF_2 at 25ºC is 1×10^{-3} M

14. What are the concentrations of Sr^{2+} and F^- in the beaker?
   a. $[Sr^{2+}] = 1.0 \times 10^{-3}$ M, $[F^-] = 1.0 \times 10^{-3}$ M
   b. $[Sr^{2+}] = 1.0 \times 10^{-3}$ M, $[F^-] = 2.0 \times 10^{-3}$ M
   c. $[Sr^{2+}] = 2.0 \times 10^{-3}$ M, $[F^-] = 1.0 \times 10^{-3}$ M
   d. $[Sr^{2+}] = 2.0 \times 10^{-3}$ M, $[F^-] = 2.0 \times 10^{-3}$ M

15. If some of the solution evaporates overnight, which of the following will occur?
   a. The mass of the solid and concentration of the ions will stay the same.
   b. The mass of the solid and concentration of the ions will stay the increase.
   c. The mass of the solid will decrease, and concentration of the ions will stay the same.
   d. The mass of the solid will increase, and concentration of the ions will stay the same.

16. How could the concentration of Sr^{2+} ions in solution be decreased?
   a. Adding some $NaF_{(s)}$ to the beaker.
   b. Adding some $Sr(NO_3)_{2(s)}$ to the beaker.
   c. By heating the solution in the beaker.
   d. By adding a small amount of water to the beaker, but not dissolving all of the solid.

17. $2HI_{(g)} + Cl_{2(g)} \leftrightharpoons 2HCl_{(g)} + I_{2(g)} + \text{energy}$

A gaseous reaction occurs and comes to equilibrium as shown above. Which of the following changes to the system will serve to increase the number of moles of I_2 present at equilibrium?

a. Increasing the volume at constant temperature.
b. Decreasing the volume at constant temperature.
c. Adding a mole of inert gas at constant volume.
d. Increasing the temperature at constant volume.
e. Decreasing the temperature at constant volume.

18. In the reaction shown below, 0.30 mole of Br_2 and 0.30 moles of I_2 are placed in an evacuated 1.00 liter vessel and allowed to reach equilibrium. What is the value of K_c if the vessel contains 0.40 moles of IBr at equilibrium?

$$I_{2(g)} + Br_{2(g)} \leftrightharpoons 2IBr_{(g)}$$

a. 1.0
b. 2.0
c. 4.0
d. 16
e. 40

19. Which is the correct set up to determine K_c for the reaction below?

$$4CuO_{(s)} + CH_{4(g)} \leftrightharpoons CO_{2(g)} + 4Cu_{(s)} + 2H_2O_{(g)}$$

a. $\dfrac{[CO_2][Cu]^4[H_2O]^2}{[CuO]^4[CH_4]}$

b. $\dfrac{[CuO]^4[CH_4]}{[CO_2][Cu]^4[H_2O]^2}$

c. $\dfrac{[CO_2][H_2O]^2}{[CH_4]}$

d. $\dfrac{[CO_2][H_2O]^2}{[CuO]^4[CH_4]}$

e. $\dfrac{[CO_2]^2[Cu]^4[H_2O]^2}{[CuO]^4[CH_4]}$

20. An evacuated 1.00-liter vessel is injected with 1.2 moles of sulfur trioxide gas, SO_3. The vessel is then heated to a high temperature where the SO_3 partially decomposes to form the products SO_2 and O_2 in the reaction shown below:

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

The temperature is kept constant, and the amount of SO_3 in the vessel at equilibrium is 0.40 mol. What is the value of K_c at this temperature?

a. 0.04
b. 0.16
c. 0.8
d. 1.0
e. 1.6

21. A sample of solid potassium nitrate is placed in water. The solid potassium nitrate comes to equilibrium with its dissolved ions by the endothermic process shown below.

$$KNO_{3(s)} + \text{energy} \leftrightharpoons K^+_{(aq)} + NO_3^-{}_{(aq)}$$

Which of the following changes to the system would increase the concentration of K^+ ions at equilibrium?

a. The volume of the solution is increased.
b. The volume of the solution is decreased.
c. Additional solid KNO_3 is added to the solution.
d. The temperature of the solution is increased.
e. The temperature of the solution is decreased

22. A 3.00-liter reaction vessel is filled with carbon monoxide gas, CO, and chlorine gas, Cl_2. The mixture is heated to 670 K and allowed to reach equilibrium according to the balanced equation shown below:

$$CO_{(g)} + Cl_{2(g)} \leftrightharpoons COCl_{2(g)}$$

At equilibrium, the mixture contains 0.3 moles CO, 0.5 moles Cl_2, and 1.5 moles $COCl_2$. What is K_c at this temperature?

a. 0.03
b. 1
c. 3
d. 10
e. 30

23. The reaction below is allowed to come to equilibrium at room temperature. At equilibrium, the partial pressure of H_2O is 296 mm Hg, Cl_2O is 15 mm Hg, and HOCl is 20 mm Hg. What is the value of K_p at this temperature?

$$H_2O_{(g)} + Cl_2O_{(g)} \leftrightharpoons 2HOCl_{(g)}$$

a. 222
b. 11
c. 0.017
d. 0.090
e. 0.0045

24. At 373 K, the reaction shown below has an equilibrium constant, $K_c = 2.19 \times 10^{-10}$.

$$COCl_2(g) \leftrightharpoons CO(g) + Cl_2(g)$$

After placing a mixture of gases in the reaction vessel, the concentrations were measured to be $[COCl_2] = 3.50 \times 10^{-3}$ M, $[CO] = 1.11 \times 10^{-5}$ M, and $[Cl_2] = 3.25 \times 10^{-6}$ M. Which statement below accurately describes the reaction?

a. The reaction is at equilibrium
b. The reaction is not at equilibrium, and it is proceeding to the left
c. The reaction is not at equilibrium, and it is proceeding to the right
d. The reaction quotient is equal to K_c
e. The reaction quotient is less than K_c

25. In the reaction $3W + X \leftrightharpoons 2Y + Z$, all substances are gases. The reaction is initiated by adding equal number of moles of W and of X. When equilibrium is reached,

a. [Y] = [Z]
b. [X] = [Y]
c. [W] = [X]
d. [X] > [W]
e. [W] + [X] = [Y] + [Z]

26. Consider the reactions below, both at the same temperature:

Rx I $X \leftrightharpoons Y$ $K = 1 \times 10^8$

Rx II $Z \leftrightharpoons Y$ $K = 1 \times 10^5$

a. I is 3 times faster than II.
b. I is 1000 times faster than II.
c. II is 3 times faster than I.
d. II is 1000 times faster than I.
e. The size of K and the time required to reach equilibrium are not directly related.

27. The reaction $3H_{2(g)} + N_{2(g)} \leftrightharpoons 2NH_{3(g)}$ is exothermic.

Increasing the temperature of this equilibrium system causes

a. an increase in $[NH_3]$
b. an increase in $[N_2]$
c. a decrease in $[H_2]$
d. an increase in K
e. a decrease in pressure at constant volume

28. Consider $N_{2(g)} + O_{2(g)} \leftrightharpoons 2NO_{(g)}$. The reaction was initiated by adding 15.0 moles of NO to a 1.0-L flask. At equilibrium, 3.0 moles of oxygen are present in the 1.0-L flask. The value of K must be

a. 0.33
b. 3.0
c. 5.0
d. 9.0
e. 81

29. At a certain temperature, the synthesis of ammonia gas from nitrogen and hydrogen gases, shown below, has a value for K of 3.0×10^{-2}.
If $[H_2] = [N_2] = 0.10$ M and $[NH_3] = 0.20$ M

$$N_2 + 3H_2 \leftrightharpoons 2NH_3$$

a. the reaction would shift toward the ammonia
b. the reaction would shift toward the N_2 and the H_2
c. the system is at equilibrium, therefore no shifting will occur
d. the reaction will shift toward a new equilibrium position, but the direction cannot be determined from these data
e. the equilibrium may shift but it is not possible to calculate Q without knowing the temperature

30. The equilibrium constant, K, may be used to determine the K of other reactions.

I. When the equation is reversed the reciprocal of the original K becomes the value of K for the new equation.
II. When the equation is doubled the square root of the original K becomes the value for the new equation.
III. When reactions are added, the sum of the equilibrium constants of each reaction added will the equilibrium constant of the new equation.

Of the above three statements, those which are always true are

a. I only
b. II only
c. III only
d. I and III
e. I, II, and III

31. The equilibrium $P_{4(g)} + 6Cl_{2(g)} \leftrightharpoons 4PCl_{3(L)}$ is established at −10°C. The equilibrium constant expression is
    a. $K = \frac{[PCl_3]}{[P_4][Cl_2]}$
    b. $K = \frac{[PCl_3]^4}{[P_4][Cl_2]^6}$
    c. $K = \frac{[P_4][Cl_2]^6}{[PCl_3]^4}$
    d. $K = \frac{1}{[PCl_3]^4}$
    e. $K = \frac{1}{[P_4][Cl_2]^6}$

32. Ammonium hydrogen sulfide will decompose into ammonia gas and hydrogen sulfide gas when heated. Consider the equilibrium system

    $$NH_4HS_{(s)} \leftrightharpoons NH_{3(g)} + H_2S_{(g)}$$

    which is developed from 1.000 mole of NH_4HS in a 100-L cylinder. At equilibrium the total pressure is found to be 0.400 atm. K_p will be equal to
    a. 2.00×10^{-1}
    b. 1.00×10^{-2}
    c. 4.00×10^{-2}
    d. 4.00
    e. a value impossible to calculate from only these data

33. At a certain temperature, it has been determined that $K = 8.0$ for the reaction below. If we also determine that the equilibrium mixture contains 0.80 mole of $H_2O_{(g)}$, 0.080 mole of $CO_{2(g)}$, and 0.080 mole of $CO_{(g)}$, in an 8.0-L flask, what must be the number of moles of $H_{2(g)}$ at equilibrium?

    $$H_2O_{(g)} + CO_{(g)} \leftrightharpoons CO_{2(g)} + H_{2(g)}$$

    a. 0.010 mole
    b. 0.80 mole
    c. 6.4 moles
    d. 8.0 moles
    e. 64 moles

34. Which of the following systems at equilibrium are not affected by a change in pressure caused by changing the volume at constant temperature?
    a. $H_{2(g)} + Cl_{2(g)} \leftrightharpoons 2HCl_{(g)}$
    b. $H_{2(g)} + I_{2(s)} \leftrightharpoons 2HI_{(g)}$
    c. $N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$
    d. $2NH_{3(g)} \leftrightharpoons N_{2(g)} + 3H_{2(g)}$
    e. $3O_{2(g)} \leftrightharpoons 2O_{3(g)}$

35. After the equilibrium represented below is established, some pure $O_{2(g)}$ is injected into the reaction vessel at constant temperature. After equilibrium is reestablished, which of the following has a lower value compared to its value at the original equilibrium?

    $$2CO_{2(g)} \leftrightharpoons 2CO_{(g)} + O_{2(g)}$$

    a. K_{eq} for the reaction
    b. The total pressure in the reaction vessel
    c. The amount of $CO_{2(g)}$ in the reaction vessel
    d. The amount of $O_{2(g)}$ in the reaction vessel
    e. The amount of $CO_{(g)}$ in the reaction vessel

36. Which of the following changes alone would cause a decrease in the value of K_{eq} for the reaction represented below?

    $$2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)} \qquad \text{Exothermic}$$

    a. Decreasing the temperature
    b. Increasing the temperature
    c. Decreasing the volume of the reaction vessel
    d. Increasing the volume of the reaction vessel
    e. Adding a catalyst

37. In which of the following systems would the number of moles of the substances present at equilibrium NOT be shifted by a change in the volume of the system at a constant temperature?
    a. $2SO_{2(g)} + O_{2(g)} \leftrightharpoons 2SO_{3(g)}$
    b. $N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$
    c. $NO_{2(g)} + SO_{2(g)} \leftrightharpoons SO_{3(g)} + NO_{(g)}$
    d. $N_2O_{4(g)} \leftrightharpoons 2NO_{2(g)}$
    e. $CO_{(g)} + 3H_{2(g)} \leftrightharpoons CH_{4(g)} + H_2O_{(g)}$

38. For: $N_{2(g)} + 3H_{2(g)} \leftrightharpoons 2NH_{3(g)}$ + energy

    Some N_2 and H_2 are mixed in a container at 200 °C, and the system reaches equilibrium according to the equation above. Which of the following causes an increase in the number of moles of NH_3 present at equilibrium?

    I. Decreasing the volume of the container
    II. Raising the temperature
    III. Adding a mole of Ar gas at constant volume

    a. I only
    b. II only
    c. I and III only
    d. II and III only
    e. I, II, and III

1. b All K values always change with varying temperature. Whether or not the equilibrium constant is larger or smaller is a function of whether the reaction is exothermic or endothermic. If the reaction is exothermic *K* will be smaller value at higher temperatures, and a larger value for lower temperatures. The converse is also true.

2. c The formula for silver phosphate is Ag_3PO_4 which leads to the reaction $Ag_3PO_4 \leftrightharpoons 3Ag^+ + PO_4^{3-}$

3. a The formula for nickel(II) hydroxide is $Ni(OH)_2$ which leads to the reaction $Ni(OH)_2 \leftrightharpoons Ni^{2+} + 2OH^-$ and the resulting solubility expression is $K_{sp} = [Ni^{2+}][OH^-]^2$.

   Since there are 2 OH^- ions for every Ni^{2+} ion upon dissolving, we can identify Ni^{2+} as "s" and OH^- as "2s" and substitute into the K_{sp} expression.

   Thus $K_{sp} = (s)(2s)^2$ and $K_{sp} = 4s^3$ solve for s $s = \sqrt[3]{\frac{1.6\times10^{-16}}{4}} = \sqrt[3]{\frac{16\times10^{-17}}{4}} = \sqrt[3]{4.0\times10^{-17}}$ without a calculator, we can't really clean this up any further.

   Since s = concentration of the Ni^{2+}, the 1:1 stoichiometry the concentration of the $Ni(OH)_2$ that actually dissolves, "s" would be the molar solubility.

4. e This problem is nearly the same as #3, however, this time the numbers are simple enough to finish the solve.

   The formula for lead(II) iodide is PbI_2 which leads to the reaction $PbI_2 \leftrightharpoons Pb^{2+} + 2\ I^-$ and the resulting solubility expression is $K_{sp} = [Pb^{2+}][I^-]^2$.

   Since there are 2 I^- ions for every Pb^{2+} ion upon dissolving, we can identify Pb^{2+} as "s" and I^- as "2s" and substitute into the K_{sp} expression.

   Thus $K_{sp} = (s)(2s)^2$ and $K_{sp} = 4s^3$ solve for s $s = \sqrt[3]{\frac{3.2\times10^{-8}}{4}} = \sqrt[3]{\frac{32\times10^{-9}}{4}} = \sqrt[3]{8\times10^{-9}} = 2\times10^{-3}$

   Since s = concentration of the Pb^{2+}, the 1:1 stoichiometry the concentration of the PbI_2 that actually dissolves, this "s" would be the molar solubility. (If you needed a gram solubility, you would have to multiply by the molar mass of PbI_2.)

5. d This problem has a twist compared to #'s 3 and 4 because instead of dissolving the "insoluble salt," lead (II) fluoride in water, it is being dissolved in a solution that already contains some lead ions, lead(II) nitrate. We call this type of problem a "common ion" problem, meaning the solution you are dissolving into already has an ion in it that is "common" to the salt that you are trying to dissolve. More information on this in your text on page 681.

   The formula for lead(II) fluoride is PbF_2 which leads to the reaction $PbF_2 \leftrightharpoons Pb^{2+} + 2F^-$ and the resulting K_{sp} expression is $K_{sp} = [Pb^{2+}][F^-]^2$. Up till now we have substituted "s" and "2s" into this equation, and solved for "s" which will be the solubility.

   Since you are given a concentration of the Pb^{2+} ion that is already dissolved in the water (and is much larger, when considering significant figures, than any lead ions that would be added to it) which you can insert directly into the K_{sp} expression. Thus $K_{sp} = (0.004)[2s]^2 = 4.0\times10^{-8}$ Solve: $4s^2 = \frac{4.0\times10^{-8}}{4\times10^{-3}} = 1\times10^{-5} = \sqrt{\frac{1\times10^{-5}}{4}} = \frac{\sqrt{1\times10^{-5}}}{2}$

   Since s = concentration of the Pb^{2+}, the 1:1 stoichiometry the concentration of the PbF_2 that actually dissolves, this "s" would be the molar solubility. (If you needed a gram solubility, you would have to multiply by the molar mass of PbF_2.)

6. c When you are told the solubility of some salt, you are being told the maximum amount of the compound that dissolves to achieve a saturated solution. For silver sulfide, Ag_2S, the solubility of 8.0×10^{-17}, this tells you the concentration of the Ag_2S that dissolves, which is equal to the sulfide ion that forms, and the silver ion is twice that due to the dissolving reaction, $Ag_2S \leftrightharpoons 2Ag^+ + S^{2-}$ thus $K_{sp} = [Ag^+]^2[S^{2-}]$. Substitute $K_{sp} = \left[16\times10^{-17}\right]^2\left[8\times10^{-17}\right]$

7. b In order to answer this question, you need to calculate Q and compare it to K_{sp}. For this combination, the possible precipitate is AgCl, thus the K_{sp} reaction of interest is $AgCl \leftrightharpoons Ag^+ + Cl^-$ and $K_{sp} = [Ag^+][Cl^-]$. The "trick" in this problem is to remember that when solutions combine, the concentrations of the ions will be diluted because there is more volume for the same number of moles. Since equal volumes are combined, each concentration is halved. $[Ag^+] = [Cl^-] = 0.8\times10^{-5}$, or 8×10^{-6}, So $Q = [8\times10^{-6}][8\times10^{-6}]$ which results in $Q = 64\times10^{-12}$ or 6.4×10^{-11} In effect, reaction shifts to the right and no precipitate forms when $K_{sp} > Q$.

8. a The formula for metal(II) sulfide is MS which leads to the reaction MS $\leftrightarrows$ $M^{2+} + S^{2-}$ and the resulting K_{sp} expression is $K_{sp} = [M^{2+}][S^{2-}]$ and thus $K_{sp} = s^2$ Change K_{sp} to 16×10^{-12} to make it easier without a calculator, and then solve for “s” which will be the solubility.

9. e The formula for metal hydroxide, XOH leads to the reaction XOH $\leftrightarrows$ $X^+ + OH^-$ and the resulting K_{sp} expression is $K_{sp} = [X^+][OH^-]$ and thus $K_{sp} = s^2$ Solve for “s” which will be the solubility, $s = 1 \times 10^{-4}$, and then take the pOH = $-\log[OH^-] = 4$, then “14-it” to solve for pH = 10.

10. b The formula for metal hydroxide is $X(OH)_2$ which leads to the reaction $M(OH)_2 \leftrightarrows M^{2+} + 2OH^-$ and the resulting K_{sp} expression is $K_{sp} = [M^{2+}][OH^-]^2$. pH of 9 should be converted to a pOH = 5, which can be unlogged to $[OH^-] = 1 \times 10^{-5}$. Due to the stoichiometry of the reaction above, the $[M^{2+}] = 0.5 \times 10^{-5}$. Now insert the values back into the K_{sp} = formula. $K_{sp} = [0.5 \times 10^{-5}][1 \times 10^{-5}]^2$

11. b It is tempting to just look at the K_{sp} and assume that the smallest K_{sp} means the lowest solubility, however, you must consider the amount of ions that the salt dissociates into. Since the Ag_2CO_3 salt has 3 ions, $K_{sp} = 4s^3$, whereas for the other two salts, the $K_{sp} = s^2$. Thus for AgBr $s = \sqrt{54 \times 10^{-14}} =\sim 7 \times 10^{-7}$, for AgCl $s = \sqrt{1.8 \times 10^{-10}} =\sim 1 \times 10^{-5}$, but for $4s^3 = 8 \times 10^{-12}$, *thus* $\sqrt[3]{2 \times 10^{-12}} =\sim 1 \times 10^{-4}$

12. d There is no need to make this problem any harder than necessary. Since all three salts are 2 ion salts, calculating the K_{sp} will be a simple squaring of the solubility. Thus the decreasing order of the K_{sp} values will be in the same order as the solubility values.

13. b HCl will react with the OH^- ion, lowering the OH^-, causing a shift to the right resulting in greater solubility, thus not the answer. This problem requires you to pick out the substance with a common ion, and both NaOH and $MgCl_2$ will satisfy that criteria. So there is a twist. Remember the solubility can be determined by solving for “s” from the equilibrium expression. $K_{sp} = \left[Mg^{2+}\right]\left[OH^-\right]^2$. For the introduction of 0.1 M $MgCl_2$, you can simply substitute the concentration of $[Mg^{2+}]$ into the equilibrium expression, and the equation to solve for “s” would look like this $6.3 \times 10^{-10} = \left[0.1\right]\left[2s\right]^2$. Divide by 0.1 for, 6.3×10^{-9}, convert to 63×10^{-10}, then divide by 4, to get $\sim 16 \times 10^{-10}$ and square root to solve for $s = \sim 4 \times 10^{-5}$.

For the introduction of 0.1 M $Mg(OH)_2$, you can simply substitute the concentration of $[OH^-]$ into the equilibrium expression, and the equation to solve for “s” would look like this $6.3 \times 10^{-10} = \left[s\right]\left[0.1\right]^2$. Square 0.1, then divide K_{sp} by 0.01, and solve for $s = 6.3 \times 10^{-8}$. It should now be obvious that the squared effect of the concentration of the common ion $[OH^-]$ causes the **NaOH to have the greater effect on solubility**.

14. b Remember that solubility describes the amount of the chemical that does dissolve. Since SrF_2 dissolves and dissociates according to the following equation, $SrF_2 \rightleftarrows Sr^{2+} + 2\,F^-$ the concentration of Sr^{2+} will be the same as the solubilty, and the concentration of F^- will be 2 × the solubility.

15. d When evaporation occurs, only water will evaporate, the salt will stay in the solution. Since there can NOT be a greater concentration in solution, that the saturation quantity, precipitation will have to occur to maintain the saturation concentrations.

16. a Addition of $NaF_{(s)}$ provides the common ion $[F^-]$,which will serve to cause the equilibrium to shift left, thus reducing the $[Sr^{2+}]$. Addition of $Sr(NO_3)_{2(s)}$ would also cause a shift to the left, but the addition of the $[Sr^{2+}]$ ions will increase $[Sr^{2+}]$, and while the $[F^-]$ will shift below the original $[F^-]$, the $[Sr^{2+}]$ will never be able to reduce below the original $[Sr^{2+}]$ concentration before the addition of $Sr(NO_3)_{2(s)}$.

17. e Changing the volume will NOT cause any shift in the equilibrium position since Q will not change due to the fact that the stoichiometry of the gas molecules is the same on both sides of the equation. Just as adding an inert gas will not Choices c” would all cause no shift. Choice “d” would cause a shift to the left. Thus “e” is the correct choice.

18. d This is a classic RICE box and then solve for K_c. It is true that the math of the K_c calculation is not so simple, yet the answers are quite spread, so approximations make it possible. So, noting that the container is 1.0 L, plug in the known info in the problem (in blue) then calculate the rest of the items with stoichiometry.

R	$I_{2(g)} + Br_{2(g)} \rightleftharpoons 2\ IBr_{(g)}$		
I	**0.3 M**	**0.3 M**	0
C	−0.2 M	−0.2 M	+0.2 M
E	0.1 M	0.1 M	**0.4 M**

$K_c = \frac{[0.4M]^2}{[0.1M][0.1M]}$ to do the easy math revise this $K_c = \frac{[0.4]^2}{[0.1][0.1]}$ you can

which will then clean up to $K_c = \frac{0.16}{1\times10^{-2}}$ which is 16, choose "d"

19. c In this problem, it is important to remember that solids (nor liquids) should be included in the solubility expression.

20. e This is a classic RICE box and then solve for K_c. So, noting that the container is 1.0 L, plug in the known info in the problem (in bold) then calculate the rest of the items with stoichiometry.

R	$2SO_{3(g)} \rightleftharpoons 2SO_{2(g)} + O_{2(g)}$		
I	**1.2 M**	**0**	**0**
C	−0.8 M	+0.8 M	+0.4 M
E	**0.4 M**	0.8 M	0.4 M

$$K_c = \frac{[0.8M]^2[0.4M]}{[0.4M]^2} \Rightarrow \frac{[0.8M]^2}{[0.4M]} \Rightarrow \frac{[0.8M][0.8M]}{[0.4M]} \Rightarrow \frac{[2M][0.8M]}{} = 1.6$$

21. d Increasing the temperature of an endothermic process will cause a shift toward products, thus causing more ions to go into solution.

22. e For this problem there is no reason to do a rice box since you were given the equilibrium mole values. It is important to note that the container is 3.0 L. I plugged in the moles and the volume directly into the equilibrium expression for the equation

$CO_{(g)} + Cl_{2(g)} \rightleftharpoons COCl_{2(g)}$ $K_c = \frac{[COCl_2]^2}{[CO][Cl_2]}$ Yikes, the math looks to be a mess, $K_c = \frac{[\frac{1.5}{3}]}{[\frac{0.3}{3}][\frac{0.5}{3}]}$ and you don't have a

calculator, so look for factors that are easy to cancel and end up with 30

23. d Again, for this problem there is no reason to do a rice box since you were given the equilibrium pressure values. I plugged in the moles and the volume directly into the equilibrium expression for the equation $H_2O_{(g)} + Cl_2O_{(g)} \rightleftharpoons 2HOCl_{(g)}$

$K_c = \frac{[HOCl]^2}{[H_2O][ClO_2]}$ This time the math isn't so bad. $K_c = \frac{[20]^2}{[15][296]}$ which approximates to $K_c = \frac{[400]}{[15][\sim 300]}$ drop some zeros

and call it $K_c = \frac{[4]}{[15][\sim 3]}$ which is $K_c = \frac{[4]}{[45]}$ which is something just smaller than 1/10th or choice "d"

24. b For this problem, you must plug into the equilibrium expression and solve for $Q = \frac{[3.25\times10^{-6}][1.11\times10^{-5}]}{[3.5\times10^{-3}]}$ you can do some estimating since you probably will be comparing orders of magnitude. I cancelled out the 3.25 and 3.5 to get

$Q = \frac{[1\times10^{-6}][1.11\times10^{-5}]}{[1\times10^{-3}]}$ Remember that 10^{-3} in the denominator is 10^{3} in the numerator, thus

$Q = (1 \times 10^{-6})(1.11 \times 10^{-5})(1 \times 10^{3})$ which is $K_{sp} < Q = (1 \times 10^{-8})$ given in the problem, thus the equilibrium will shift left to reduce the value of Q.

25. d Starting with only W and X, Y could never be equal to Z due to the 2:1 stoichiometry. We do not have enough information to know where the equilibrium position will land, so we cannot compare X to Z. W could not be equal to X because of starting with equal quantities of them, and the 3:1 ratio between them, thus we can tell that X will be greater than W because W reacts 3x the rate of X. Choice "e" is certainly not what equilibrium means.

26. e rate laws and rate *k* constants gives us info about speed. K_{eq} tells us nothing about speed of reaction, only where the "position" ends up.

27. b Since we know from the reaction is exothermic, think of energy on the right side. Raising the temp of the system, and the reaction will shift to use some of this incoming heat, that's left.

28. d It is a 1.0 L flask, so moles are equal to molarity. This is easy math compared to those others. $K_c = \frac{[9]^2}{[3][3]}$

29. b Plug in values to solve for $Q = \frac{[0.2]^2}{[0.1][0.1]^3}$ which is $Q = \frac{[0.04]}{[0.1][.001]}$ which is $Q = \frac{[4\times10^{-2}]}{[1\times10^{-4}]}$ which is 400 which is greater than the K given in the problem and thus the shift will be left. (Note: If you have trouble multiplying or squaring decimals, turn 0.02 into 2×10^{-1} which is might be easier to cope with.)

30. a Only I is true. For II when the stoichiometry of a reaction is changed, the equilibrium constant of the original equation should be raised to the power of the stoichiometry change. For III, when reactions are added, the product of the equilibrium constants of each reaction added will the equilibrium constant of the new equation.

31. e Note that PCl_3 in this problem is a liquid, so it should not be included in the equilibrium expression.

32. c In this problem, take note that the reactant is a solid. The product gases come only from the solid reactant, thus the two products must be present in equal quantities, this allows the total equilibrium pressure to be cut in half to get each gases partial pressure, 0.2 atm for each gas. Then you can square the partial pressures to solve for K_p.= (0.2)(0.2)

33. c Plug into the K_c expression and solve for $[H_2]$ $8 = \frac{[0.01][H_2]}{[0.01][0.1]}$, thus $[H_2] = 0.8$ and multiply by 8.0 L to get 6.4 moles.

34. a You are looking for a reaction in which the number of moles of gases are the same on both sides.

35. e If oxygen is injected at equilibrium, “stress” has been applied and the system must shift to restore equilibrium. Thus a shift toward reactants will occur, which will reduce the quantity of carbon monoxide.

36. b Since we know that the reaction is exothermic, think of energy on the right side. Raising the temperature of the system, and the reaction will shift to use some of this incoming heat, that’s left, thus reducing the value of *K*, since *K* is constant only for one particular temperature.

37. c You are looking for a reaction in which the number of moles of gases are the same on both sides. (Note there was a typo in the original version of the test in choice “a”

38. a Decreasing the volume of the container will increase the partial pressure of all the gases and cause the equilibrium to shift toward the products where there are less gas molecules. Raising the temp would cause a shift left, and adding argon would have no effect on partial pressures (even though the total pressure is increased), and thus would create no shift at all.