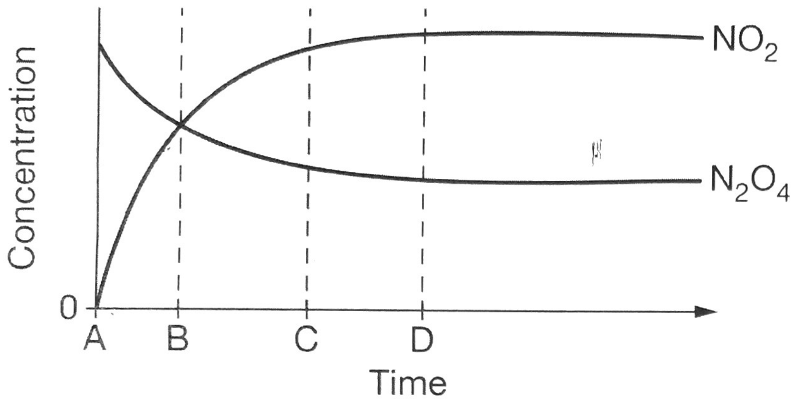
Name: KEY
Date:

Unit 7 Review: Equilibrium PPC

Question 1



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

The graph above represents the data collected under certain conditions for the decomposition of $N_2O_4(g)$ according to the chemical equation above. Based on the graph, at approximately which time is equilibrium established?

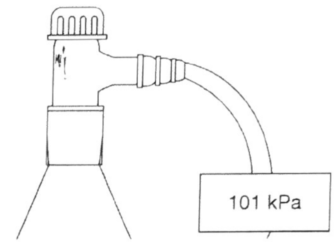
A At time A, because $N_2O_4(g)$ is expanding to fill the container.

B At time B, because the reaction is reversible and $[NO_2] = [N_2O_4]$.

C At time C, because the reaction is about to reach completion and $[NO_2] > [N_2O_4]$.

D At time D, because there are no observable changes in $[NO_2]$ and $[N_2O_4]$.

Question 2



Time (min)	Pressure (kPa)
0.25	101
0.50	108
0.75	116
1.00	124
1.25	128
1.50	130

130
130
130

A sample of acetone is placed into a container. The container is sealed and attached to a pressure sensor, as shown in the diagram above. The container is allowed to sit on the lab table for a few minutes as the pressure in the container is monitored at regular intervals. At the end of 2.00 minutes, some acetone liquid remains in the container. Which of the following best explains the pressure data presented in the table above?

A The acetone heats up over time, causing more of it to vaporize at an increasing rate. *(No heat added)*

B The acetone has completely vaporized after 1.50 minutes, so the pressure becomes constant. *(some liquid remains)*

C The acetone vaporizes from the liquid at a constant rate and the rate of condensation increases until it becomes equal to the rate of evaporation and the pressure stays constant.

D The acetone vaporizes from the liquid at a rate that is fast in the beginning but then slows down until the vaporization process stops completely. *(not true will continue during equilibrium)*

Question 3

2 gas → 1 gas

Brown Colorlesss

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

A sample of pure $NO_2(g)$ in a sealed tube at $20\,^{\circ}C$ is placed in a temperature bath at $30\,^{\circ}C$. Observations of changes in the color, pressure, and mass of the mixture are recorded as a function of time. Which of the following is an observation that would best support the claim that the reaction represented above has reached equilibrium at $30\,^{\circ}C$?

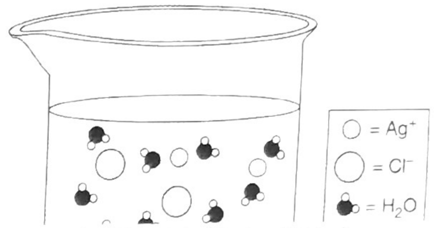
A. The total mass of the system remains constant because at equilibrium the amounts of reactant and product do not change with time. True... but wouldn't tell us when it's @ equilibrium mass is always constant in this case

B. The color of the system changes from brown to completely colorless because at equilibrium only the product will be present. not true

C. The total pressure of the system decreases then reaches a constant value because at equilibrium the amounts of reactant and product no longer change.

D. The temperature of the system remains constant because at equilibrium the temperature must be constant. only due to temp bath

Question 4



$AgCl(s) \rightleftarrows Ag^+(aq) + Cl^-(aq)$

When $AgCl(s)$ is placed in water, it dissolves according to the chemical equation above. The particle diagram above was proposed to represent an aqueous solution in which $AgCl(s)$ is in equilibrium with its ions. Which of the following best explains whether or not the diagram provides a good representation of this dynamic equilibrium at the microscopic level?

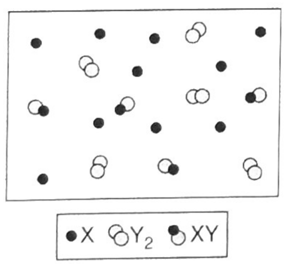
A. It is a good representation because it shows that the dipoles of the H_2O molecules are oriented around the Ag^+ and Cl^- ions in solution. irrelevant to equilibrium

B. It is a good representation because it shows that the concentrations of Ag^+ ions and Cl^- ions are equal at equilibrium. Doesn't have to be true

C. It is not a good representation because it does not show that the concentration of $AgCl(s)$ is constant at equilibrium. ? solid's don't have a [ ]

D. It is not a good representation because it does not illustrate the dynamic equilibrium in which the rates of the forward and reverse reactions are equal. → key to equilibrium

Question 5



The particle diagram above shows the system represented by the equation $2\,X(g) + Y_2(g) \rightleftarrows 2\,XY(g)$. Which of the following explains whether the particle diagram indicates that the system is at equilibrium?

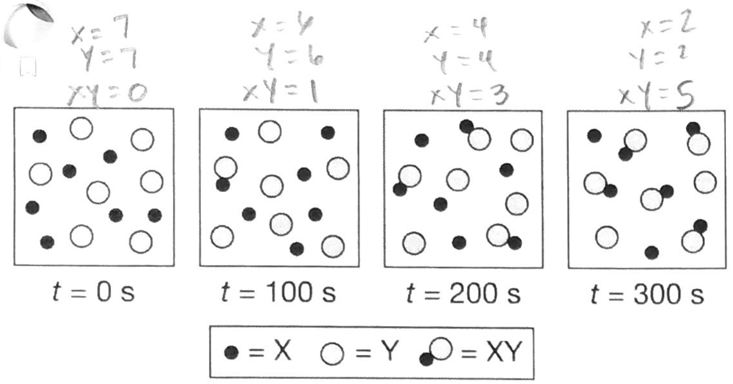
A. The particle diagram does not indicate that the system is at equilibrium because it shows the system only at one point in time.

B. The particle diagram does not indicate that the system is at equilibrium because the ratio $\frac{[XY]_{eq}}{[X]_{eq}}$ is not equal to 1.

C. The particle diagram indicates that the system is at equilibrium because the value of K is small.

D. The particle diagram indicates that the system is at equilibrium because $[X]_{eq} = 2 \times [Y_2]_{eq}$.

Question 6



$$X(g) + Y(g) \rightleftarrows XY(g)$$

The particle diagrams above show the changes that occurred after an equimolecular mixture of $X(g)$ and $Y(g)$ was placed inside a rigid container at constant temperature. Which of the following statements is be best supported by the particle diagrams?

(A) The rate of the reverse reaction is zero because the concentration of XY keeps increasing.
not necessarily true

(B) The reverse reaction has a higher rate than the forward reaction between 200 s and 300 s because [XY]>>[X].
not true → forward rxn is faster

(C) The forward reaction has a faster rate than the reverse reaction between 0 s and 300 s because more products were being formed. *True*

(D) The rates of the forward and reverse reactions were the same after 100 s because the rate of formation of XY was constant. *not true XY kept forming*

Question 7

$$2\,HI(g) \rightleftarrows H_2(g) + I_2(g)$$

Substance	Initial Concentration (M)	Equilibrium Concentration (M)
HI	0.30	0.24
H_2	0.00	0.030
I_2	0.00	?

In an experiment involving the reaction shown above, a sample of pure HI was placed inside a rigid container at a certain temperature. The table above provides the initial and equilibrium concentrations for some of the substances in the reaction. Based on the data, which of the following is the value of the equilibrium constant (K_{eq}) for the reaction, and why?

(A) $K_{eq} = 2.5 \times 10^{-1}$, because $[I_2]_{eq} = 2 \times [HI]_{eq}$.

(B) $K_{eq} = 6.3 \times 10^{-2}$, because $[I_2]_{eq} = \frac{1}{2}[HI]_{eq}$.

(C) $K_{eq} = 1.6 \times 10^{-2}$, because $[I_2]_{eq} = [H_2]_{eq}$.

(D) $K_{eq} = 3.1 \times 10^{-2}$, because $[I_2]_{eq} = 2 \times [H_2]_{eq}$.

$2HI(g) \rightleftarrows H_2(g) + I_2(g)$; I: .30, 0, 0; C: −2x, +x, +x; E: .24, .030, .030
$K = \frac{(.030)(.030)}{.24^2} = .016$

Question 8

$$SO_2(g) + Cl_2(g) \rightleftarrows SO_2Cl_2(g)$$

$SO_2(g)$ reacts with $Cl_2(g)$ to produce $SO_2Cl_2(g)$, according to the equation above. A mixture of 1.00 atm of $SO_2(g)$ and 1.00 atm of $Cl_2(g)$ is placed in a rigid, evacuated 1.00 L container, and the reaction mixture is allowed to reach equilibrium at a certain temperature. The table below shows the initial partial pressure of each gas and the equilibrium partial pressures of $SO_2(g)$ and $SO_2Cl_2(g)$ at the same temperature.

	P_{SO_2}	P_{Cl_2}	$P_{SO_2Cl_2}$
Initial pressure (atm)	1.00	1.00	0.00
Equilibrium pressure (atm)	0.20	?	0.80

Which of the following indicates the equilibrium value of P_{Cl_2} and the approximate value of the equilibrium constant K_p ?

(A) $P_{Cl_2} = 0.10$ atm and $K_p = 5.0$

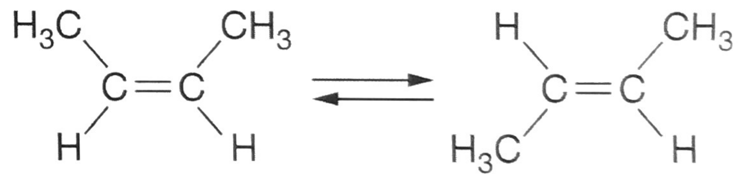
(B) $P_{Cl_2} = 0.20$ atm and $K_p = 0.050$

(C) $P_{Cl_2} = 0.20$ atm and $K_p = 20.$

(D) $P_{Cl_2} = 0.80$ atm and $K_p = 0.025$

$SO_2 + Cl_2 \rightleftarrows SO_2Cl_2$; I: 1.00, 1.00, 0; C: −x, −x, +x; E: .20, .20, .80
$K = \frac{.8}{(.20)(.20)}$

Question 9



Cis-2-Butene *Trans*-2-Butene

The diagram above represents the equilibrium between two isomers of 2-butene. The equilibrium constant, K_c, is 1.2 at a certain temperature. Two identical vessels each contain an equilibrium mixture of the two gases at that temperature. The concentration of *cis*-2-butene in the second vessel is twice the concentration in the first vessel. What is the concentration of *trans*-2-butene in the second vessel compared to that in the first vessel?

(A) Half the concentration of that in the first vessel

(B) The same concentration as that in the first vessel

(C) Twice the concentration of that in the first vessel

(D) Four times the concentration of that in the first vessel

>1 more product; both @ equil.; 1: cis, trans; 2: x2 cis, ?trans. For K to be same for both, x2 trans too

Question 10

$$HBrO(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + BrO^-(aq) \qquad K_{eq} = 2.8 \times 10^{-9}$$

<1 more reactants

The equilibrium reaction in $0.100\ M$ $HBrO(aq)$ at equilibrium is represented by the equation above. Based on the magnitude of the equilibrium constant, which of the following correctly compares the equilibrium concentrations of substances involved in the reaction, and why?

A. The equilibrium concentration of BrO^- will be much smaller than the equilibrium concentration of H_3O^+, because H_2O is the solvent and is present in the largest amount.

B. The equilibrium concentration of BrO^- will be much smaller than the equilibrium concentration of $HBrO$, because $K_{eq} << 1$.

C. The equilibrium concentration of H_3O^+ will be much smaller than the equilibrium concentration of BrO^-, because all the $HBrO$ will react to produce BrO^-.

D. The equilibrium concentration of H_3O^+ will be much larger than the equilibrium concentration of $HBrO$, because $K_{eq} << 1$.

Question 11

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$
$$K_p \approx 2 \times 10^5$$

At a certain temperature, $SO_2(g)$ and $O_2(g)$ react to produce $SO_3(g)$ according to the chemical equation shown above. An evacuated rigid vessel is originally filled with $SO_2(g)$ and $O_2(g)$, each with a partial pressure of 1 atm. Which of the following is closest to the partial pressure of $O_2(g)$ after the system has reached equilibrium, and why?

A. 0 atm; because K_p is very large, nearly all the $SO_2(g)$ and $O_2(g)$ are consumed before the system reaches equilibrium.

B. 0.5 atm; because K_p is very large, nearly all the $SO_2(g)$ is consumed before the system reaches equilibrium, but an excess amount of $O_2(g)$ remains at equilibrium.

C. 1 atm; because K_p is very large, the system is already near equilibrium, and there will be very little change to the partial pressure of $O_2(g)$.

D. 1.5 atm; because K_p is very large, the decomposition of any $SO_3(g)$ that forms increases the amount of $O_2(g)$ at equilibrium.

Question 12

.10 mol .50 mol

$$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Zn^{2+}(aq) + Cu(s) \qquad K_{eq} = 2 \times 10^{37}$$

Huge!

A 0.10 mol sample of solid zinc was added to 500.0 mL of $1.0\ M$ $Cu(NO_3)_2(aq)$. After the mixture sits overnight, which of the following best describes what was most likely observed and measured the next morning and why?

A. Almost all of the $Zn(s)$ will still be in the beaker with no visible $Cu(s)$, because equilibrium was not reached due to the very large K_{eq}.

B. About half of the $Zn(s)$ will have disappeared and $Cu(s)$ will have appeared in the beaker, because the system reached equilibrium.

C. About two-thirds of the $Zn(s)$ had disappeared and $Cu(s)$ will have appeared in the beaker, because the system reached equilibrium.

D. Virtually all of the $Zn(s)$ will have disappeared and $Cu(s)$ will have appeared in the beaker, because the reaction went almost to completion at equilibrium due to the very large K_{eq}.

Question 13

Reaction	Chemical Equation	K_{eq}
1	$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$	$K_1 = 0.010$
2	$H_2B(aq) \rightleftharpoons H^+(aq) + HB^-(aq)$	$K_2 = 0.000010$

reverse

The table above provides the K_{eq} values for two reactions. Which of the following is the correct mathematical expression needed to determine the equilibrium constant of the reaction shown below?

Reaction 3:

$$HA(aq) + HB^-(aq) \rightleftharpoons H_2B(aq) + A^-(aq)$$
$$K_{eq} = K_3$$

A. $K_3 = K_1 \times \frac{1}{K_2}$

B. $K_3 = \frac{1}{K_1} \times K_2$

C. $K_3 = K_1 - K_2$

D. $K_3 = -K_1 + K_2$

Question 14

Reaction 1: $PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,Cl^-(aq)$ $K_{sp} = K_1$ good

Reaction 2: $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ $K_{sp} = K_2$ reverse and x2

Based on the information given above, which of the following is the expression for K_{eq} for the reaction that occurs when a $0.1\ M\ AgNO_3(aq)$ is added to a saturated solution of $PbCl_2(aq)$, as represented by the following chemical equation?

$$PbCl_2(s) + 2\,Ag^+(aq) \rightleftharpoons Pb^{2+}(aq) + 2\,AgCl(s)$$

A $K_{eq} = K_1 + (2 \times K_2)$

B $K_{eq} = K_1 - (2 \times K_2)$

C $K_{eq} = K_1 \times (K_2)^2$

D $K_{eq} = \frac{K_1}{(K_2)^2}$

Question 15

Equation 1: $2\,NO_2(g) \rightleftharpoons N_2O_4(g)$ K_1

Equation 2: $N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$ K_2

In a large reaction vessel at a constant temperature, nitrogen dioxide (NO_2) and dinitrogen tetroxide (N_2O_4) are in a state of dynamic equilibrium, as represented by the chemical equations shown above. The equilibrium constants for the reactions are K_1 and K_2. Which of the following quantities can most easily be used to find the value of K_2?

A The value of ΔH for the reaction

B The temperature of the system

C The volume of the system

D The value of K_1

If have K_1, K_2 is simply $\frac{1}{K_1}$ since they are reverse of each other

Question 16

Hey → we did this in lab!

heat + $Co(H_2O)_6^{2+}(aq) + 4\,Cl^-(aq) \rightleftharpoons CoCl_4^{2-}(aq) + 6\,H_2O(l)$ $\Delta H > 0$ (+)

pink blue

A student poured 10 mL of $CoCl_2(aq)$ into a test tube and added a few drops of concentrated HCl, which resulted in a deep-blue solution. The reaction that occurred is represented by the chemical equation shown above. Then, the student placed the test tube inside a beaker that contained ice and water for about five minutes. Which of the following describes what the student most likely observed next, and why?

A The color of the solution changed from blue to pink, because lowering the temperature increased the collision frequency between $CoCl_4^{2-}$ and H_2O.

B The color of the solution changed from blue to pink, because cooling caused the equilibrium to shift to form the pink-colored $Co(H_2O)_6^{2+}$.

C The color of the solution did not change, because more water was not added to the solution.

D The color of the solution did not change, because heat is not released from this reaction.

Question 17

more

$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$

colorless colorless red-brown

The formation of $FeSCN^{2+}$ in an aqueous solution is represented by the chemical equilibrium shown above. A light red-brown solution is prepared combining $12.50\ mL$ of $0.5\ M\ Fe(NO_3)_3$, $0.5\ mL$ of $0.002\ M\ KSCN$, and $37.0\ mL$ of water that had been slightly acidified. If an additional $1.0\ mL$ of $0.5\ M\ Fe(NO_3)_3$ is added, which of the following predicts and explains correctly whether the darkness of the color of the solution will change?

A The color of the solution will lighten because most of the volume added is solvent.

B The color of the solution will darken because the equilibrium will favor the dissociation of $FeSCN^{2+}$.

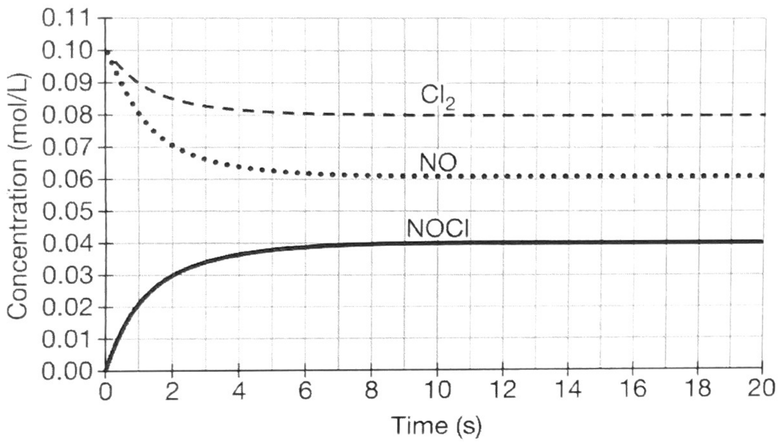
C The color of the solution will darken because the equilibrium will favor the formation of more $FeSCN^{2+}$.

D The color of the solution will not change because the solution already contains Fe^{3+} ions.

Question 18

$$2\,NO(g) + Cl_2(g) \rightleftharpoons 2\,NOCl(g)$$

The reaction of $NO(g)$ with $Cl_2(g)$ is represented by the balanced equation above.



A chemist carried out the reaction at $573\ K$, starting with $0.100\ mol$ of each reactant in a $1.00\ L$ container with variable volume. The reaction mixture quickly reached equilibrium, as indicated in the graph above. The chemist reduces the volume of the equilibrium system by half while keeping the temperature constant. Which of the following predictions about the yield of the reaction is best, and why?

A. By halving the volume, the pressure doubles. The system will respond to the increase in pressure by decreasing the total number of moles of gas in the system. Thus, the yield will increase because the reaction will shift toward more product.

B. By halving the volume, the pressure doubles. The system will respond to the increase in pressure by increasing the total number of moles of gas in the system. Thus, the yield will decrease because the reaction will shift toward more reactants.

C. By halving the volume, the pressure is reduced by half. The system will respond to the decrease in pressure by increasing the total number of moles of gas in the system. Thus, the yield will decrease because the reaction will shift toward more reactants.

D. By halving the volume, the pressure is reduced by half. The system will respond to the decrease in pressure by decreasing the total number of moles of gas in the system. Thus, the yield will increase because the reaction will shift toward more products.

Question 19

$$CO(g) + 2\,H_2(g) \rightleftharpoons CH_3OH(g) \qquad K_c = 11$$

Substance	Initial Concentration (M)
CO	0.10
H_2	0.10
CH_3OH	0.10

The table lists the initial concentrations of each substance in the system represented by the equation above at a given temperature. Which of the following best predicts what will occur as the system approaches equilibrium?

A. The rate of the reverse reaction will be less than the rate of the forward reaction and additional $CH_3OH(g)$ will be consumed, because $K_c < Q_c$.

B. The rate of the reverse reaction will be greater than the rate of the forward reaction and additional $CH_3OH(g)$ will be consumed because $K_c < Q_c$.

C. The rate of the forward reaction will be less than the rate of the reverse reaction and additional $CH_3OH(g)$ will be produced, because $K_c < Q_c$.

D. The rate of the forward reaction will be greater than the rate of the reverse reaction and additional $CH_3OH(g)$ will be produced, because $K_c < Q_c$.

Question 20

Substance	Concentration (M)
X_2	0.10
Y_2	0.30
XY	0.20

$$X_2(g) + Y_2(g) \rightleftharpoons 2\,XY(g) \qquad K_c = 25$$

The chemical reaction shown above took place inside a rigid container at constant temperature. The table provides the concentrations of reactants and products at some point during the reaction. Based on this information, which of the following explains whether or not the reaction has reached equilibrium, and why?

A. The reaction is not at equilibrium because $Q_c > K_c$; the forward reaction is favored in order to form more XY.

B. The reaction is not at equilibrium because $Q_c < K_c$; the reverse reaction is favored in order to form more X_2 and Y_2.

C. The reaction is not at equilibrium because $Q_c < K_c$; the forward reaction is favored in order to form more XY.

D. The reaction is at equilibrium because $Q_c = K_c$ and more XY will not be formed.

Question 21

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

At a given temperature, the system represented by the chemical equation above is at equilibrium inside a rigid container. Which of the following explains how the system will restore equilibrium, based on the correct relationship between Q and K, after a certain amount of $Cl_2(g)$ is added?

A. Since $Q = \frac{[Cl_2][CO]}{[COCl_2]}$, when $Cl_2(g)$ is added to the system $Q > K$ and the system will restore equilibrium by producing more $COCl_2(g)$.

B. Since $Q = \frac{[Cl_2][CO]}{[COCl_2]}$, when $Cl_2(g)$ is added to the system $Q < K$ and the system will restore equilibrium by producing more $CO(g)$.

C. Since $Q = \frac{[COCl_2]}{[CO][Cl_2]}$, when $Cl_2(g)$ is added to the system $Q > K$ and the system will restore equilibrium by producing more $COCl_2(g)$.

D. Since $Q = \frac{[COCl_2]}{[CO][Cl_2]}$, when $Cl_2(g)$ is added to the system $Q > K$ and the system will restore equilibrium by producing more $CO(g)$.

Question 22

For which of the following salts would the relationship between molar solubility, s, in mol/L, and the value of K_{sp} be represented by the equation $K_{sp} = 4s^3$?

A. $PbCO_3$ x^2

B. $Mg_3(PO_4)_2$ $(3x)^3(2x)^2 = 36x^5$

C. Ag_2SO_4 $4x^3$

D. MnS x^2

$A \rightarrow 2B + C$

$+2x$ $+x$

$(2x)^2\ x$

$4x^3$

Question 23

$$Hg_2I_2(s) \rightleftarrows Hg_2^{2+}(aq) + 2I^-(aq) \qquad K_{sp} = [Hg_2^{2+}][I^-]^2$$

A saturated solution of Hg_2I_2 is at equilibrium at $25^\circ C$ as represented by the equation above. If $[I^-] = 4.6 \times 10^{-10}\ M$ at equilibrium, which of the following gives the correct molar solubility, S, and K_{sp} for Hg_2I_2 ?

A. $S = 4.6 \times 10^{-10}\ M;\ K_{sp} = (2.3 \times 10^{-10})(4.6 \times 10^{-10})^2$

B. $S = 4.6 \times 10^{-10}\ M;\ K_{sp} = (4.6 \times 10^{-10})(9.2 \times 10^{-10})^2$

C. $S = 2.3 \times 10^{-10}\ M;\ K_{sp} = (2.3 \times 10^{-10})(4.6 \times 10^{-10})^2$

D. $S = 2.3 \times 10^{-10}\ M;\ K_{sp} = (4.6 \times 10^{-10})(9.2 \times 10^{-10})^2$

$Hg_2I_2(s) \rightleftarrows Hg_2^{+2} + 2I^-$

$+x$ $+2x$

x $2x$

$2x = 4.6 \times 10^{-10}\ M$

Question 24

$$AgI(s) \rightleftarrows Ag^+(aq) + I^-(aq) \qquad K_{sp} = 8.3 \times 10^{-17} \text{ at } 298\ K$$

The dissolution of AgI is represented above. Which of the following shows the mathematical relationship between the molar solubility, S, of AgI and the K_{sp} at $298\ K$?

A. $S = 8.3 \times 10^{-17}\ mol/L$

B. $S = \left(\frac{8.3 \times 10^{-17}}{2}\right) mol/L$

C. $S = \sqrt{8.3 \times 10^{-17}}\ mol/L$

D. $S = 2\sqrt{8.3 \times 10^{-17}}\ mol/L$

$K_{sp} = S^2$

$S = \sqrt{K_{sp}}$

Question 25

$$Ca(OH)_2(s) \rightleftarrows Ca^{2+}(aq) + 2\,OH^-(aq)$$
$$K_{sp} = 5.5 \times 10^{-6} \text{ at } 298\text{ K}$$

The equilibrium in a saturated solution of $Ca(OH)_2$ is represented above. In an experiment, a student places 5.0 g of $Ca(OH)_2(s)$ into 100.0 mL of distilled water and stirs the mixture. How would the results be affected if the student repeats the experiment but this time places 5.0 g of $Ca(OH)_2(s)$ into 100.0 mL of 0.0010 M $NaOH(aq)$ instead of distilled water?

common ion OH^-

A. Less solid will dissolve, because the larger value of $[OH^-]$ will cause the equilibrium position to lie farther to the right.

B. Less solid will dissolve, because the larger value of $[OH^-]$ will cause the equilibrium position to lie farther to the left.

C. More solid will dissolve, because the larger value of $[OH^-]$ will cause the equilibrium position to lie farther to the right.

D. More solid will dissolve, because the smaller value of $[OH^-]$ will cause the equilibrium position to lie farther to the left.

Question 26

← equilibrium shifts left

$$AgCN(s) \rightleftarrows Ag^+(aq) + CN^-(aq)$$

The dissolution of solid $AgCN$ is represented by the chemical equation above. In pure water, the equilibrium concentration of Ag^+ ions in a saturated solution is 7.7×10^{-9} M. If a small amount of solid $NaCN$ is added to the saturated $AgCN$ solution, which of the following would be observed?

add

A. The K_{sp} increases and more $AgCN$ dissolves.

B. The K_{sp} increases and some $AgCN$ precipitates.

C. The molar solubility of $AgCN$ becomes smaller than 7.7×10^{-9} M and some $AgCN$ precipitates.

D. The molar solubility of $AgCN$ becomes larger than 7.7×10^{-9} M and more $AgCN$ dissolves.

Question 27

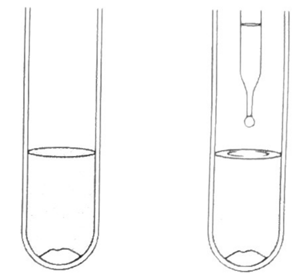


Figure 1
pH = 10.35

Figure 2

$$Mg(OH)_2(s) \rightleftarrows Mg^{2+}(aq) + 2\,OH^-(aq)$$

A student prepared a saturated aqueous solution of $Mg(OH)_2$ and measured its pH, as shown in Figure 1 above. Then the student added a few drops of an unknown solution to the test tube and observed cloudiness in the solutions as shown in Figure 2. On the basis of this information and the equilibrium represented above, which of the following is most likely the identity of the reagent added from the dropper?

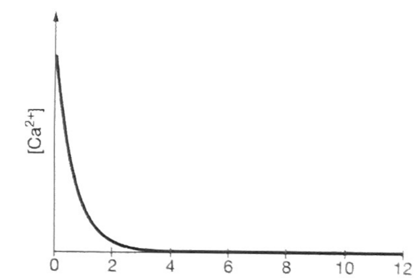
A. Distilled water → dilute, but not part of rxn

B. $NaNO_3(aq)$ irrelevant

C. $HCl(aq)$ → more of basic anion can dissolve shift →

D. $KOH(aq)$ ↑ common ion OH^- shift ←

Question 28



$CaF_2(s)$ dissolves in water according to the equation $CaF_2(s) \rightleftarrows Ca^{2+}(aq) + 2\,F^-(aq)$. The value of K_{sp} for the dissolution is 3.5×10^{-11}. A student measures the concentration of Ca^{2+} ions in a saturated solution of CaF_2 at various pH values and uses those values to generate the graph above. Based on the data, which of the following observations about the solubility of CaF_2 is most valid?

A. It does not depend on pH because $[Ca^{2+}]$ does not change between pH 4 and pH 12.

B. It does not depend on pH because $K_{sp} = [Ca^{2+}][F^-]^2$, so as $[Ca^{2+}]$ decreases, $[F^-]$ increases to compensate, keeping K_{sp} constant.

C. It is higher at a lower pH; there are more H^+ ions in solution at low pH, so HF forms and shifts the equilibrium reaction above to the right.

D. It is lower at a higher pH; there are more H^+ ions in solution at high pH, so HF forms and shifts the equilibrium reaction above to the right.

Question 29

pH of saturated $Cu(OH)_2$ solution	Concentration of H_3O^+ ions	Absorbance
7.50	$3.2 \times 10^{-8}\ M$	0.047
7.25	$5.6 \times 10^{-8}\ M$	0.149
7.00	$1.0 \times 10^{-7}\ M$	0.470

more H^+ — more soluble

$Cu^{2+}(aq)$ absorbs a certain frequency of visible light. Absorbance was measured for three saturated solutions of $Cu(OH)_2$, each at a different pH at 298 K. Based on the data recorded in the table above, which of the following conclusions about the effect of pH on the solubility of $Cu(OH)_2$ can be made?

A When the concentration of H^+ ions is increased, the solubility of $Cu(OH)_2$ increases.

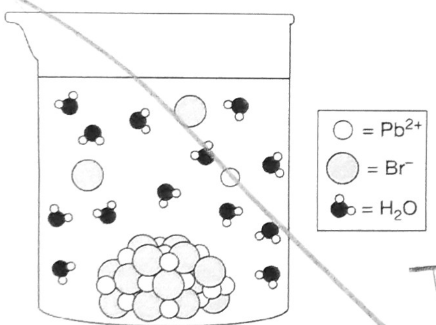
B When the concentration of H^+ ions is decreased, the solubility of $Cu(OH)_2$ increases.

C The solubility of $Cu(OH)_2$ is independent of pH.

D $Cu(OH)_2$ is soluble only at a pH of 7.00.

Question 30

$$PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,Br^-(aq)$$



$\Delta H^\circ > 0$	enthalpy (heat)
$\Delta S^\circ > 0$	entropy (randomness)
$\Delta G^\circ > 0$	free energy (is a process favorable or not)

This Q should be in Unit 9

Shown above are a chemical equation that represents the dissolution of $PbBr_2$ in pure water, a table of the changes in some thermodynamic properties for the process, and a particle diagram. Which of the following explains which relative change in a thermodynamic property is best illustrated by the particle diagram?

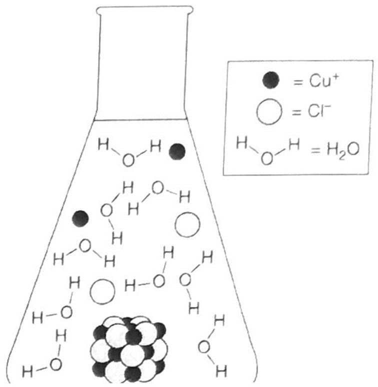
A The reorganization of the water molecules around the ions illustrates that $\Delta H^\circ > 0$ because forming strong ion-dipole interactions releases energy.

B The very small amount of Pb^{2+} and Br^- ions illustrates that $\Delta S^\circ > 0$ because entropy decreases when $PbBr_2$ dissolves.

C The very small amount of Pb^{2+} and Br^- ions illustrates that $\Delta G^\circ > 0$ because the dissolution of $PbBr_2$ is not a favorable process.

D The very large amount of solid that remains undissolved illustrates that $\Delta G^\circ > 0$ because the dissolution of $PbBr_2$ is a favorable process.

Question 31



The particle diagram shown above represents the dissolution of $CuCl(s)$ assuming an equilibrium concentration for Cu^+ ions of about $4 \times 10^{-4} M$ in a saturated solution at $25^\circ C$. The equilibrium being represented is shown in the following chemical equation.

$$CuCl(s) \rightleftharpoons Cu^+(aq) + Cl^-(aq)$$

add

Which of the following changes to the particle diagram will best represent the effect of adding 1.0 mL of 4 M NaCl to the solution?

A Some of the Cu^+ and Cl^- ions combine to form $CuCl(s)$ because the K_{sp} will be lower than 1.6×10^{-7}.

B Some of the Cu^+ and Cl^- ions combine to form $CuCl(s)$ because the molar solubility will be lower than $4 \times 10^{-4} M$.

C More Cu^+ and Cl^- ions will be in solution because the molar solubility will be higher than $4 \times 10^{-4} M$.

D More Cu^+ and Cl^- ions will be in solution because the K_{sp} will be higher than 1.6×10^{-7}.

The particle diagram shown above represents the dissolution of $CuCl(s)$ assuming an equilibrium concentration for Cu^+ ions of about $4 \times 10^{-4} M$ in a saturated solution at $25^\circ C$. The equilibrium being represented is shown in the following chemical equation.

$$CuCl(s) \rightleftharpoons Cu^+(aq) + Cl^-(aq)$$

Unit 9

Which of the following best explains what the particle diagram is able to show about the entropy change for the dissolution of $CuCl(s)$ in pure water?

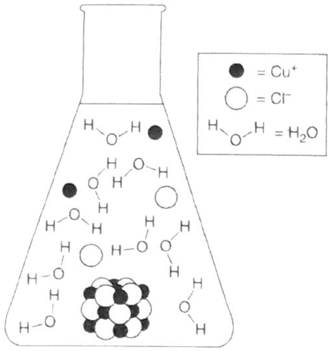
A The particle diagram shows that the dissociation of $CuCl(s)$ into ions contributes to an increase in the entropy for the dissolution.

B The particle diagram shows that the dissociation of $CuCl(s)$ into ions contributes to a decrease in the entropy for the dissolution.

C The particle diagram shows that there is no reorganization of the water molecules around the ions and the change in entropy for the dissolution is zero.

D The particle diagram shows that there are no interactions between the water molecules and the change in entropy for the dissolution is zero.

Question 33



The particle diagram shown above represents the dissolution of $CuCl(s)$ assuming an equilibrium concentration for Cu^+ ions of about $4 \times 10^{-4} M$ in a saturated solution at 25°C. The equilibrium being represented is shown in the following chemical equation.

$CuCl(s) \rightleftharpoons Cu^+(aq) + Cl^-(aq)$

Which of the following best explains whether or not the particle diagram can predict the relative value of the enthalpy change for the dissolution of $CuCl(s)$?

A. The value of the enthalpy change for the dissolution of $CuCl(s)$ cannot be predicted from the particle diagram because it fails to illustrate the amount of energy required to overcome the forces between solute particles and between solvent particles. + or − need to know!

B. The value of the enthalpy change for the dissolution of $CuCl(s)$ cannot be predicted from the particle diagram because it fails to illustrate the amount of energy released when the water molecules form hydrogen bonds with Cl^- ions.

C. The value of the enthalpy change for the dissolution of $CuCl(s)$ is positive (endothermic) because energy is released to overcome the forces between solute particles, as shown in the particle diagram.

D. The value of the enthalpy change for the dissolution of $CuCl(s)$ is negative (exothermic) because energy is required when the bonds between the ions are broken, as shown in the particle diagram.

Question 1

For parts of the free-response question that require calculations, clearly show the method used and the steps involved in arriving at your answers. You must show your work to receive credit for your answer. Examples and equations may be included in your answers where appropriate.

$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq)$ $K_{sp} = 7 \times 10^{-9}$

$7\times10^{-9} = 4x^3$

The dissolution of $PbI_2(s)$ is represented above.

(a) Write a mathematical expression that can be used to determine the value of S, the molar solubility of $PbI_2(s)$. (Do not do any numerical calculations.)

(b) If $PbI_2(s)$ is dissolved in 1.0 M $NaI(aq)$, is the maximum possible concentration of $Pb^{2+}(aq)$ in the solution greater than, less than, or equal to the concentration of $Pb^{2+}(aq)$ in the solution in part (a) ? Explain.

Common ion → less dissolves due to I^- already in solution less Pb^{+2} before Ksp is reached

Compound	K_{sp}
$PbCl_2$	2×10^{-5}
PbI_2	7×10^{-9}
$Pb(IO_3)_2$	3×10^{-13}

(c) A table showing K_{sp} values for several lead compounds is given above. A saturated solution of which of the compounds has the greatest molar concentration of $Pb^{2+}(aq)$? Explain. (Do not do any numerical calculations.)

$PbCl_2$ has highest b/c largest K_{sp}

$PbI_2(s) \rightleftharpoons Pb^{+2} + 2I^-$

I	/	0	0
C	/	+x	+2x
E	/	x	2x

$(x)(2x)^2$

Question 2

For parts of the free-response question that require calculations, clearly show the method used and the steps involved in arriving at your answers. You must show your work to receive credit for your answer. Examples and equations may be included in your answers where appropriate.

Gas	Initial Concentration (M)
H_2	0.030
I_2	0.015
HI	?

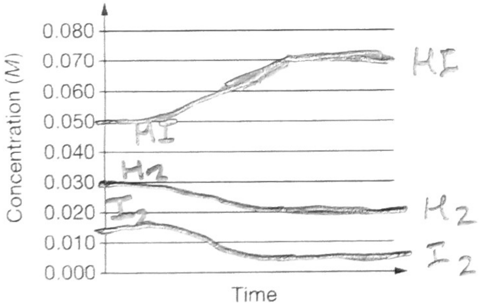
_ H_2 + _ I_2 ⇌ 2 HI
I .03 .015
C −x −x +2x
E

Samples of three gases, $H_2(g)$, $I_2(g)$, and $HI(g)$, were combined in a rigid vessel. The initial concentrations of $H_2(g)$ and $I_2(g)$ are given in the table above.

(a) The original value of the reaction quotient, Q_c, for the reaction of $H_2(g)$ and $I_2(g)$ to form $HI(g)$ (before any reactions take place and before equilibrium is established), was 5.56. On the following graph, plot the points representing the initial concentrations of all three gases. Label each point with the formula of the gas.

$Q = \frac{[HI]^2}{[H_2][I_2]}$ $5.56 = \frac{[HI]^2}{(.03)(.015)}$

[HI] = .050



Equilibrium was established at a certain temperature according to the following chemical equation.

$$H_2(g) + I_2(g) \rightleftarrows 2\,HI(g) \qquad \Delta H^\circ_{rxn} = -9.4\ \text{kJ/mol}_{rxn}; \quad K_c = 49$$

After equilibrium was established, the concentration of $H_2(g)$ was 0.020 M.

(b) On the graph above, carefully draw three curves, one for each of the three gases, starting from the initial points you drew in part (a). The curves must show how the concentration of each of the three gases changed as equilibrium was established.

H_2 + I_2 ⇌ 2 HI + heat
I .030 .015 .050
C −x −x +2x
E .020 .005 .070

so...

x = .010

$H_2(g)$, $I_2(g)$, and $HI(g)$ are at equilibrium at a different temperature in a different vessel.

(c) When the temperature in the vessel is decreased, does the equilibrium shift to the right, favoring the product, or to the left, favoring the reactants? Justify your answer.

→ + heat

b/c forward reaction is exothermic

shift right

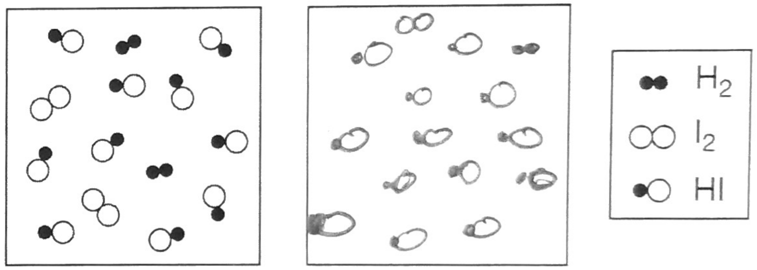
(d) Does the value of K_c increase, decrease, or remain the same when the temperature is decreased? Justify your answer based on the expression for K_c and the concentrations of the product and reactants.

$K_c = \frac{[HI]^2}{[H_2][I_2]}$

shift right means more products on top

K_c increases

(e) In the following empty box, draw an appropriate number of each type of molecule to represent a possible new equilibrium at the lower temperature.



10 HI
2 H_2
2 I_2

lower temp shifts reaction forward so more HI in box

12 HI
1 H_2
1 I_2