

Equilibrium

2003B.1 – Expression, Graph, ICE/K, QvsK
2004B.1 – Expression, QvsK, ManK
2004B.8 – K (no math), QvsK, LeChat
2007.6 – VSEPR, hybrid, Catalyst
2008.1 – Expression, IGL, Dalton, K, Catalyst, QvsK
2008B.1 – Molarity, IGL, Expression, K, Mole Fraction
2010B.1 – IMF, Expression, IGL, K LeChat
2013C.3 – IGL, Dalton, Expression, K, LeChat, QvsK
2014.4 – IGL, LeChat
2015B.5 – VSEPR, IGL, Mole Fraction, K
2016.6 – KInt, QvsK
2018B.2 – ox #, Stoich, LeChat, ICE/K, KInt, IMF, IGL, Lab

K_{sp}

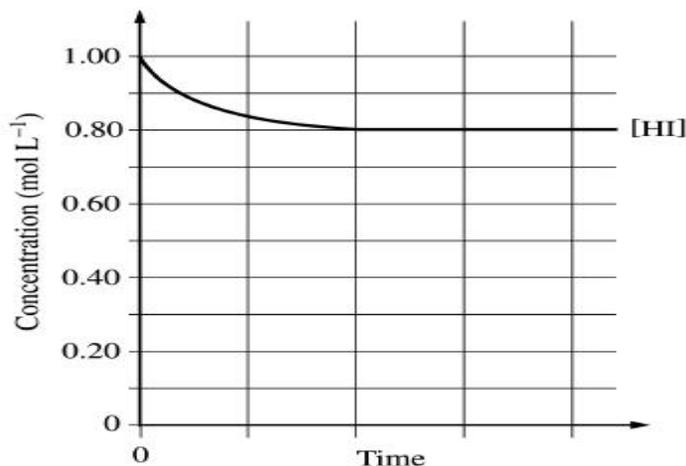
2001.1 – Rxn, ICE/K, QvsK, SP
2004.1 – Expression, ICE/K, Molarity, LeChat, Rxn
2006.1 – Expression, ICE/K, LeChat, Molarity QvsKP
2006B.5 – Beers, Expression, ICE, KInt
2010.1 – Expression, ICE/K, LeChat, Molarity, QvsKP, Rxn, KInt
2013A.1 – Expression, Molarity (Stoich), ICE/K, SP, MVMV, Lab
2013C.1 – Rxn, Electrolytes, Titration, Molarity, ICE/K, LeChat
2014.1 – Rxn, Lab, Stoich, KInt
2015.4 – Rxn, ICE/K, Particle Diagram
2015B.2 – Trend (Size), KInt, Rxn, Lab (Procedures), % by mass
2017.6 – ICE/K, Lattice Energy
2017B.6 – ICE/K, LeChat

KEY

ManK - Manipulation of K
QvsK – Q vs. K
KInt – Interpretation of K
LeChat – Le Chatelier and/or Common Ion Effect
ICE/K – ICE Table or Calculation involving K and/or Molar Solubility
QvsKP – Q vs. K (Precipitation)
SP – Selective Precipitation



After a 1.0 mole sample of $\text{HI}(g)$ is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of $\text{HI}(g)$ as a function of time is shown below.



- Write the expression for the equilibrium constant, K_c , for the reaction.
- What is $[\text{HI}]$ at equilibrium?
- Determine the equilibrium concentrations of $\text{H}_2(g)$ and $\text{I}_2(g)$.
- On the graph above, make a sketch that shows how the concentration of $\text{H}_2(g)$ changes as a function of time.
- Calculate the value of the following equilibrium constants at 700. K.
 - K_c
 - REMOVED**
- At 1,000 K, the value of K_c for the reaction is 2.6×10^{-2} . In an experiment, 0.75 mole of $\text{HI}(g)$, 0.10 mole of $\text{H}_2(g)$, and 0.50 mole of $\text{I}_2(g)$ are placed in a 1.0 L container and allowed to reach equilibrium at 1,000 K. Determine whether the equilibrium concentration of $\text{HI}(g)$ will be greater than, equal to, or less than the initial concentration of $\text{HI}(g)$. Justify your answer.



For the reaction represented above, the value of the equilibrium constant, K_p , is 3.1×10^{-4} at 700. K.

(a) Write the expression for the equilibrium constant, K_p , for the reaction.

(b) Assume that the initial partial pressures of the gases are as follows:

$$p_{\text{N}_2} = 0.411 \text{ atm}, \quad p_{\text{H}_2} = 0.903 \text{ atm}, \quad \text{and} \quad p_{\text{NH}_3} = 0.224 \text{ atm}.$$

(i) Calculate the value of the reaction quotient, Q , at these initial conditions.

(ii) Predict the direction in which the reaction will proceed at 700. K if the initial partial pressures are those given above. Justify your answer.

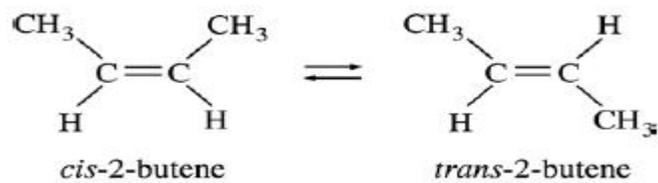
(c) **REMOVED**

(d) The value of K_p for the reaction represented below is 8.3×10^{-3} at 700. K.



Calculate the value of K_p at 700. K for each of the reactions represented below.





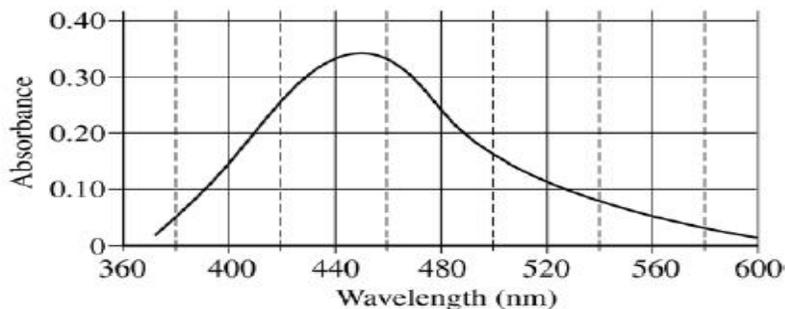
The gas-phase conversion reaction between the geometric isomers *cis*-2-butene and *trans*-2-butene is represented by the equation above. The value of the equilibrium constant, K_{eq} , for the reaction is 3.2 at 298 K and 1.0 atm.

- In a mixture of the isomers at equilibrium at 298 K and 1.0 atm, which is present at a higher concentration, *cis*-2-butene or *trans*-2-butene? Justify your answer.
- If 1.00 mol of pure *cis*-2-butene and 1.0 mol of pure *trans*-2-butene were introduced into an evacuated container at 298 K, in which direction (to the right or to the left) would the reaction proceed to establish equilibrium? Justify your answer.
- Given that K_{eq} for the reaction at 400 K has the value 1.3, predict whether the reaction is endothermic or exothermic. Justify your answer.
- REMOVED**

2006B.5

A student carries out an experiment to determine the equilibrium constant for a reaction by colorimetric (spectrophotometric) analysis. The production of the red-colored species $\text{FeSCN}^{2+}(aq)$ is monitored.

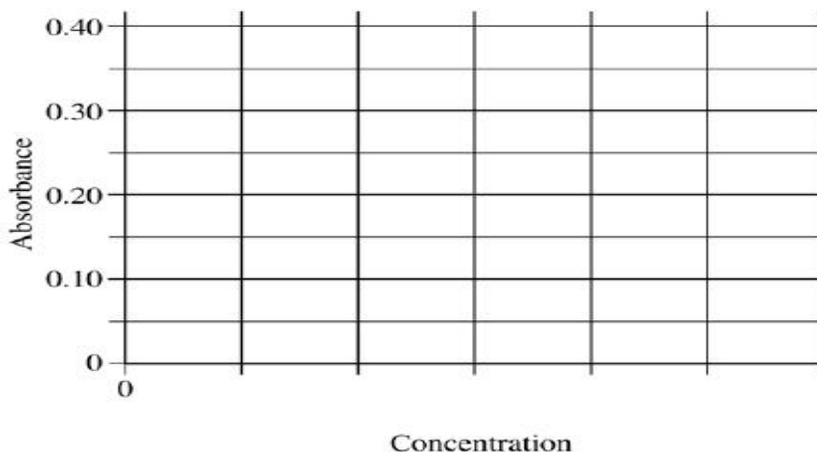
- (a) The optimum wavelength for the measurement of $[\text{FeSCN}^{2+}]$ must first be determined. The plot of absorbance, A , versus wavelength, λ , for $\text{FeSCN}^{2+}(aq)$ is given below. What is the optimum wavelength for this experiment? Justify your answer.



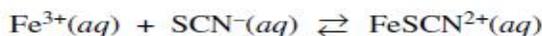
- (b) A calibration plot for the concentration of $\text{FeSCN}^{2+}(aq)$ is prepared at the optimum wavelength. The data below give the absorbances measured for a set of solutions of known concentration of $\text{FeSCN}^{2+}(aq)$.

Concentration (mol L ⁻¹)	Absorbance
1.1×10^{-4}	0.030
3.0×10^{-4}	0.065
8.0×10^{-4}	0.160
12×10^{-4}	0.239
18×10^{-4}	0.340

- (i) Draw a Beer's law calibration plot of all the data on the grid below. Indicate the scale on the horizontal axis by labeling it with appropriate values.



- (ii) An $\text{FeSCN}^{2+}(aq)$ solution of unknown concentration has an absorbance of 0.300. Use the plot you drew in part (i) to determine the concentration, in moles per liter, of this solution.
- (c) The purpose of the experiment is to determine the equilibrium constant for the reaction represented below.



- (i) Write the equilibrium-constant expression for K_c .
- (ii) The student combines solutions of $\text{Fe}(\text{NO}_3)_3$ and KSCN to produce a solution in which the initial concentrations of $\text{Fe}^{3+}(aq)$ and $\text{SCN}^{-}(aq)$ are both $6.0 \times 10^{-3} M$. The absorbance of this solution is measured, and the equilibrium $\text{FeSCN}^{2+}(aq)$ concentration is found to be $1.0 \times 10^{-3} M$. Determine the value of K_c .
- (d) If the student's equilibrium $\text{FeSCN}^{2+}(aq)$ solution of unknown concentration fades to a lighter color before the student measures its absorbance, will the calculated value of K_c be too high, too low, or unaffected? Justify your answer.

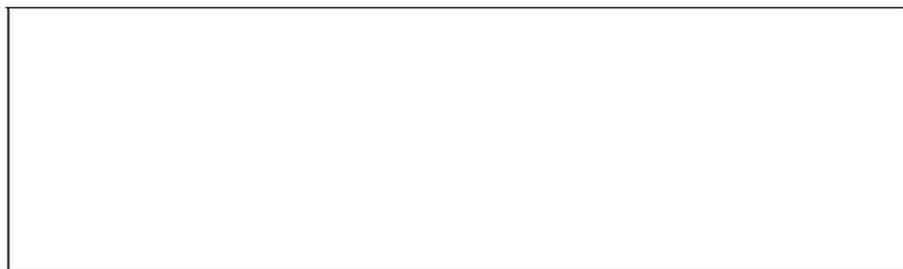
2007.6

Answer the following questions, which pertain to binary compounds.

- (a) In the box provided below, draw a complete Lewis electron-dot diagram for the IF_3 molecule.



- (b) On the basis of the Lewis electron-dot diagram that you drew in part (a), predict the molecular geometry of the IF_3 molecule.
- (c) In the SO_2 molecule, both of the bonds between sulfur and oxygen have the same length. Explain this observation, supporting your explanation by drawing in the box below a Lewis electron-dot diagram (or diagrams) for the SO_2 molecule.



- (d) On the basis of your Lewis electron-dot diagram(s) in part (c), identify the hybridization of the sulfur atom in the SO_2 molecule.

The reaction between $\text{SO}_2(\text{g})$ and $\text{O}_2(\text{g})$ to form $\text{SO}_3(\text{g})$ is represented below.



The reaction is exothermic. The reaction is slow at 25°C ; however, a catalyst will cause the reaction to proceed faster.

- (e) Using the axes provided on the next page, draw the complete potential-energy diagram for both the catalyzed and uncatalyzed reactions. Clearly label the curve that represents the catalyzed reaction.



- (f) Predict how the ratio of the equilibrium pressures, $\frac{P_{\text{SO}_2}}{P_{\text{SO}_3}}$, would change when the temperature of the uncatalyzed reaction mixture is increased. Justify your prediction.
- (g) How would the presence of a catalyst affect the change in the ratio described in part (f)? Explain.



Solid carbon and carbon dioxide gas at 1,160 K were placed in a rigid 2.00 L container, and the reaction represented above occurred. As the reaction proceeded, the total pressure in the container was monitored. When equilibrium was reached, there was still some $\text{C}(s)$ remaining in the container. Results are recorded in the table below.

Time (hours)	Total Pressure of Gases in Container at 1,160 K (atm)
0.0	5.00
2.0	6.26
4.0	7.09
6.0	7.75
8.0	8.37
10.0	8.37

- (a) Write the expression for the equilibrium constant, K_p , for the reaction.
- (b) Calculate the number of moles of $\text{CO}_2(g)$ initially placed in the container. (Assume that the volume of the solid carbon is negligible.)
- (c) For the reaction mixture at equilibrium at 1,160 K, the partial pressure of the $\text{CO}_2(g)$ is 1.63 atm. Calculate
- the partial pressure of $\text{CO}(g)$, and
 - the value of the equilibrium constant, K_p .
- (d) If a suitable solid catalyst were placed in the reaction vessel, would the final total pressure of the gases at equilibrium be greater than, less than, or equal to the final total pressure of the gases at equilibrium without the catalyst? Justify your answer. (Assume that the volume of the solid catalyst is negligible.)

In another experiment involving the same reaction, a rigid 2.00 L container initially contains 10.0 g of $\text{C}(s)$, plus $\text{CO}(g)$ and $\text{CO}_2(g)$, each at a partial pressure of 2.00 atm at 1,160 K.

- (e) Predict whether the partial pressure of $\text{CO}_2(g)$ will increase, decrease, or remain the same as this system approaches equilibrium. Justify your prediction with a calculation.

2008B.1

Answer the following questions regarding the decomposition of arsenic pentafluoride, $\text{AsF}_5(g)$.

(a) A 55.8 g sample of $\text{AsF}_5(g)$ is introduced into an evacuated 10.5 L container at 105°C .

(i) What is the initial molar concentration of $\text{AsF}_5(g)$ in the container?

At 105°C , $\text{AsF}_5(g)$ decomposes into $\text{AsF}_3(g)$ and $\text{F}_2(g)$ according to the following chemical equation.



(b) In terms of molar concentrations, write the equilibrium-constant expression for the decomposition of $\text{AsF}_5(g)$.

(c) When equilibrium is established, 27.7 percent of the original number of moles of $\text{AsF}_5(g)$ has decomposed.

(i) Calculate the molar concentration of $\text{AsF}_5(g)$ at equilibrium.

(ii) Using molar concentrations, calculate the value of the equilibrium constant, K_{eq} , at 105°C .

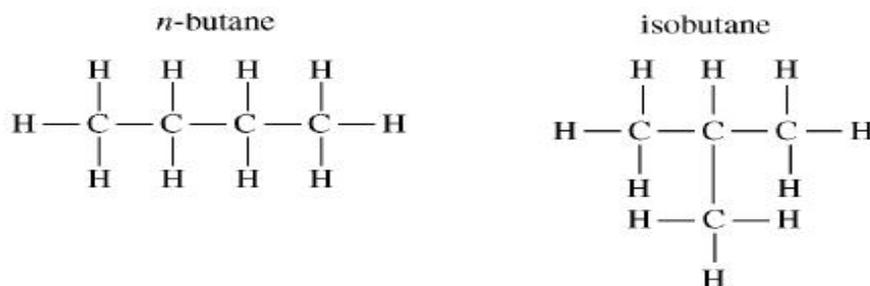
(d) Calculate the mole fraction of $\text{F}_2(g)$ in the container at equilibrium.

2010B.1

The compound butane, C_4H_{10} , occurs in two isomeric forms, *n*-butane and isobutane (2-methyl propane). Both compounds exist as gases at $25^\circ C$ and 1.0 atm.

(a) **REMOVED**

(b) On the basis of molecular structures, identify the isomer that has the higher boiling point. Justify your answer.



The two isomers exist in equilibrium as represented by the equation below.



Suppose that a 0.010 mol sample of pure *n*-butane is placed in an evacuated 1.0 L rigid container at $25^\circ C$.

- (c) Write the expression for the equilibrium constant, K_c , for the reaction.
- (d) Calculate the initial pressure in the container when the *n*-butane is first introduced (before the reaction starts).
- (e) The *n*-butane reacts until equilibrium has been established at $25^\circ C$.
- Calculate the total pressure in the container at equilibrium. Justify your answer.
 - Calculate the molar concentration of each species at equilibrium.
 - If the volume of the system is reduced to half of its original volume, what will be the new concentration of *n*-butane after equilibrium has been reestablished at $25^\circ C$? Justify your answer.

Suppose that in another experiment a 0.010 mol sample of pure isobutane is placed in an evacuated 1.0 L rigid container and allowed to come to equilibrium at $25^\circ C$.

- (f) Calculate the molar concentration of each species after equilibrium has been established.



A 4.32 g sample of liquid SO_2Cl_2 is placed in a rigid, evacuated 1.50 L reaction vessel. As the container is heated to 400. K, the sample vaporizes completely and starts to decompose according to the equation above. The decomposition reaction is endothermic.

- (a) If no decomposition occurred, what would be the pressure, in atm, of the $\text{SO}_2\text{Cl}_2(\text{g})$ in the vessel at 400. K ?
- (b) When the system has reached equilibrium at 400. K, the total pressure in the container is 1.26 atm. Calculate the partial pressures, in atm, of $\text{SO}_2\text{Cl}_2(\text{g})$, $\text{SO}_2(\text{g})$, and $\text{Cl}_2(\text{g})$ in the container at 400. K.
- (c) For the decomposition reaction at 400. K,
- write the equilibrium-constant expression for K_p for the reaction, and
 - calculate the value of the equilibrium constant, K_p .
- (d) The temperature of the equilibrium mixture is increased to 425 K. Will the value of K_p increase, decrease, or remain the same? Justify your prediction.
- (e) In another experiment, the original partial pressures of $\text{SO}_2\text{Cl}_2(\text{g})$, $\text{SO}_2(\text{g})$, and $\text{Cl}_2(\text{g})$ are 1.0 atm each at 400. K. Predict whether the amount of $\text{SO}_2\text{Cl}_2(\text{g})$ in the container will increase, decrease, or remain the same. Justify your prediction.

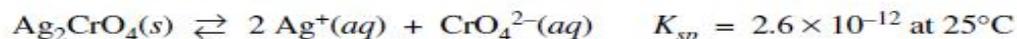
2001.1

Answer the following questions relating to the solubility of the chlorides of silver and lead.

- (a) At 10°C , 8.9×10^{-5} g of $\text{AgCl}(s)$ will dissolve in 100. mL of water.
- Write the equation for the dissociation of $\text{AgCl}(s)$ in water.
 - Calculate the solubility, in mol L^{-1} , of $\text{AgCl}(s)$ in water at 10°C .
- (b) At 25°C , the value of K_{sp} for $\text{PbCl}_2(s)$ is 1.6×10^{-5} and the value of K_{sp} for $\text{AgCl}(s)$ is 1.8×10^{-10} .
- If 60.0 mL of 0.0400 M $\text{NaCl}(aq)$ is added to 60.0 mL of 0.0300 M $\text{Pb}(\text{NO}_3)_2(aq)$, will a precipitate form? Assume that volumes are additive. Show calculations to support your answer.
 - Calculate the equilibrium value of $[\text{Pb}^{2+}(aq)]$ in 1.00 L of saturated PbCl_2 solution to which 0.250 mole of $\text{NaCl}(s)$ has been added. Assume that no volume change occurs.
 - If 0.100 M $\text{NaCl}(aq)$ is added slowly to a beaker containing both 0.120 M $\text{AgNO}_3(aq)$ and 0.150 M $\text{Pb}(\text{NO}_3)_2(aq)$ at 25°C , which will precipitate first, $\text{AgCl}(s)$ or $\text{PbCl}_2(s)$? Show calculations to support your answer.

Answer the following questions relating to the solubilities of two silver compounds, Ag_2CrO_4 and Ag_3PO_4 .

Silver chromate dissociates in water according to the equation shown below.



- Write the equilibrium-constant expression for the dissolving of $\text{Ag}_2\text{CrO}_4(s)$.
- Calculate the concentration, in mol L^{-1} , of $\text{Ag}^+(aq)$ in a saturated solution of Ag_2CrO_4 at 25°C .
- Calculate the maximum mass, in grams, of Ag_2CrO_4 that can dissolve in 100. mL of water at 25°C .
- A 0.100 mol sample of solid AgNO_3 is added to a 1.00 L saturated solution of Ag_2CrO_4 . Assuming no volume change, does $[\text{CrO}_4^{2-}]$ increase, decrease, or remain the same? Justify your answer.

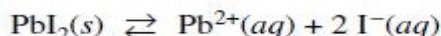
In a saturated solution of Ag_3PO_4 at 25°C , the concentration of $\text{Ag}^+(aq)$ is $5.3 \times 10^{-5} \text{ M}$. The equilibrium-constant expression for the dissolving of $\text{Ag}_3\text{PO}_4(s)$ in water is shown below.



- Write the balanced equation for the dissolving of Ag_3PO_4 in water.
- Calculate the value of K_{sp} for Ag_3PO_4 at 25°C .
- A 1.00 L sample of saturated Ag_3PO_4 solution is allowed to evaporate at 25°C to a final volume of 500. mL. What is $[\text{Ag}^+]$ in the solution? Justify your answer.

Answer the following questions that relate to solubility of salts of lead and barium.

- (a) A saturated solution is prepared by adding excess $\text{PbI}_2(s)$ to distilled water to form 1.0 L of solution at 25°C . The concentration of $\text{Pb}^{2+}(aq)$ in the saturated solution is found to be $1.3 \times 10^{-3} M$. The chemical equation for the dissolution of $\text{PbI}_2(s)$ in water is shown below.



- (i) Write the equilibrium-constant expression for the equation.
 - (ii) Calculate the molar concentration of $\text{I}^{-}(aq)$ in the solution.
 - (iii) Calculate the value of the equilibrium constant, K_{sp} .
- (b) A saturated solution is prepared by adding $\text{PbI}_2(s)$ to distilled water to form 2.0 L of solution at 25°C . What are the molar concentrations of $\text{Pb}^{2+}(aq)$ and $\text{I}^{-}(aq)$ in the solution? Justify your answer.
- (c) Solid NaI is added to a saturated solution of PbI_2 at 25°C . Assuming that the volume of the solution does not change, does the molar concentration of $\text{Pb}^{2+}(aq)$ in the solution increase, decrease, or remain the same? Justify your answer.
- (d) The value of K_{sp} for the salt BaCrO_4 is 1.2×10^{-10} . When a 500. mL sample of $8.2 \times 10^{-6} M$ $\text{Ba}(\text{NO}_3)_2$ is added to 500. mL of $8.2 \times 10^{-6} M$ Na_2CrO_4 , no precipitate is observed.- (i) Assuming that volumes are additive, calculate the molar concentrations of $\text{Ba}^{2+}(aq)$ and $\text{CrO}_4^{2-}(aq)$ in the 1.00 L of solution.
- (ii) Use the molar concentrations of $\text{Ba}^{2+}(aq)$ ions and $\text{CrO}_4^{2-}(aq)$ ions as determined above to show why a precipitate does not form. You must include a calculation as part of your answer.

2010.1

Several reactions are carried out using AgBr, a cream-colored silver salt for which the value of the solubility-product constant, K_{sp} , is 5.0×10^{-13} at 298 K.

- (a) Write the expression for the solubility-product constant, K_{sp} , of AgBr.
- (b) Calculate the value of $[\text{Ag}^+]$ in 50.0 mL of a saturated solution of AgBr at 298 K.
- (c) A 50.0 mL sample of distilled water is added to the solution described in part (b), which is in a beaker with some solid AgBr at the bottom. The solution is stirred and equilibrium is reestablished. Some solid AgBr remains in the beaker. Is the value of $[\text{Ag}^+]$ greater than, less than, or equal to the value you calculated in
- (d) Calculate the minimum volume of distilled water, in liters, necessary to completely dissolve a 5.0 g sample of AgBr(s) at 298 K. (The molar mass of AgBr is 188 g mol^{-1} .)
- (e) A student mixes 10.0 mL of $1.5 \times 10^{-4} \text{ M AgNO}_3$ with 2.0 mL of $5.0 \times 10^{-4} \text{ M NaBr}$ and stirs the resulting mixture. What will the student observe? Justify your answer with calculations.
- (f) The color of another salt of silver, AgI(s), is yellow. A student adds a solution of NaI to a test tube containing a small amount of solid, cream-colored AgBr. After stirring the contents of the test tube, the student observes that the solid in the test tube changes color from cream to yellow.
 - (i) Write the chemical equation for the reaction that occurred in the test tube.
 - (ii) Which salt has the greater value of K_{sp} : AgBr or AgI? Justify your answer.

2013A.1

Answer the following questions about the solubility of some fluoride salts of alkaline earth metals.

- (a) A student prepares 100. mL of a saturated solution of MgF_2 by adding 0.50 g of solid MgF_2 to 100. mL of distilled water at 25°C and stirring until no more solid dissolves. (Assume that the volume of the undissolved MgF_2 is negligibly small.) The saturated solution is analyzed, and it is determined that $[\text{F}^-]$ in the solution is $2.4 \times 10^{-3} \text{ M}$.

- (i) Write the chemical equation for the dissolving of solid MgF_2 in water.
- (ii) Calculate the number of moles of MgF_2 that dissolved.
- (iii) Determine the value of the solubility-product constant, K_{sp} , for MgF_2 at 25°C .

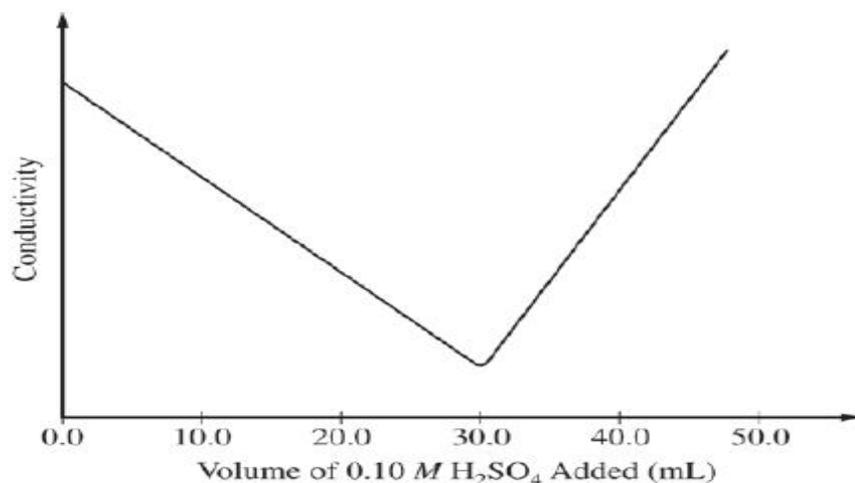
- (b) A beaker contains 500. mL of a solution in which both $\text{Ca}^{2+}(\text{aq})$ and $\text{Ba}^{2+}(\text{aq})$ are present at a concentration of 0.10 M at 25°C . A student intends to separate the ions by adding 0.20 M NaF solution one drop at a time from a buret. At 25°C the value of K_{sp} for CaF_2 is 3.5×10^{-11} ; the value of K_{sp} for BaF_2 is 1.8×10^{-6} .

- (i) Which salt will precipitate first, CaF_2 or BaF_2 ? Justify your answer.

For parts (b)(ii) and (b)(iii) below, assume that the addition of the NaF solution does not significantly affect the total volume of the liquid in the beaker.

- (ii) Calculate the minimum concentration of $\text{F}^-(\text{aq})$ necessary to initiate precipitation of the salt selected in part (b)(i).
- (iii) Calculate the minimum volume of 0.20 M NaF that must be added to the beaker to initiate precipitation of the salt selected in part (b)(i).

- (c) There are several ways to dissolve salts that have limited solubility. Describe one procedure to redissolve the precipitate formed in part (b).



A student performs an experiment in which the conductivity of a solution of $\text{Ba}(\text{OH})_2$ is monitored as the solution is titrated with $0.10\text{ M H}_2\text{SO}_4$. The original volume of the $\text{Ba}(\text{OH})_2$ solution is 25.0 mL . A precipitate of BaSO_4 ($K_{sp} = 1.0 \times 10^{-10}$) formed during the titration. The data collected from the experiment are plotted in the graph above.

- (a) As the first 30.0 mL of $0.10\text{ M H}_2\text{SO}_4$ are added to the $\text{Ba}(\text{OH})_2$ solution, two types of chemical reactions occur simultaneously. On the lines provided below, write the balanced net-ionic equations for (i) the neutralization reaction and (ii) the precipitation reaction.
- (i) Equation for neutralization reaction: _____
- (ii) Equation for precipitation reaction: _____
- (b) The conductivity of the $\text{Ba}(\text{OH})_2$ solution decreases as the volume of added $0.10\text{ M H}_2\text{SO}_4$ changes from 0.0 mL to 30.0 mL .
- (i) Identify the chemical species that enable the solution to conduct electricity as the first 30.0 mL of $0.10\text{ M H}_2\text{SO}_4$ are added.
- (ii) On the basis of the equations you wrote in part (a), explain why the conductivity decreases.
- (c) Using the information in the graph, calculate the molarity of the original $\text{Ba}(\text{OH})_2$ solution.
- (d) Calculate the concentration of $\text{Ba}^{2+}(\text{aq})$ in the solution at the equivalence point (after exactly 30.0 mL of $0.10\text{ M H}_2\text{SO}_4$ are added).
- (e) The concentration of $\text{Ba}^{2+}(\text{aq})$ in the solution decreases as the volume of added $0.10\text{ M H}_2\text{SO}_4$ increases from 30.0 mL to 31.0 mL . Explain.

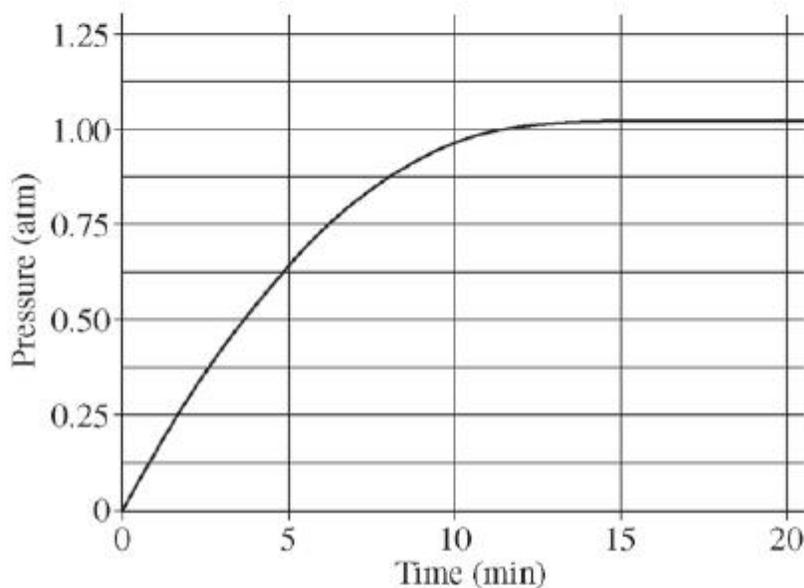
Mass of KI tablet	0.425 g
Mass of thoroughly dried filter paper	1.462 g
Mass of filter paper + precipitate after first drying	1.775 g
Mass of filter paper + precipitate after second drying	1.699 g
Mass of filter paper + precipitate after third drying	1.698 g

A student is given the task of determining the I^- content of tablets that contain KI and an inert, water-soluble sugar as a filler. A tablet is dissolved in 50.0 mL of distilled water, and an excess of 0.20 M $Pb(NO_3)_2(aq)$ is added to the solution. A yellow precipitate forms, which is then filtered, washed, and dried. The data from the experiment are shown in the table above.

- For the chemical reaction that occurs when the precipitate forms,
 - write a balanced, net-ionic equation for the reaction, and
 - explain why the reaction is best represented by a net-ionic equation.
- Explain the purpose of drying and weighing the filter paper with the precipitate three times.
- In the filtrate solution, is $[K^+]$ greater than, less than, or equal to $[NO_3^-]$? Justify your answer.
- Calculate the number of moles of precipitate that is produced in the experiment.
- Calculate the mass percent of I^- in the tablet.
- In another trial, the student dissolves a tablet in 55.0 mL of water instead of 50.0 mL of water. Predict whether the experimentally determined mass percent of I^- will be greater than, less than, or equal to the amount calculated in part (e). Justify your answer.
- A student in another lab also wants to determine the I^- content of a KI tablet but does not have access to $Pb(NO_3)_2$. However, the student does have access to 0.20 M $AgNO_3$, which reacts with $I^-(aq)$ to produce $AgI(s)$. The value of K_{sp} for AgI is 8.5×10^{-17} .
 - Will the substitution of $AgNO_3$ for $Pb(NO_3)_2$ result in the precipitation of the I^- ion from solution? Justify your answer.
 - The student only has access to one KI tablet and a balance that can measure to the nearest 0.01 g. Will the student be able to determine the mass of AgI produced to three significant figures? Justify your answer.



When heated, calcium carbonate decomposes according to the equation above. In a study of the decomposition of calcium carbonate, a student added a 50.0 g sample of powdered $\text{CaCO}_3(s)$ to a 1.00 L rigid container. The student sealed the container, pumped out all the gases, then heated the container in an oven at 1100 K. As the container was heated, the total pressure of the $\text{CO}_2(g)$ in the container was measured over time. The data are plotted in the graph below.



The student repeated the experiment, but this time the student used a 100.0 g sample of powdered $\text{CaCO}_3(s)$. In this experiment, the final pressure in the container was 1.04 atm, which was the same final pressure as in the first experiment.

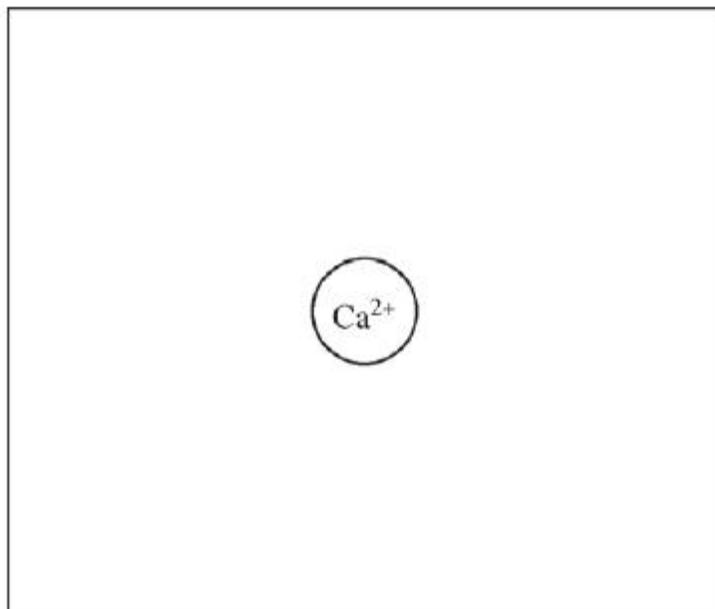
- Calculate the number of moles of $\text{CO}_2(g)$ present in the container after 20 minutes of heating.
- The student claimed that the final pressure in the container in each experiment became constant because all of the $\text{CaCO}_3(s)$ had decomposed. Based on the data in the experiments, do you agree with this claim? Explain.
- After 20 minutes some $\text{CO}_2(g)$ was injected into the container, initially raising the pressure to 1.5 atm. Would the final pressure inside the container be less than, greater than, or equal to 1.04 atm? Explain your reasoning.
- Are there sufficient data obtained in the experiments to determine the value of the equilibrium constant, K_p , for the decomposition of $\text{CaCO}_3(s)$ at 1100 K? Justify your answer.

2015.4

Answer the following questions about the solubility of $\text{Ca}(\text{OH})_2$ ($K_{sp} = 1.3 \times 10^{-6}$).

- (a) Write a balanced chemical equation for the dissolution of $\text{Ca}(\text{OH})_2(s)$ in pure water.
- (b) Calculate the molar solubility of $\text{Ca}(\text{OH})_2$ in 0.10 M $\text{Ca}(\text{NO}_3)_2$.
- (c) In the box below, complete a particle representation diagram that includes four water molecules with proper orientation around the Ca^{2+} ion.

Represent water molecules as .



Answer the following questions about Fe and Al compounds.

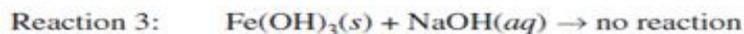
- (a) $\text{Fe}_2\text{O}_3(s)$ and $\text{Al}_2\text{O}_3(s)$ have similar chemical properties; some similarities are due to the oxides having similar lattice energies. Give two reasons why the lattice energies of the oxides are similar.

Use the following reactions that involve Fe and Al compounds to answer parts (b) and (c).

In distilled water



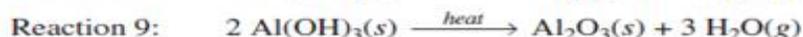
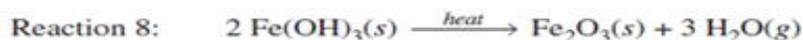
In base



In acid



When heated



Compound	K_{sp}
$\text{Fe}(\text{OH})_3$	4×10^{-38}
$\text{Al}(\text{OH})_3$	1×10^{-33}

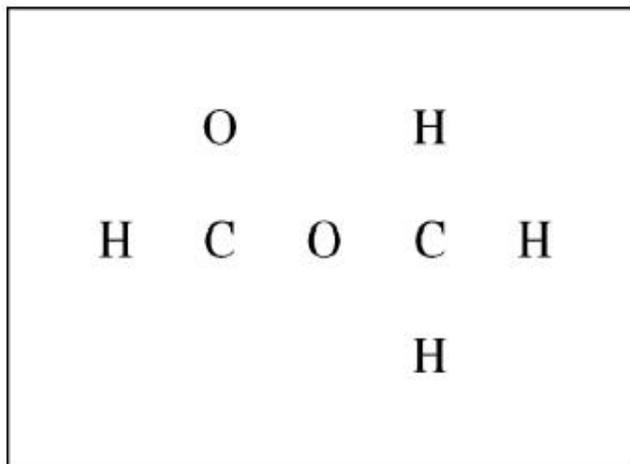
- (b) The K_{sp} values for $\text{Fe}(\text{OH})_3$ and $\text{Al}(\text{OH})_3$ are given in the table above. A 1.0 g sample of powdered $\text{Fe}_2\text{O}_3(s)$ and a 1.0 g sample of powdered $\text{Al}_2\text{O}_3(s)$ are mixed together and placed in 1.0 L of distilled water.
- Which ion, $\text{Fe}^{3+}(aq)$ or $\text{Al}^{3+}(aq)$, will be present in the higher concentration? Justify your answer with respect to the K_{sp} values provided.
 - Write a balanced chemical equation for the dissolution reaction that results in the production of the ion that you identified in part (i).
- (c) Students are asked to develop a plan for separating $\text{Al}_2\text{O}_3(s)$ from a mixture of powdered $\text{Fe}_2\text{O}_3(s)$ and powdered $\text{Al}_2\text{O}_3(s)$ using chemical reactions and laboratory techniques.
- One student proposes that $\text{Al}_2\text{O}_3(s)$ can be separated from the mixture by adding water to the mixture and then filtering. Explain why this approach is not reasonable.
 - A second student organizes a plan using a table. The first two steps have already been entered in the table, as shown below. Complete the plan by listing the additional steps that are needed to recover the $\text{Al}_2\text{O}_3(s)$. List the steps in the correct order and refer to the appropriate reaction by number, if applicable.

Step	Description	Reaction(s)
1	Add $\text{NaOH}(aq)$ to convert $\text{Al}_2\text{O}_3(s)$ to $\text{Al}(\text{OH})_3(s)$ and then to $\text{NaAl}(\text{OH})_4(aq)$.	2 and 4
2	Filter out the solid $\text{Fe}(\text{OH})_3$ from the mixture and save the filtrate.	—
3		
4		
5		

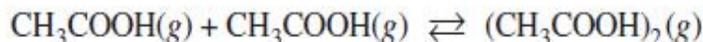
- The second student recovers 5.5 g of $\text{Al}_2\text{O}_3(s)$ from a 10.0 g sample of the mixture. Calculate the percent of Al by mass in the mixture of the two powdered oxides. (The molar mass of Al_2O_3 is 101.96 g/mol, and the molar mass of Fe_2O_3 is 159.70 g/mol.)

Answer the following questions about two isomers, methyl methanoate and ethanoic acid. The molecular formula of the compounds is $C_2H_4O_2$.

- (a) Complete the Lewis electron-dot diagram of methyl methanoate in the box below. Show all valence electrons.



A student puts 0.020 mol of methyl methanoate into a previously evacuated rigid 1.0 L vessel at 450 K. The pressure is measured to be 0.74 atm. When the experiment is repeated using 0.020 mol of ethanoic acid instead of methyl methanoate, the measured pressure is lower than 0.74 atm. The lower pressure for ethanoic acid is due to the following reversible reaction.



- (b) Assume that when equilibrium has been reached, 50. percent of the ethanoic acid molecules have reacted.
- (i) Calculate the total pressure in the vessel at equilibrium at 450 K.
 - (ii) Calculate the value of the equilibrium constant, K_p , for the reaction at 450 K.



The polyatomic ion $\text{C}_{10}\text{H}_{12}\text{N}_2\text{O}_8^{4-}$ is commonly abbreviated as EDTA^{4-} . The ion can form complexes with metal ions in aqueous solutions. A complex of EDTA^{4-} with Ba^{2+} ion forms according to the equation above. A 50.0 mL volume of a solution that has an $\text{EDTA}^{4-}(\text{aq})$ concentration of 0.30 M is mixed with 50.0 mL of 0.20 M $\text{Ba}(\text{NO}_3)_2$ to produce 100.0 mL of solution.

- Considering the value of K for the reaction, determine the concentration of $\text{Ba}(\text{EDTA})^{2-}(\text{aq})$ in the 100.0 mL of solution. Justify your answer.
- The solution is diluted with distilled water to a total volume of 1.00 L. After equilibrium has been reestablished, is the number of moles of $\text{Ba}^{2+}(\text{aq})$ present in the solution greater than, less than, or equal to the number of moles of $\text{Ba}^{2+}(\text{aq})$ present in the original solution before it was diluted? Justify your answer.

2017.6

Answer the following questions about $\text{Mg}(\text{OH})_2$. At 25°C , the value of the solubility product constant, K_{sp} , for $\text{Mg}(\text{OH})_2(s)$ is 1.8×10^{-11} .

- (a) Calculate the number of grams of $\text{Mg}(\text{OH})_2$ (molar mass 58.32 g/mol) that is dissolved in 100. mL of a saturated solution of $\text{Mg}(\text{OH})_2$ at 25°C .
- (b) The energy required to separate the ions in the $\text{Mg}(\text{OH})_2$ crystal lattice into individual $\text{Mg}^{2+}(g)$ and $\text{OH}^-(g)$ ions, as represented in the table below, is known as the lattice energy of $\text{Mg}(\text{OH})_2(s)$. As shown in the table, the lattice energy of $\text{Sr}(\text{OH})_2(s)$ is less than the lattice energy of $\text{Mg}(\text{OH})_2(s)$. Explain why in terms of periodic properties and Coulomb's law.

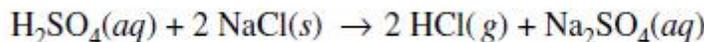
Reaction	Lattice Energy (kJ/mol)
$\text{Mg}(\text{OH})_2(s) \rightarrow \text{Mg}^{2+}(g) + 2 \text{OH}^-(g)$	2900
$\text{Sr}(\text{OH})_2(s) \rightarrow \text{Sr}^{2+}(g) + 2 \text{OH}^-(g)$	2300

Answer the following questions about the solubility of $\text{AgCl}(s)$. The value of K_{sp} for $\text{AgCl}(s)$ is 1.8×10^{-10} .

- (a) Calculate the value of $[\text{Ag}^+]$ in a saturated solution of AgCl in distilled water.
- (b) The concentration of $\text{Cl}^-(aq)$ in seawater is $0.54 M$.
 - (i) Calculate the molar solubility of $\text{AgCl}(s)$ in seawater.
 - (ii) Explain why $\text{AgCl}(s)$ is less soluble in seawater than in distilled water.

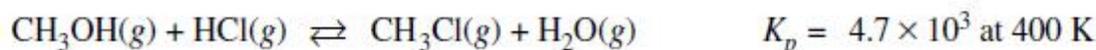
Answer the following questions relating to HCl, CH₃Cl, and CH₃Br.

- (a) HCl(g) can be prepared by the reaction of concentrated H₂SO₄(aq) with NaCl(s), as represented by the following equation.



- (i) A student claims that the reaction is a redox reaction. Is the student correct? Justify your answer.
 (ii) Calculate the mass, in grams, of NaCl(s) needed to react with excess H₂SO₄(aq) to produce 3.00 g of HCl(g). Assume that the reaction goes to completion.

HCl(g) can react with methanol vapor, CH₃OH(g), to produce CH₃Cl(g), as represented by the following equation.



- (b) CH₃OH(g) and HCl(g) are combined in a 10.00 L sealed reaction vessel and allowed to reach equilibrium at 400 K. The initial partial pressure of CH₃OH(g) in the vessel is 0.250 atm and that of HCl(g) is 0.600 atm.
- (i) Does the total pressure in the vessel increase, decrease, or remain the same as equilibrium is approached? Justify your answer in terms of the reaction stoichiometry.
 (ii) Considering the value of K_p , calculate the final partial pressure of HCl(g) after the system inside the vessel reaches equilibrium at 400 K.
 (iii) The student claims that the final partial pressure of CH₃OH(g) at equilibrium is very small but not exactly zero. Do you agree or disagree with the student's claim? Justify your answer.

- (c) The table below shows some data for the compounds CH₃Cl and CH₃Br.

Compound	Molar Mass (g/mol)	Dipole Moment (D)	Boiling Point (K)	$\Delta H_{\text{vaporization}}^{\circ}$ (kJ/mol)
CH ₃ Cl	50.5	1.87	249	18.9
CH ₃ Br	94.9	1.81	277	22.8

- (i) Identify all the types of intermolecular forces that exist among molecules in CH₃Cl(l).
 (ii) In terms of intermolecular forces, explain why the boiling point of CH₃Br(l) is greater than that of CH₃Cl(l).
- (d) A 2.00 mL sealed glass vial containing a 1.00 g sample of CH₃Cl(l) is stored in a freezer at 233 K.
- (i) Calculate the pressure in the vial at 298 K assuming that all the CH₃Cl(l) vaporizes.
 (ii) Explain why it would be unsafe to remove the vial from the freezer and leave it on a lab bench at 298 K.