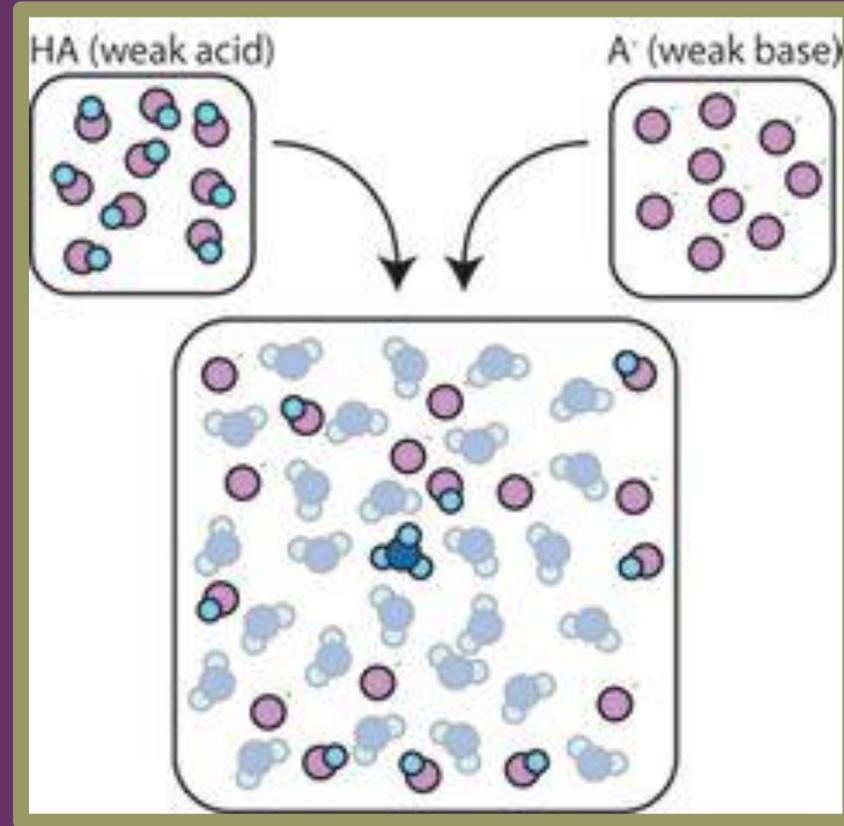


+

Big Idea #6

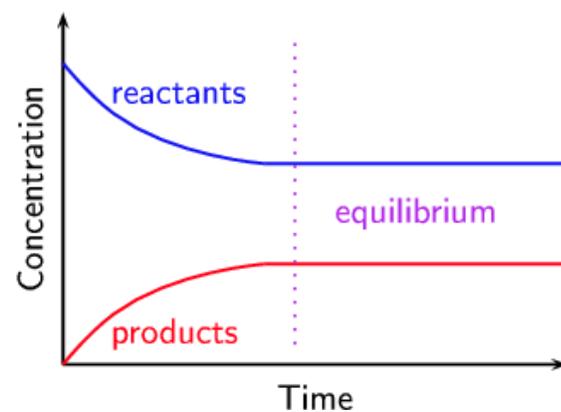
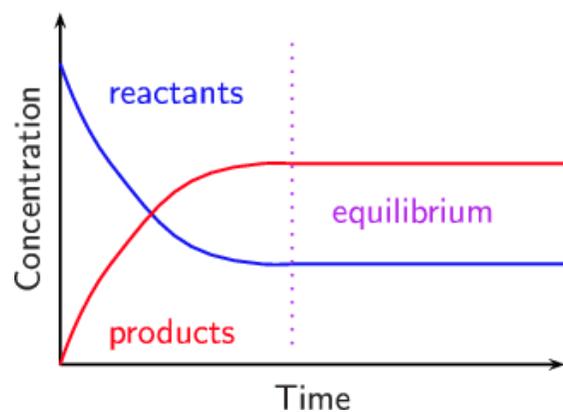
Equilibrium





What is chemical equilibrium?

- Systems that have reached the state where the rates of the forward reaction and the reverse reaction are constant and equal.
- It is a dynamic process where reactants continuously form products and vice versa, but the net amounts of reactants and products remain constant.
- The proportions of products and reactants formed in a system at a specific temperature that has achieved equilibrium is represented by K , the equilibrium constant.

[Video](#)


In the reaction:
 $aA + bB \rightleftharpoons cC + dD$

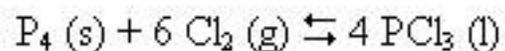
$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

[A] = concentration of A in mol dm^{-3}
 a = number of moles of A

LO 6.1: Given a set of experimental observations regarding processes that are reversible, construct an explanation that connects the observations to the reversibility of the underlying chemical reactions or processes.

+ Manipulating Q and K

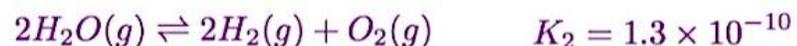
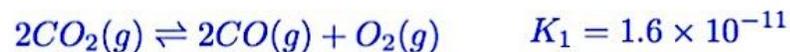
- K (equilibrium constant) represents the relative amounts of products to reactants at equilibrium at a given temperature.
- Q (reaction progress) describes the relative amounts of products to reactants present at any point in the reaction at a given temperature.
- Q and K only include substances that are gases or in aqueous solutions. No solids or liquids are ever included in these expressions.



$$K = \frac{1}{[Cl_2]^6}$$

$$K_p = \frac{1}{P_{Cl_2}^6}$$

1)



Click reveals answer.

- Similar reactions will have related K values at the same temperature.

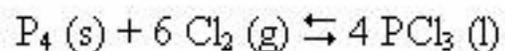
Reaction Run in Reverse	Inverse the value of K (i.e. raise it to the -1 power)
Multiplying by a Constant	Raise the equilibrium constant to the power that was multiplied by
Adding Reactions	Multiply the equilibrium constant by one another

[Video](#)

LO 6.2: The student can, given a manipulation of a chemical reaction or set of reactions (e.g., reversal of reaction or addition of two reactions), determine the effects of that manipulation on Q or K .

+ Manipulating Q and K

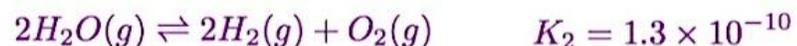
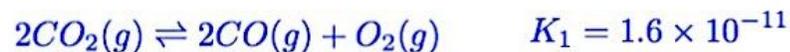
- K (equilibrium constant) represents the relative amounts of products to reactants at equilibrium at a given temperature.
- Q (reaction progress) describes the relative amounts of products to reactants present at any point in the reaction at a given temperature.
- Q and K only include substances that are gases or in aqueous solutions. No solids or liquids are ever included in these expressions.



$$K = \frac{1}{[Cl_2]^6}$$

$$K_p = \frac{1}{P_{Cl_2}^6}$$

1)



.35

- Similar reactions will have related K values at the same temperature.

Reaction Run in Reverse	Inverse the value of K (i.e. raise it to the -1 power)
Multiplying by a Constant	Raise the equilibrium constant to the power that was multiplied by
Adding Reactions	Multiply the equilibrium constant by one another

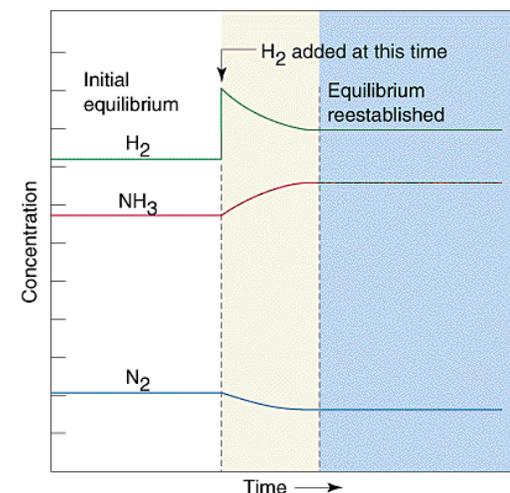
Video

LO 6.2: The student can, given a manipulation of a chemical reaction or set of reactions (e.g., reversal of reaction or addition of two reactions), determine the effects of that manipulation on Q or K .



Kinetics and Equilibrium

- Kinetics examines the rate at which reactions proceed. Rate laws are used to describe how reactant concentrations affect a reaction's rate. Rate constants (k) in rate law expressions are determined experimentally at a given temperature.
- Equilibrium describes the state at which the rates of the forward reaction and the reverse reaction are constant and equal.
- If the rates are initially unequal (the system is not at equilibrium), the faster direction depletes its reactants, which feeds back to slow down that direction.
- At the same time, the slower direction accumulates its reactants, speeding up the slower direction.
- These loops continue until the faster rate and the slower rate have become equal.
- In the graph to the right, after equilibrium has been achieved, additional hydrogen gas is added to the system. The system then consumes both H_2 and N_2 to form additional NH_3 molecules, eventually reestablishing equilibrium.

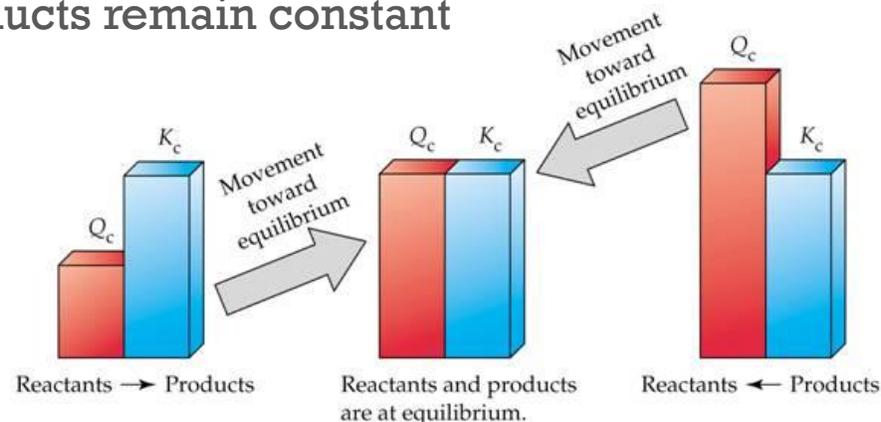


LO 6.3: The student can connect kinetics to equilibrium by using reasoning, such as LeChatelier's principle, to infer the relative rates of the forward and reverse reactions.



Q vs. K

- Equilibrium is reached when the rates of the forward reaction and the rates of the reverse reaction are equal, which is when Q is equal to K .
- Comparing Q to K enables us to determine if a chemical system has achieved equilibrium or will need to move towards reactants or products to reach equilibrium.
 - if $Q < K$, the reaction will proceed in the forward direction until $Q = K$
 - if $Q > K$, the reaction will proceed in the reverse direction until $Q = K$
 - if $Q = K$, the reaction is at equilibrium, and the concentrations of reactants and products remain constant

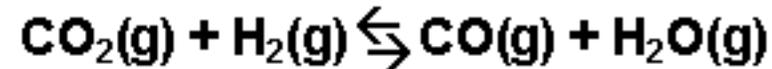
[Video](#)

LO 6.4: Given a set of initial conditions and the equilibrium constant, K , use the tendency of Q to approach K to predict and justify the prediction as to whether the reaction will proceed toward products or reactants as equilibrium is approached.



Calculating K

- Equilibrium constants can be determined using experimental concentrations of reactants and products at equilibrium.
 - Steps:
 - 1) Write an equilibrium expression
 - 2) Determine equilibrium molar concentrations or partial pressures for all substances in expression
 - 3) Substitute quantities into equilibrium expression and solve.
- 2) ■ Example: Calculate K for the following system if 0.1908 mol CO_2 , 0.0908 mol H_2 , 0.0092 CO , and 0.0092 mol H_2O vapor are present in a 2.00 L vessel at equilibrium:



Click reveals answer and explanation.

[Video](#)

LO 6.5: The student can, given data (tabular, graphical, etc.) from which the state of a system at equilibrium can be obtained, calculate the equilibrium constant, K .



Calculating K

2) Write the [equilibrium expression](#) for the reaction system.

$$K_c = \frac{[\text{CO}][\text{H}_2\text{O}]}{[\text{CO}_2][\text{H}_2]}$$

- Since K_c is being determined, check to see if the given equilibrium amounts are expressed in moles per liter ([molarity](#)). In this example they are not; conversion of each is required.

$$[\text{CO}_2] = 0.1908 \text{ mol CO}_2 / 2.00 \text{ L} = 0.0954 \text{ M}$$

$$[\text{H}_2] = 0.0454 \text{ M}$$

$$[\text{CO}] = 0.0046 \text{ M}$$

$$[\text{H}_2\text{O}] = 0.0046 \text{ M}$$

- Substitute each concentration into the equilibrium expression and calculate the value of the equilibrium constant.

$$K_c = \frac{[0.0046][0.0046]}{[0.0954][0.0454]} = 0.0049 \text{ or } 4.9 \times 10^{-3}$$

LO 6.5: The student can, given data (tabular, graphical, etc.) from which the state of a system at equilibrium can be obtained, calculate the equilibrium constant, K .



Calculating K

- Equilibrium constants can also be determined from both initial and equilibrium concentrations using an ICE chart.
- Steps:
 - 1) Write an equilibrium expression
 - 2) Determine molar concentrations or partial pressures for all substances in expression
 - 3) Determine equilibrium concentrations or partial pressures using an [ICE chart](#).
 - 4) Substitute quantities into equilibrium expression and solve.
- 3) ■ Example: Initially, a mixture of 0.100 M NO, 0.050 M H₂, and 0.100 M H₂O was allowed to reach equilibrium. No N₂ was present initially. At equilibrium, NO had a concentration of 0.062 M. Determine the value of K for the reaction: $2 \text{NO}(\text{g}) + 2 \text{H}_2(\text{g}) \rightleftharpoons \text{N}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$

[Video](#)

Click reveals answer and explanation.

LO 6.5: Given data (tabular, graphical, etc.) from which the state of a system at equilibrium can be obtained, calculate the equilibrium constant, K .

3) Write the [equilibrium expression](#) for the reaction.

$$K_c = \frac{[\text{N}_2][\text{H}_2\text{O}]^2}{[\text{NO}]^2[\text{H}_2]^2}$$

- Check to see if the amounts are expressed in moles per liter (molarity) since K_c is being . In this example they are.
- Create an [ICE chart](#) that expresses the initial concentration, the change in concentration, and the equilibrium concentration for each species in the reaction. From the chart you can determine the changes in the concentrations of each species and the equilibrium concentrations. From the example, we start with the following information.

	NO	H ₂	N ₂	H ₂ O
Initial Concentration (M)	0.100	0.0500	0	0.100
Change in Concentration (M)	-2 x	-2 x	+ x	+ 2 x
Equilibrium Concentration (M)	0.062			

The change in concentration of the NO was $(0.062 \text{ M} - 0.100\text{M}) = -0.038 \text{ M}$. Thus $-2x = -0.038$ and $x = 0.019$. Note: the negative sign indicates a decreasing concentration, not a negative concentration. The changes in the other species must agree with the stoichiometry dictated by the balance equation. The hydrogen will also change by -0.038 M , while the nitrogen will increase by $+0.019 \text{ M}$ and the water will increase by $+0.038 \text{ M}$. From these changes we can complete the chart to find the equilibrium concentrations for each species.

	NO	H ₂	N ₂	H ₂ O
Initial Concentration (M)	0.100	0.0500	0	0.100
Change in Concentration (M)	-0.038	-0.038	+0.019	+0.038
Equilibrium Concentration (M)	0.062	0.012	0.019	0.138

- Substitute the equilibrium concentrations into the equilibrium expression and solve for K_c .

$$K_c = \frac{[0.019][0.138]^2}{[0.062]^2[0.012]^2} = 650 \text{ or } 6.5 \times 10^2$$

LO 6.5: Given data (tabular, graphical, etc.) from which the state of a system at equilibrium can be obtained, calculate the equilibrium constant, K .



Calculating Equilibrium Concentrations with K

[Source](#)



- Equilibrium concentrations can be calculated using a K expression, the K constant, and initial concentrations or partial pressures of substances.

- Steps:

- 1) Write an equilibrium expression for the reaction
- 2) Set up an ICE table and fill in “initial” quantities
- 3) Determine “changes” in the system in terms of x needed for the system to achieve equilibrium
- 4) Determine the “equilibrium” values for the system by adding the “initial” and “change” values together
- 5) Solve for x using the K expression and the “equilibrium” values. Verify if the change in initial concentrations is negligible using the 5% rule.
- 6) Determine all equilibrium quantities using the value of x

[Video](#)

4)

- Example: Given the following reaction at 1373 K: $\text{Cl}_2(\text{g}) \rightleftharpoons 2\text{Cl}(\text{g})$, determine the equilibrium partial pressures of all species if 0.500 atm Cl_2 is present initially. $K = 1.13 \times 10^{-4}$ for the reaction at 1373 K.

Click reveals
answer.

LO 6.6: Given a set of initial conditions (concentrations or partial pressures) and K , use stoichiometric relationships and the law of mass action (Q equals K at equilibrium) to determine qualitatively and/or quantitatively the conditions at equilibrium for a system involving a single reversible reaction.



Calculating Equilibrium Concentrations with K

[Source](#)



- Equilibrium concentrations can be calculated using a K expression, the K constant, and initial concentrations or partial pressures of substances.

- Steps:

- 1) Write an equilibrium expression for the reaction
- 2) Set up an ICE table and fill in “initial” quantities
- 3) Determine “changes” in the system in terms of x needed for the system to achieve equilibrium
- 4) Determine the “equilibrium” values for the system by adding the “initial” and “change” values together
- 5) Solve for x using the K expression and the “equilibrium” values. Verify if the change in initial concentrations is negligible using the 5% rule.
- 6) Determine all equilibrium quantities using the value of x

[Video](#)

4)

- Example: Given the following reaction at 1373 K: $\text{Cl}_2(\text{g}) \rightleftharpoons 2\text{Cl}(\text{g})$, determine the equilibrium partial pressures of all species if 0.500 atm Cl_2 is present initially. $K = 1.13 \times 10^{-4}$ for the reaction at 1373 K.

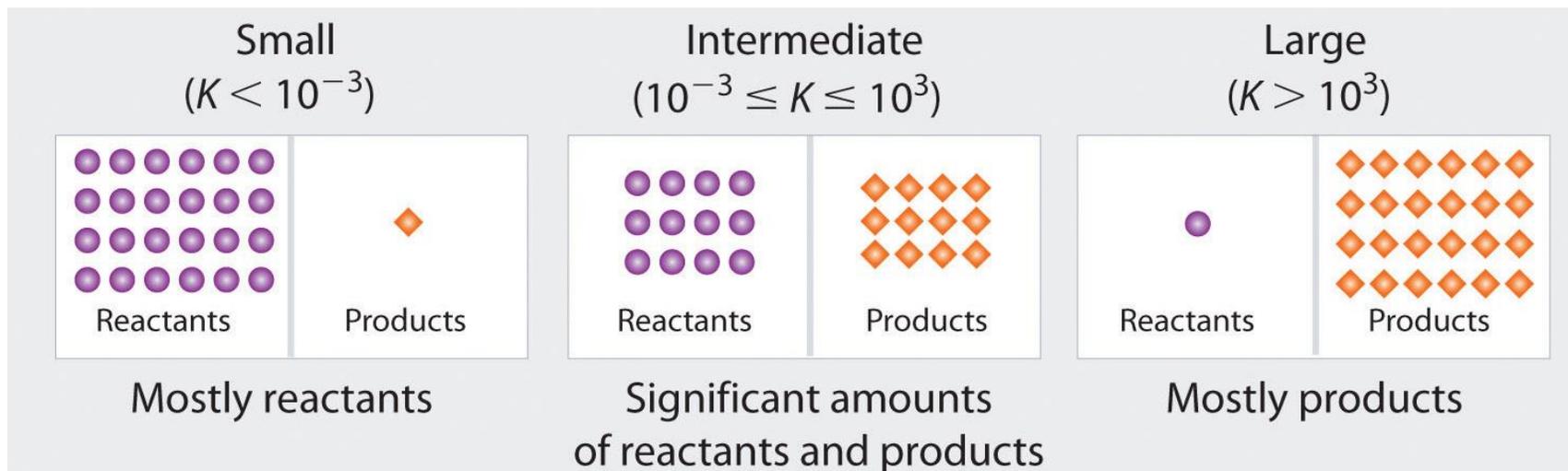
**0.497 atm Cl_2 and
0.00752 atm Cl**

LO 6.6: Given a set of initial conditions (concentrations or partial pressures) and K , use stoichiometric relationships and the law of mass action (Q equals K at equilibrium) to determine qualitatively and/or quantitatively the conditions at equilibrium for a system involving a single reversible reaction.



Magnitude of K

- For many reactions involving aqueous solutions, K is either very large (favoring the forward reaction) or very small (favoring the reverse reaction)
- The size of K can be used to describe the relationship between the numbers of reactant and product particles present at equilibrium.

[Video](#)

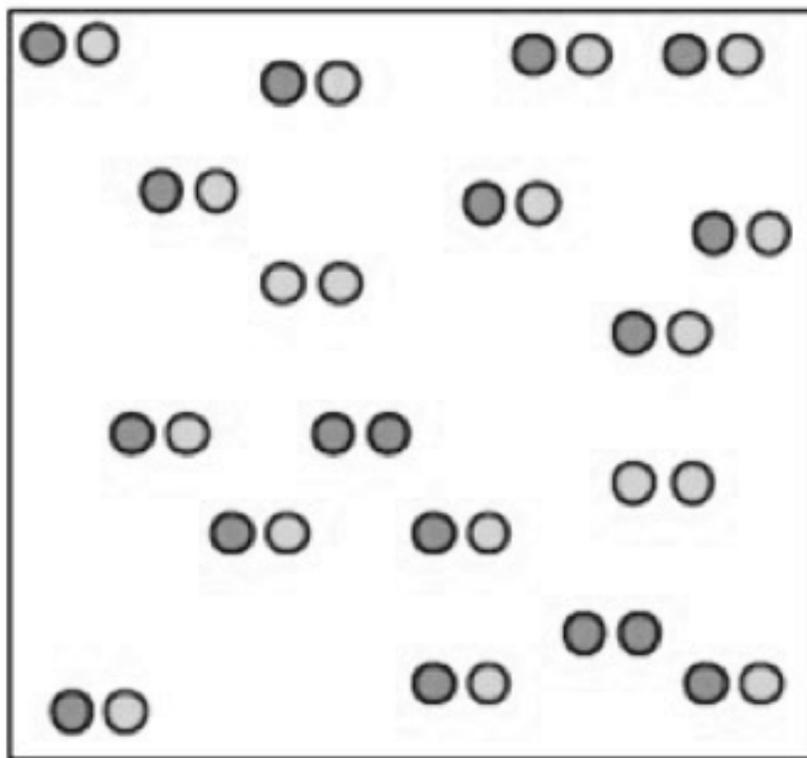
LO 6.7: The student is able, for a reversible reaction that has a large or small K , to determine which chemical species will have very large versus very small concentrations at equilibrium.

+

Magnitude of K

5) The diagram below represents a certain equilibrium position at a certain temperature for the reaction, $\text{H}_{2(g)} + \text{I}_{2(g)} \leftrightarrow 2\text{HI}_{(g)}$.

H_2 is represented by , I_2 is represented by , and HI is represented by .



What can be said about the equilibrium position in terms of reactants and products, and the likely magnitude of K ?

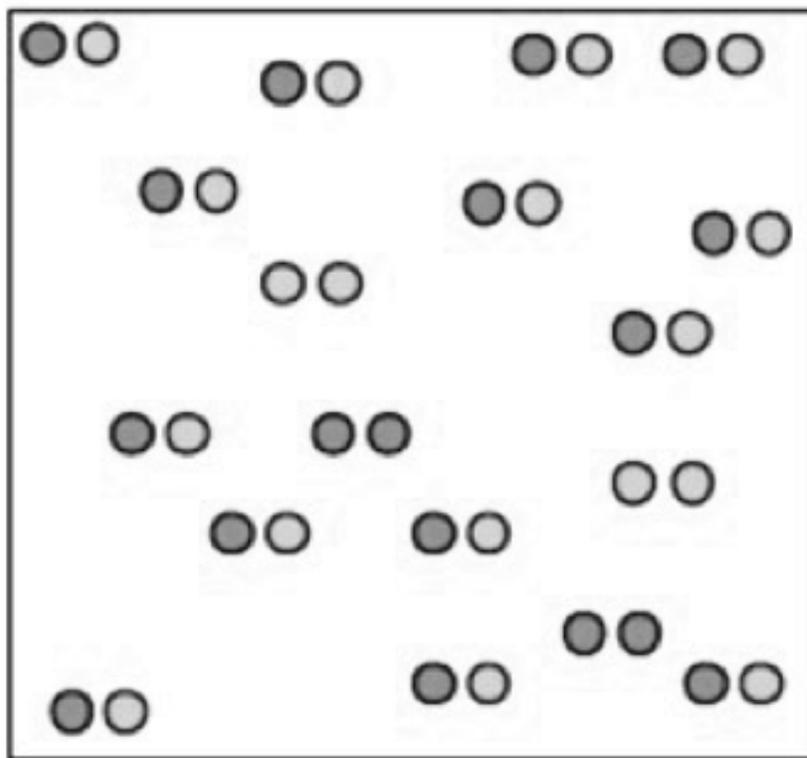
- A. K is large and products are favored
- B. K is small and products are favored
- C. K is small and reactants are favored
- D. K is large and reactants are favored

determine which chemical species will have very large versus very small concentrations at equilibrium.

+ Magnitude of K

5) The diagram below represents a certain equilibrium position at a certain temperature for the reaction, $\text{H}_{2(g)} + \text{I}_{2(g)} \leftrightarrow 2\text{HI}_{(g)}$.

H_2 is represented by , I_2 is represented by , and HI is represented by .



What can be said about the equilibrium position in terms of reactants and products, and the likely magnitude of K ?

- A. K is large and products are favored
- B. K is small and products are favored
- C. K is small and reactants are favored
- D. K is large and reactants are favored

determine which chemical species will have very large versus very small concentrations at equilibrium.



Le Chatelier's Principle

[Animation](#) [Source](#)

- This principle is used to describe changes that occur in a system that has achieved equilibrium. There are three factors that can cause shifts in a system at equilibrium: concentration, pressure, and temperature.

<u>Change</u>	<u>Direction System Shifts to Reestablish Equilibrium</u>
Adding a reactant	Shifts towards products
Adding a product	Shifts towards reactants
Removing a reactant	Shifts towards reactants
Removing a product	Shifts towards products
Increasing pressure (decreasing volume)	Shifts toward less gas molecules
Decreasing pressure (increasing volume)	Shifts towards more gas molecules
Adding an inert gas	No effect
Increasing the temperature	Endothermic: shifts towards products Exothermic: shifts towards reactants
Decreasing the temperature	Endothermic: shifts towards reactants Exothermic: shifts towards products

[Video](#)



LO 6.8: The student is able to use LeChatelier's principle to predict the direction of the shift resulting from various possible stresses on a system at chemical equilibrium.



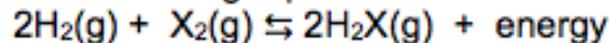
Le Chatelier's Principle

- This principle states that if a system at equilibrium is subjected to a change in concentration, pressure, or temperature, the system will shift its equilibrium position to counteract the change.

- Le Chatelier's Principle
- Equilibrium constants
- Equilibrium constants and Gibbs free energy
- Reaction quotient
- Reaction quotient and Gibbs free energy
- Increasing the concentration of a reactant
- Decreasing the concentration of a reactant
- Adding a catalyst
- Increasing the pressure
- Decreasing the pressure

6)

Consider the following equation:



Addition of X_2 to a system described by the above equilibrium will cause the original:

- $[\text{H}_2]$ to decrease
- $[\text{X}_2]$ to decrease
- $[\text{H}_2]$ to increase
- $[\text{H}_2\text{X}]$ to decrease

Increasing the pressure by decreasing the volume will cause

- $[\text{H}_2]$ and $[\text{H}_2\text{X}]$ to decrease and $[\text{X}_2]$ to increase
- $[\text{H}_2]$ and $[\text{X}_2]$ to decrease and $[\text{H}_2\text{X}]$ to increase
- $[\text{H}_2\text{X}]$ and $[\text{X}_2]$ to decrease and $[\text{H}_2]$ to increase
- $[\text{H}_2\text{X}]$, $[\text{X}_2]$ and $[\text{H}_2]$ to increase

Increasing the temperature will cause

- $[\text{H}_2\text{X}]$ and $[\text{X}_2]$ to decrease and $[\text{H}_2]$ to increase
- $[\text{H}_2\text{X}]$, $[\text{X}_2]$ and $[\text{H}_2]$ to increase
- $[\text{H}_2]$ and $[\text{H}_2\text{X}]$ to decrease and $[\text{X}_2]$ to increase
- $[\text{H}_2]$ and $[\text{X}_2]$ to increase and $[\text{H}_2\text{X}]$ to decrease

LO 6.8: The student will be able to predict the direction of the shift resulting from a change in concentration, pressure, or temperature.

[Animation](#) [Source](#)

a
factors
ration,

[Reestablish](#)

[Video](#)

ts

ts

ts

ts

ecules

olecules

products

products

reactants

products

direction of the shift
orium.



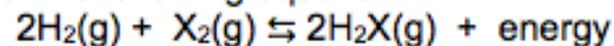


Le C

- This p
system
that c
press

6)

Consider the following equation:



Addition of X_2 to a system described by the above equilibrium will cause the original:

- a) $[\text{H}_2]$ to decrease
- b) $[\text{X}_2]$ to decrease
- c) $[\text{H}_2]$ to increase
- d) $[\text{H}_2\text{X}]$ to decrease

Increasing the pressure by decreasing the volume will cause

- a) $[\text{H}_2]$ and $[\text{H}_2\text{X}]$ to decrease and $[\text{X}_2]$ to increase
- b) $[\text{H}_2]$ and $[\text{X}_2]$ to decrease and $[\text{H}_2\text{X}]$ to increase
- c) $[\text{H}_2\text{X}]$ and $[\text{X}_2]$ to decrease and $[\text{H}_2]$ to increase
- d) $[\text{H}_2\text{X}]$, $[\text{X}_2]$ and $[\text{H}_2]$ to increase

Increasing the temperature will cause

- a) $[\text{H}_2\text{X}]$ and $[\text{X}_2]$ to decrease and $[\text{H}_2]$ to increase
- b) $[\text{H}_2\text{X}]$, $[\text{X}_2]$ and $[\text{H}_2]$ to increase
- c) $[\text{H}_2]$ and $[\text{H}_2\text{X}]$ to decrease and $[\text{X}_2]$ to increase
- d) $[\text{H}_2]$ and $[\text{X}_2]$ to increase and $[\text{H}_2\text{X}]$ to decrease

[Animation](#) [Source](#)

a
factors
ration,

[reestablish](#)

[Video](#)

ts

ts

ts

ts

ecules

olecules

products

products

reactants

products

direction of the shift
orium.

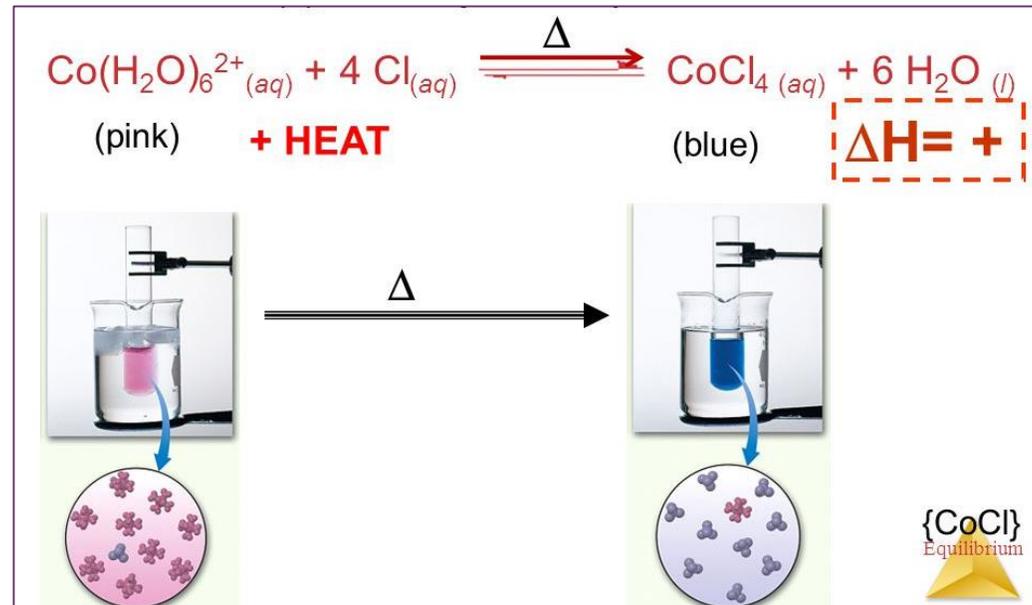
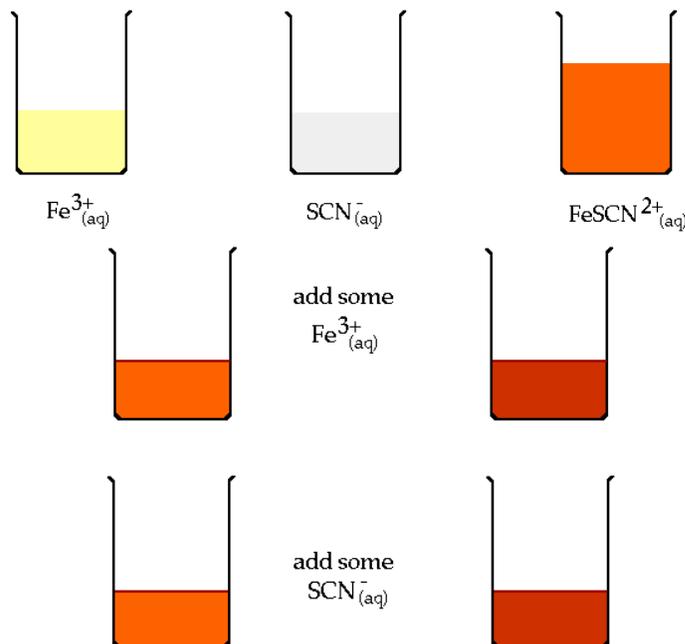
LO 6.8: The stu
resulting from



Experimentally Examining Le Chatelier's Principle

- Systems at equilibrium can be examined using Le Chatelier's Principle by measuring its properties, including pH, temperature, solution color (absorbance)

[Video](#)

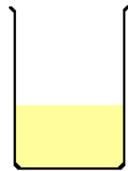
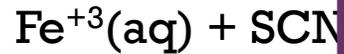


LO 6.9: The student is able to use LeChatelier's principle to design a set of conditions that will optimize a desired outcome, such as product yield.

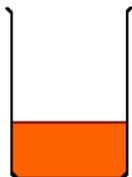
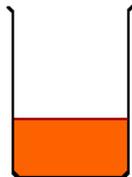
+

Experimentally Examining Le Chatelier's Principle

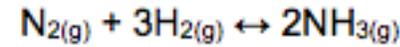
- Systems at Equilibrium
- Principle of Le Chatelier
- Solution Color



$\text{Fe}^{3+}_{(\text{aq})}$



The Haber process is the industrial process that is used to manufacture ammonia gas. It relies upon the establishment of an equilibrium according to the reaction shown below.



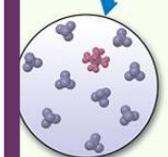
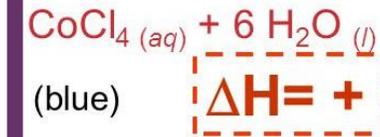
Temperature	Equilibrium constant
200 K	x
300 K	0.0100 x
400 K	0.00025 x

Which of the following will maximize the ammonia formation in equilibrium mixture?

- A. Adding more hydrogen gas
- B. Heating the reaction mixture to 400 K
- C. Increasing the size of the reaction vessel
- D. Adding a catalyst

ure,

[Video](#)



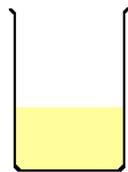
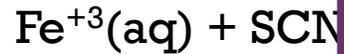
{CoCl}
Equilibrium

LO 6.9: The student is able to use LeChatelier's principle to design a set of conditions that will optimize a desired outcome, such as product yield.

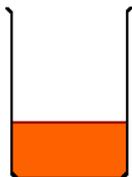
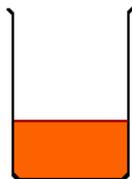
+

Experimentally Examining Le Chatelier's Principle

- Systems at Equilibrium
- Principle of Le Chatelier
- Solution Concentration

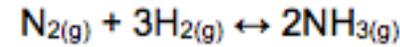


$\text{Fe}^{3+}(\text{aq})$



7)

The Haber process is the industrial process that is used to manufacture ammonia gas. It relies upon the establishment of an equilibrium according to the reaction shown below.



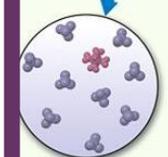
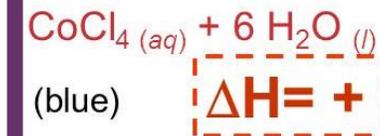
Temperature	Equilibrium constant
200 K	x
300 K	0.0100 x
400 K	0.00025 x

Which of the following will maximize the ammonia formation in equilibrium mixture?

- A. Adding more hydrogen gas
- B. Heating the reaction mixture to 400 K
- C. Increasing the size of the reaction vessel
- D. Adding a catalyst

ure,

[Video](#)



{CoCl}
Equilibrium

LO 6.9: The student is able to use LeChatelier's principle to design a set of conditions that will optimize a desired outcome, such as product yield.



Changes to Q and K for a System at Equilibrium

[Animation](#)



- Some changes that occur to a system at equilibrium will affect the reaction's current position (Q). Others will affect the value of K

<u>Change</u>	<u>Direction System Shifts to Reestablish Equilibrium</u>	<u>Effect on Q or K</u>
Adding a reactant	Shifts towards products	Q decreases
Adding a product	Shifts towards reactants	Q increases
Removing a reactant	Shifts towards reactants	Q increases
Removing a product	Shifts towards products	Q decreases
Increasing pressure (decreasing volume)	Shifts toward less gas molecules	Q can increase, decrease, or remain constant depending on ratio of gas molecules between reactants and products
Decreasing pressure (increasing volume)	Shifts towards more gas molecules	
Adding an inert gas	No effect	Q doesn't change
Increasing the temperature	Endothermic: shifts towards products Exothermic: shifts towards reactants	Endothermic: K increases Exothermic: K decreases
Decreasing the temperature	Endothermic: shifts towards reactants Exothermic: shifts towards products	Endothermic: K decreases Exothermic: K increases

[Video Source](#)

LO 6.10: The student is able to connect LeChatelier's principle to the comparison of Q to K by explaining the effects of the stress on Q and K .

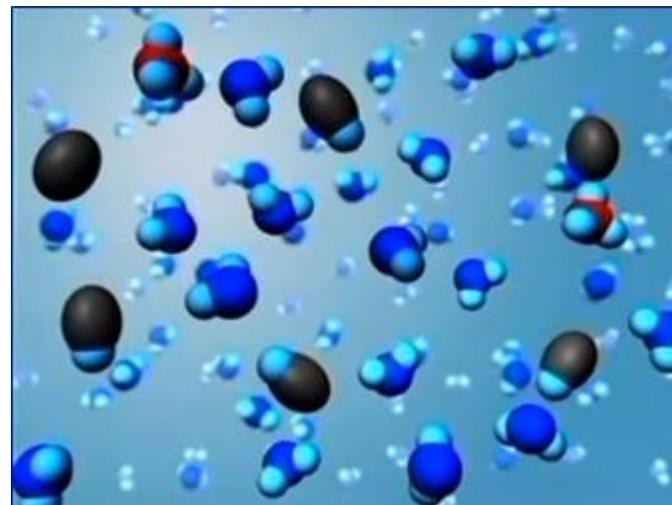
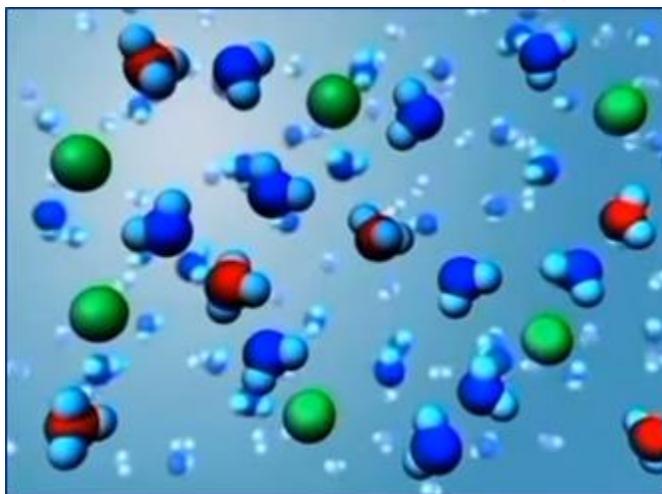
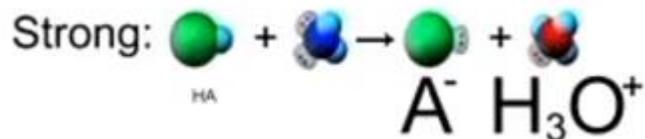
+

Acid/Base Particulates



[Source](#)

Select Acid-Base Ionization



[Video](#)

Strong: Since $K_a = \frac{[H_3O^+][A^-]}{[HA]} \gg 1$, at equilibrium strong acids are molecules that essentially ionize to *completion* in aqueous solution, disassociating into H_3O^+ ions and the additional anion.

Weak: Since $K_a = \frac{[H_3O^+][A^-]}{[HA]} \ll 1$, at equilibrium weak acids are molecules that only partially ionize in aqueous solution, disassociating into few H_3O^+ ions and the additional anion.

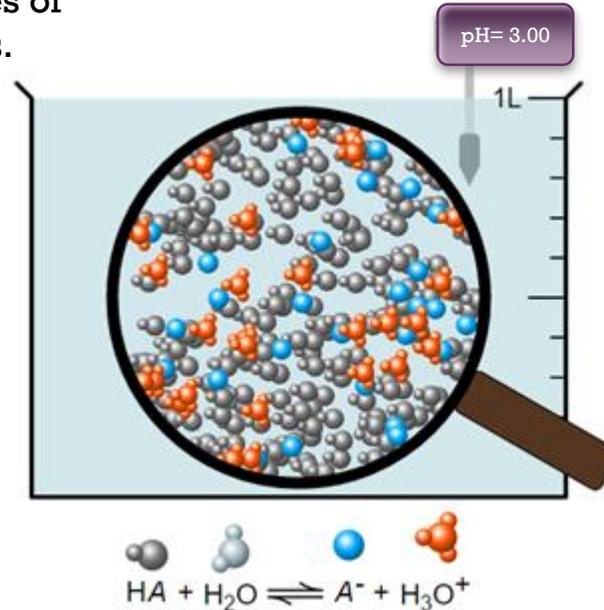
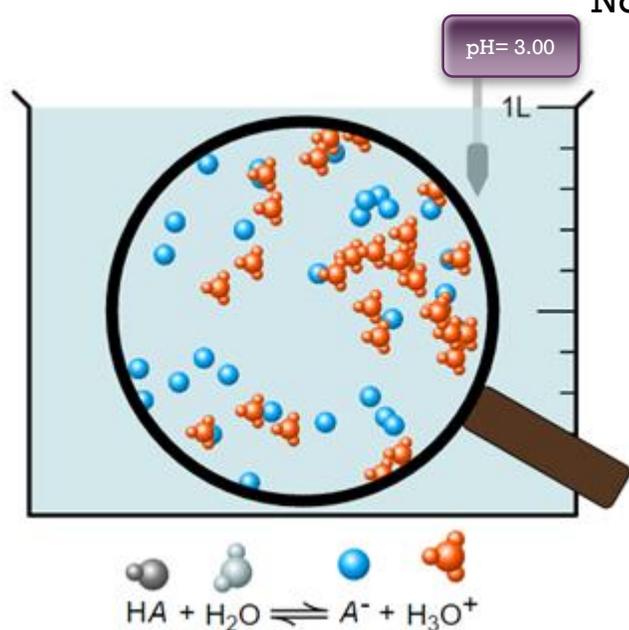
LO 6.11: The student can generate or use a particulate representation of an acid (strong or weak or polyprotic) and a strong base to explain the species that will have large versus small concentrations at equilibrium.

+

pH of Weak or Strong Acid

[Source](#)

Note the similar pH values of both monoprotic acids.



- This is a particulate picture of a strong acid whose $[HA] = 0.00100M$.
- Note the 100% ionization of this acid.

- This is a particulate picture of a weak acid whose $[HA] = 1.00M$ and $K_a = 1.00 \times 10^{-6}$.
- pH is a measure of the $[H^+]$ in solution. More moles of a weak acid are needed to achieve equivalent $[H^+]$ values of a strong acid of the same pH, since a weak acid only partially ionizes.

- If similar volumes of both acids above were titrated with the same strong base, the weak acid would require a larger volume of base to reach its equivalence point.

LO 6.12: Reason about the distinction between strong and weak acid solutions with similar values of pH, including the percent ionization of the acids, the concentrations needed to achieve the same pH, and the amount of base needed to reach the equivalence point in a titration.

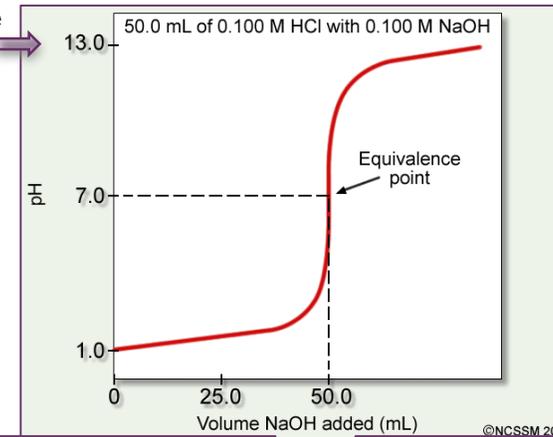
[Video](#)





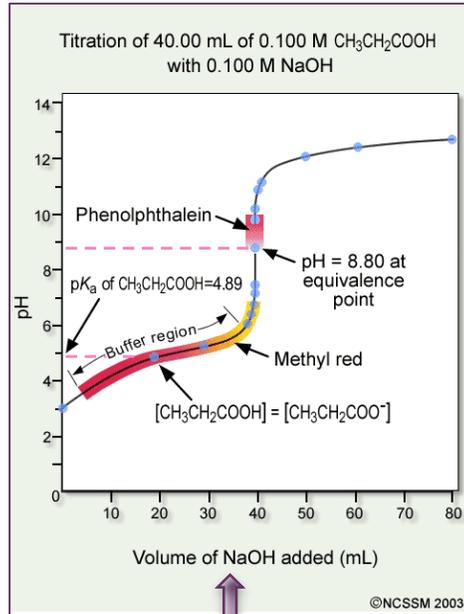
Titrations

This illustration shows the titration curve of a strong acid with a strong base.

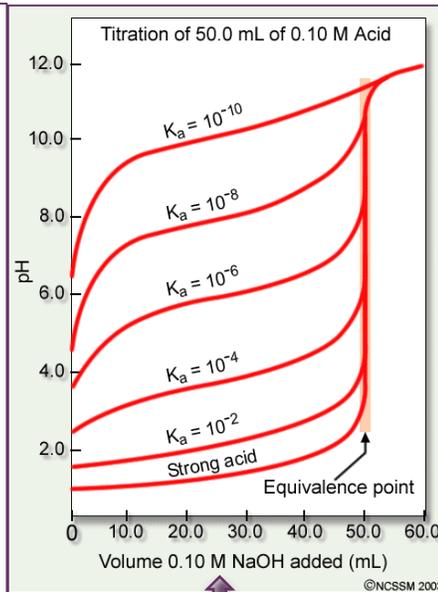


[Source](#)

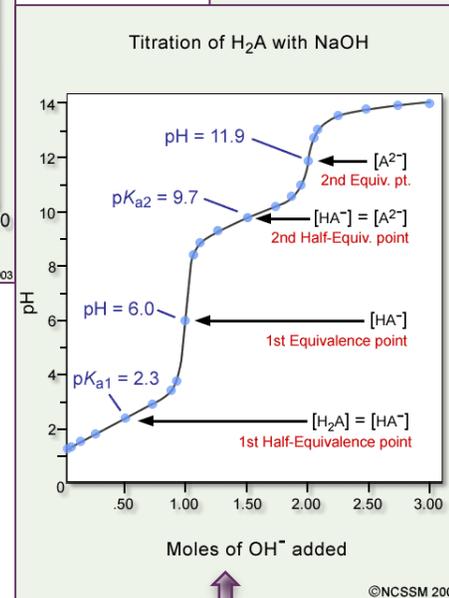
[Video](#)



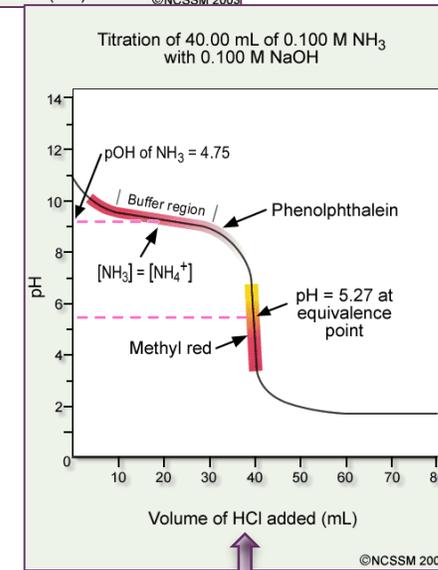
This illustration shows the titration curve of a weak acid with a strong base with indicator changes.



This illustration shows the titration curves of several weak acids with a strong base



This illustration shows the titration curve of a polyprotic weak acid with a strong base.



This illustration shows the titration curve of a weak base with a strong acid with indicator changes.

See Source link to review titration calculations.

LO 6.13: The student can interpret titration data for monoprotic or polyprotic acids involving titration of a weak or strong acid by a strong base (or a weak or strong base by a strong acid) to determine the concentration of the titrant and the pK_a for a weak acid, or the pK_b for a weak base.



8)

2015 AP[®] CHEMISTRY FREE-RESPONSE QUESTIONS

3. Potassium sorbate, $\text{KC}_6\text{H}_7\text{O}_2$ (molar mass 150. g/mol) is commonly added to diet soft drinks as a preservative. A stock solution of $\text{KC}_6\text{H}_7\text{O}_2(aq)$ of known concentration must be prepared. A student titrates 45.00 mL of the stock solution with 1.25 M $\text{HCl}(aq)$ using both an indicator and a pH meter. The value of K_a for sorbic acid, $\text{HC}_6\text{H}_7\text{O}_2$, is 1.7×10^{-5} .

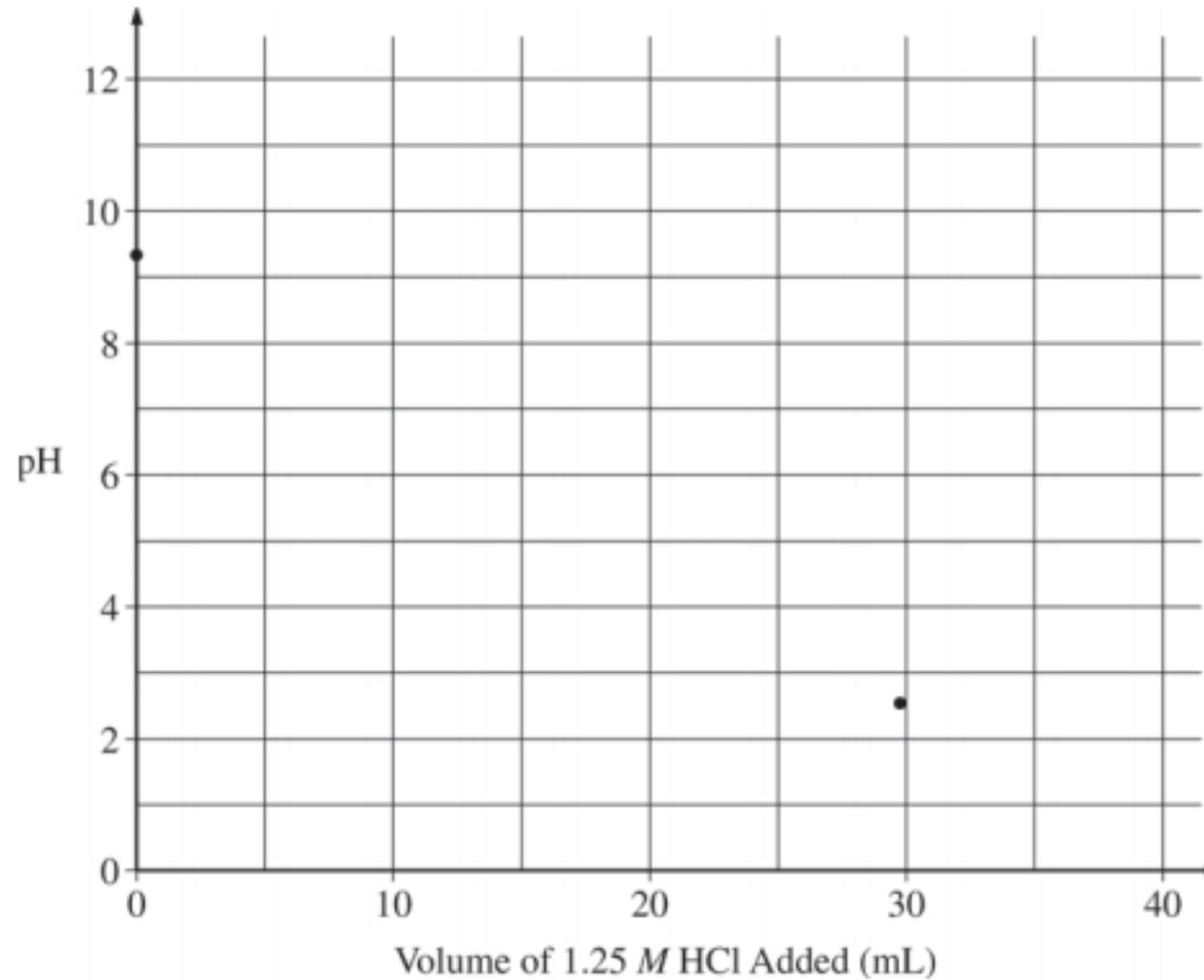
- (a) Write the net-ionic equation for the reaction between $\text{KC}_6\text{H}_7\text{O}_2(aq)$ and $\text{HCl}(aq)$.
- (b) A total of 29.95 mL of 1.25 M $\text{HCl}(aq)$ is required to reach the equivalence point. Calculate $[\text{KC}_6\text{H}_7\text{O}_2]$ in the stock solution.
- (c) The pH at the equivalence point of the titration is measured to be 2.54. Which of the following indicators would be the best choice for determining the end point of the titration? Justify your answer.

Indicator	$\text{p}K_a$
Phenolphthalein	9.3
Bromothymol blue	7.0
Methyl red	5.0
Thymol blue	2.0
Methyl violet	0.80

- (d) Calculate the pH at the half-equivalence point.



- (e) The initial pH and the equivalence point are plotted on the graph below. Accurately sketch the titration curve on the graph below. Mark the position of the half-equivalence point on the curve with an X.
- 8)





(e) The
on t

8)

(a) Write the net-ionic equation for the reaction between $\text{KC}_6\text{H}_7\text{O}_2(aq)$ and $\text{HCl}(aq)$.

$\text{H}^+ + \text{C}_6\text{H}_7\text{O}_2^- \rightleftharpoons \text{HC}_6\text{H}_7\text{O}_2$	1 point is earned the net-ionic equation.
--	---

(b) A total of 29.95 mL of 1.25 M $\text{HCl}(aq)$ is required to reach the equivalence point. Calculate $[\text{KC}_6\text{H}_7\text{O}_2]$ in the stock solution.

$\frac{1.25 \text{ mol HCl}}{1000 \text{ mL}} = \frac{x \text{ mol HCl}}{29.95 \text{ mL}} \quad x = 0.0374 \text{ mol HCl}$	1 point is earned for the moles of HCl at the equivalence point.
$\frac{0.0374 \text{ mol C}_6\text{H}_7\text{O}_2^-}{45.0 \text{ mL}} = \frac{x \text{ mol C}_6\text{H}_7\text{O}_2^-}{1000 \text{ mL}} \Rightarrow 0.832 \text{ M}$	1 point is earned for the correct answer.

(c) The pH at the equivalence point of the titration is measured to be 2.54. Which of the following indicators would be the best choice for determining the end point of the titration? Justify your answer.

Indicator	$\text{p}K_a$
Phenolphthalein	9.3
Bromothymol blue	7.0
Methyl red	5.0
Thymol blue	2.0
Methyl violet	0.80

Thymol blue; it has a $\text{p}K_a$ close to the pH at the equivalence point, so it will change color near the equivalence point.	1 point is earned for the correct indicator. 1 point is earned for correct justification.
---	--

(d) Calculate the pH at the half-equivalence point.

$\text{pH} = \text{p}K_a = -\log(1.7 \times 10^{-5}) = 4.77$	1 point is earned for the correct pH.
--	---------------------------------------

Source

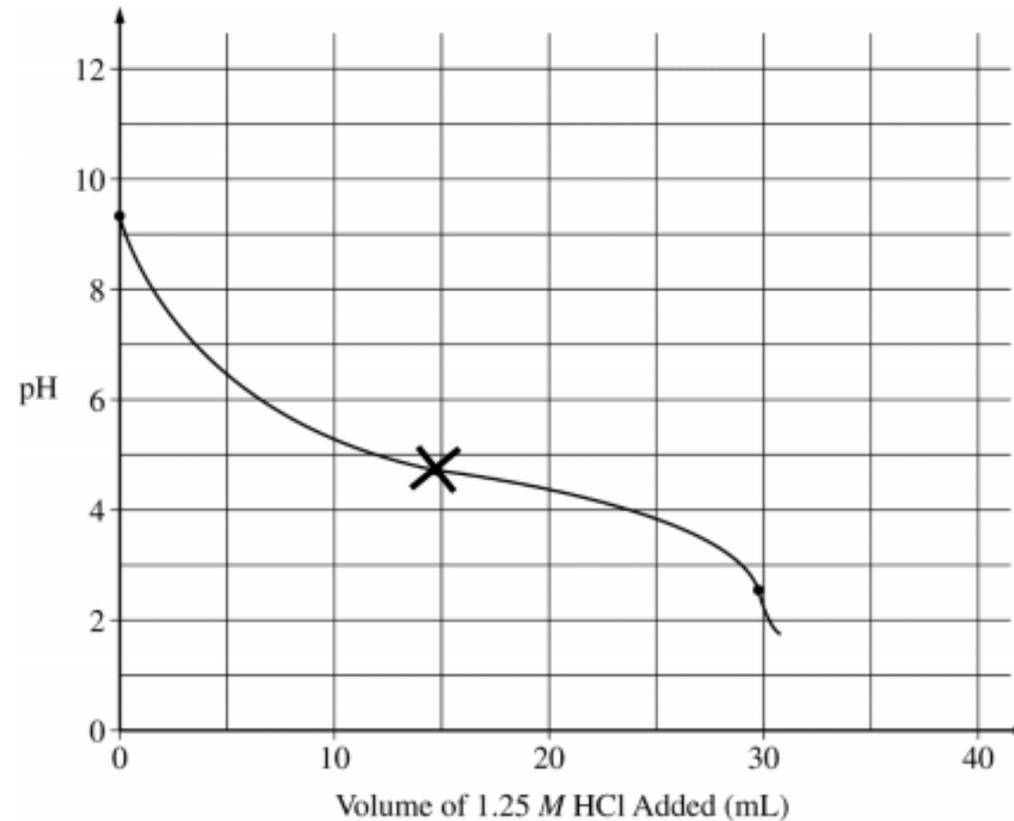
n curve





(e) The
on
8)

(e) The initial pH and the equivalence point are plotted on the graph below. Accurately sketch the titration curve on the graph below. Mark the position of the half-equivalence point on the curve with an X.



[The pH curve should have the correct shape.]

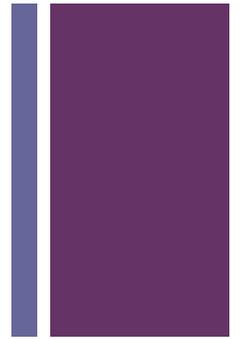
1 point is earned for a half-equivalence point consistent with the answer to part (d) and at the correct volume.

1 point is earned for a curve that levels off to a relatively horizontal slope through the half-equivalence point.

1 point is earned for a relatively steep negative slope through the equivalence point.

source

curve



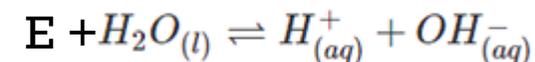
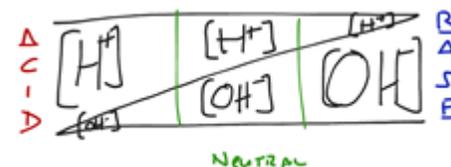
+ K_w and Temperature

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

T (°C)	K_w (mol ² dm ⁻⁶)	pH
0	0.114×10^{-14}	7.47
10	0.293×10^{-14}	7.27
20	0.681×10^{-14}	7.08
25	1.008×10^{-14}	7.00
30	1.471×10^{-14}	6.92
40	2.916×10^{-14}	6.77
50	5.476×10^{-14}	6.63
100	51.3×10^{-14}	6.14

- As T increases, pH of pure water decreases. The water is NOT becoming more acidic. A solution is only acidic if $[\text{H}^+] > [\text{OH}^-]$.
- At 50°C, the pH of pure water is 6.63, which is defined as “neutral”, when $[\text{H}^+] = [\text{OH}^-]$. A solution with a pH of 7 at this temperature is slightly basic b/c it is higher than the neutral value of 6.63.

[Video](#)



The dissociation of water is endothermic. An increase of energy will shift the reaction to the right, increasing the forward reaction, and increase the value of K_w .

LO 6.14: The student can, based on the dependence of K_w on temperature, reason that neutrality requires $[\text{H}^+] = [\text{OH}^-]$ as opposed to requiring $\text{pH} = 7$, including especially the application to biological systems.



Acid/Base Mixtures and its pH

9)

A 25 mL sample of hydrofluoric acid (HF) is titrated with 25 mL of 0.30M sodium hydroxide (NaOH). At the equivalence point of the titration, what would the pH of the solution be? Justify with a reaction.

- a. $\text{pH} < 7$
- b. $\text{pH} = 7$
- c. $\text{pH} > 7$
- d. $\text{pH} = \text{pK}_a$

[Video](#)

Click reveals answer and explanation.

LO 6.15: The student can identify a given solution as containing a mixture of strong acids and/or bases and calculate or estimate the pH (and concentrations of all chemical species) in the resulting solution.



Acid/Base Mixtures and its pH

9)

A 25 mL sample of hydrofluoric acid (HF) is titrated with 25 mL of 0.30M sodium hydroxide (NaOH). At the equivalence point of the titration, what would the pH of the solution be? Justify with a reaction.

- a. $\text{pH} < 7$
- b. $\text{pH} = 7$
- c. $\text{pH} > 7$
- d. $\text{pH} = \text{pK}_a$

[Video](#)

The correct answer is “c” $\text{pH} > 7$. At the equivalence point, the moles of acid equal the moles of base. The remaining species would be Na^+ and F^- . The conjugate base, F^- will hydrolyze with water, producing OH^- in solution: $\text{F}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HF}(aq) + \text{OH}^-(aq)$
Select the Source link to see more calculations in this titration.

LO 6.15: The student can identify a given solution as containing a mixture of strong acids and/or bases and calculate or estimate the pH (and concentrations of all chemical species) in the resulting solution.

+

10)

pH and Acid/Base Equilibria

a. Vinegar is 0.50M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, with a $K_a = 1.8 \times 10^{-5}$. What would be the pH of this solution?

1st click reveals answer and explanation.

Source link for more calculations.

[Source](#)

[Video](#)

b. Identify and compare the relative strengths of the two acids and the two bases in this neutralization reaction: $\text{OH}^-(aq) + \text{NH}_4^+(aq) \rightleftharpoons \text{H}_2\text{O}(l) + \text{NH}_3(aq)$.

2nd click reveals answer and explanation.

LO 6.16: The student can identify a given solution as being the solution of a monoprotic weak acid or base (including salts in which one ion is a weak acid or base), calculate the pH and concentration of all species in the solution, and/or infer the relative strengths of the weak acids or bases from given equilibrium concentrations.

+

10)

pH and Acid/Base Equilibria

a. Vinegar is 0.50M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, with a $K_a = 1.8 \times 10^{-5}$.

What would be the pH of this solution?

$\text{HC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$

	$\text{HC}_2\text{H}_3\text{O}_2$	H_2O	$\text{C}_2\text{H}_3\text{O}_2^-$	H_3O^+
Initial	0.5	---	0	0
Change	-x	---	+x	+x
Equilibrium	0.5 - x	---	x	x

For every acetic acid molecule that dissociates, one acetate ion and one hydronium ion is produced. This can be represented by subtracting "x" from the original acetic acid concentration, and adding "x" to the original concentrations of the dissociated ions.

You can create a modified equilibrium constant expression

$$K_a = \frac{[\text{C}_2\text{H}_3\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

and then plug in the concentration values you found in the ICE table

$$1.8 \times 10^{-5} = \frac{x^2}{0.5 - x}$$

so

$$x^2 + (1.8 \times 10^{-5})x - (9 \times 10^{-6}) = 0$$

then use the quadratic formula to calculate

$$x = 0.0030 \text{ M} = [\text{H}_3\text{O}^+]$$

which can be plugged into the formula

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$-\log(0.0030) = \text{pH} = 2.5$$

See Source link for more calculations.

[Source](#)

[Video](#)

b. Identify and compare the relative strengths of the two acids and the two bases in this neutralization reaction: $\text{OH}^-(\text{aq}) + \text{NH}_4^+(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\text{l}) + \text{NH}_3(\text{aq})$.

Answer: Hydroxide(OH^-), and ammonia(NH_3) are the two bases, with ammonia being the weaker, while ammonium(NH_4^+) and water are the acids, with water the weaker.

The correct formulation of acid strength is based on a comparison to the K_a 's of hydronium (=1) and water (= 1×10^{-14}). Any acid that is stronger than hydronium (e.g. HCl) is a strong acid and its conjugate base (chloride) has no effect on pH because it does not hydrolyze in water. Any acid whose strength is between hydronium and water (e.g. ammonium) is a weak acid, and its conjugate is a weak base. Similarly, for bases, any base with K_b greater than that of hydroxide, which is = 1, (e.g., oxide ion as K_2O) is a strong base, and its conjugate acid (K^+) has no effect on pH because it does not hydrolyze in water. Any base with K_b between that of hydroxide and water (= 1×10^{-14}) is a weak base, and its conjugate is a weak acid (e.g., acetate).

LO 6.16: The student can identify a given solution as being the solution of a monoprotic weak acid or base (including salts in which one ion is a weak acid or base), calculate the pH and concentration of all species in the solution, and/or infer the relative strengths of the weak acids or bases from given equilibrium concentrations.

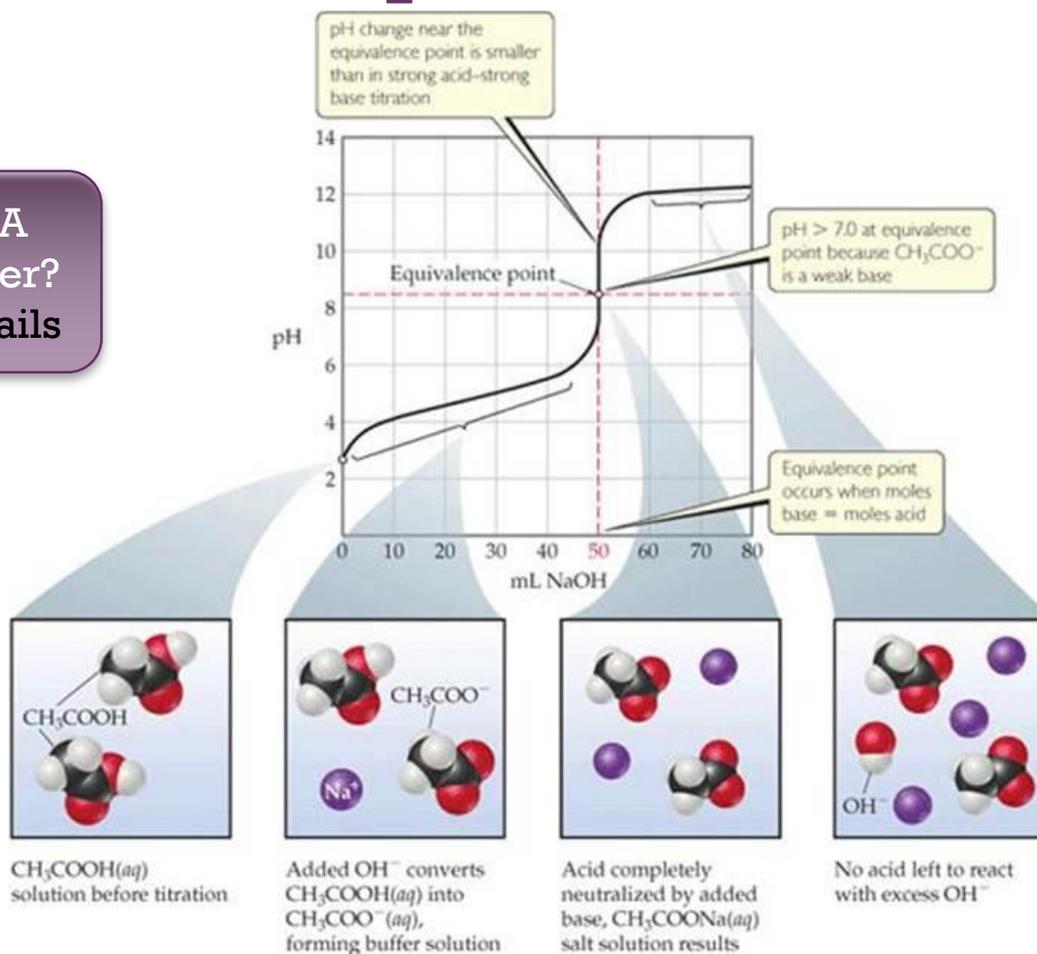
+

Acid/Base reaction species

Ask yourself:

Is it strong? Weak? A salt? A buffer?
What will it do in water?
See Source link for more details

Deal w/strong A/B first. These will react to completion with the available species.



Video

LO 6.17: The student can, given an arbitrary mixture of weak and strong acids and bases (including polyprotic systems), determine which species will react strongly with one another (i.e., with $K > 1$) and what species will be present in large concentrations at equilibrium.



How to Build a Buffer:

Getting the pH correct:

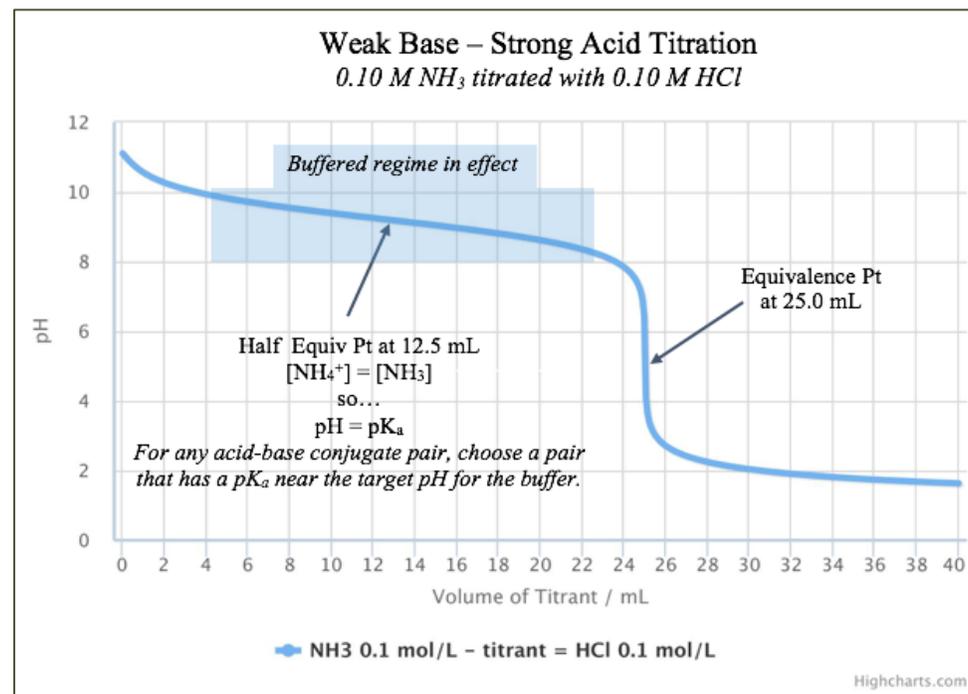
- The pH of a buffer is primarily determined by the pK_a of the weak acid in the conjugate acid-base pair.
- When both species in the conjugate acid-base pair have equal concentrations, the pH of the buffer is equal to the pK_a.
- Choose a conjugate acid-base pair that has a pK_a closest to the pH you desire and then adjust concentrations to fine tune from there.

Estimating Buffer Capacity:

- A buffer is only effective as long as it has sufficient amounts of both members of the conjugate acid-base pair to allow equilibrium to shift during a stress.

[Source](#)

[Video](#)



LO 6.18: The student can design a buffer solution with a target pH and buffer capacity by selecting an appropriate conjugate acid-base pair and estimating the concentrations needed to achieve the desired capacity.



How to Build a Buffer:

Source

Estimating Buffer Capacity:

- A buffer is only effective as long as it has sufficient amounts of both members of the conjugate acid-base pair to allow

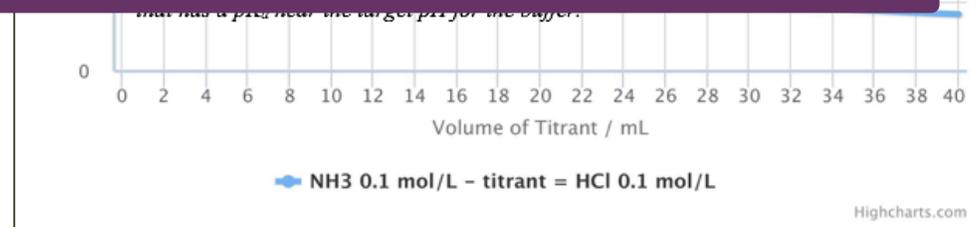
Getting the pH correct:

1 1)

A student must create a buffer with a pH of 6.2 from an acid, HA ($pK_a = 6.5$). They need to _____.

- add a strong base to the solution
- add a strong acid to the solution
- make sure the solution has $[A^-] > [HA]$
- make sure the solution has $[A^-] < [HA]$

pH you desire and then adjust concentrations to fine tune from there.



LO 6.18: The student can design a buffer solution with a target pH and buffer capacity by selecting an appropriate conjugate acid-base pair and estimating the concentrations needed to achieve the desired capacity.



How to Build a Buffer:

[Source](#)

Estimating Buffer Capacity:

- A buffer is only effective as long as it has sufficient amounts of both members of the conjugate acid-base pair to allow

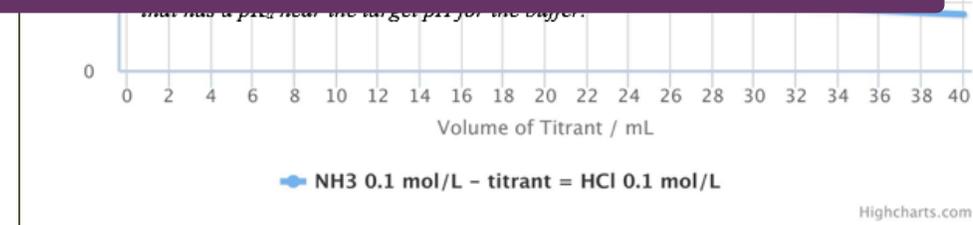
Getting the pH correct:

11)

A student must create a buffer with a pH of 6.2 from an acid, HA ($pK_a = 6.5$). They need to _____.

- a) add a strong base to the solution
- b) add a strong acid to the solution
- c) make sure the solution has $[A^-] > [HA]$
- d) make sure the solution has $[A^-] < [HA]$

pH you desire and then adjust concentrations to fine tune from there.



LO 6.18: The student can design a buffer solution with a target pH and buffer capacity by selecting an appropriate conjugate acid-base pair and estimating the concentrations needed to achieve the desired capacity.



Finding the Major Species

12)

- A 50.0 mL sample of 0.50 M $\text{HC}_2\text{H}_3\text{O}_2$ is titrated to the half equivalence point with 25.0 mL of 0.50 M NaOH. Which of the following options shows the correct ranking of the molarities of the species in solution?

(pKa for $\text{HC}_2\text{H}_3\text{O}_2$ is 4.7)

- $[\text{HC}_2\text{H}_3\text{O}_2] > [\text{C}_2\text{H}_3\text{O}_2^{1-}] > [\text{H}^+] > [\text{OH}^-]$
- $[\text{HC}_2\text{H}_3\text{O}_2] = [\text{C}_2\text{H}_3\text{O}_2^{1-}] > [\text{H}^+] > [\text{OH}^-]$
- $[\text{HC}_2\text{H}_3\text{O}_2] > [\text{C}_2\text{H}_3\text{O}_2^{1-}] = [\text{H}^+] > [\text{OH}^-]$
- $[\text{C}_2\text{H}_3\text{O}_2^{1-}] > [\text{HC}_2\text{H}_3\text{O}_2] > [\text{OH}^-] > [\text{H}^+]$

Option B is correct.

At the half equivalence point, the concentrations of the weak acid and its conjugate base are equal because the OH^- ion has reacted with half of the original acetic acid. The pH of the buffer is equal to the pKa at that point, so the $[\text{H}^+]$ is $10^{-4.7}$ which is far lower than the concentrations of the conjugates but far higher than the $[\text{OH}^-]$ which has a value of $10^{-9.3}$.

LO 6.19: The student can relate the predominant form of a chemical species involving a labile proton (i.e., protonated/deprotonated form of a weak acid) to the pH of a solution and the pKa associated with the labile proton.

[Video](#)



Finding the Major Species

[Source](#)



12)

- A 50.0 mL sample of 0.50 M $\text{HC}_2\text{H}_3\text{O}_2$ is titrated to the half equivalence point with 25.0 mL of 0.50 M NaOH. Which of the following options shows the correct ranking of the molarities of the species in solution?

(pKa for $\text{HC}_2\text{H}_3\text{O}_2$ is 4.7)

- a. $[\text{HC}_2\text{H}_3\text{O}_2] > [\text{C}_2\text{H}_3\text{O}_2^{1-}] > [\text{H}^+] > [\text{OH}^-]$
- b. $[\text{HC}_2\text{H}_3\text{O}_2] = [\text{C}_2\text{H}_3\text{O}_2^{1-}] > [\text{H}^+] > [\text{OH}^-]$
- c. $[\text{HC}_2\text{H}_3\text{O}_2] > [\text{C}_2\text{H}_3\text{O}_2^{1-}] = [\text{H}^+] > [\text{OH}^-]$
- d. $[\text{C}_2\text{H}_3\text{O}_2^{1-}] > [\text{HC}_2\text{H}_3\text{O}_2] > [\text{OH}^-] > [\text{H}^+]$

[Video](#)

Click reveals answer and explanation.

LO 6.19: The student can relate the predominant form of a chemical species involving a labile proton (i.e., protonated/deprotonated form of a weak acid) to the pH of a solution and the pKa associated with the labile proton.



The Buffer Mechanism

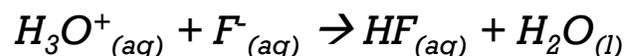
[Source](#)



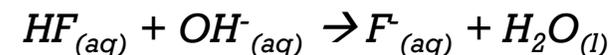
A buffer is able to resist pH change because the two components (conjugate acid and conjugate base) are both present in appreciable amounts at equilibrium and are able to neutralize small amounts of other acids and bases (in the form of H_3O^+ and OH^-) when they are added to the solution. Take, for example, a fluoride buffer made from hydrofluoric acid and NaF. An ideal fluoride buffer would contain equimolar concentrations of HF and NaF. Since they are a weak acid and a weak base, respectively, the amount of hydrolysis is minimal and both buffer species are present at, effectively, their initial supplied concentrations.

[Video](#)

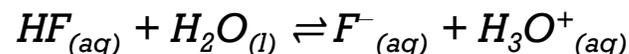
If a strong acid is added to the HF/ F^- buffer, then the added acid will react completely with the available base, F^- . This results in a nearly unchanged $[\text{H}_3\text{O}^+]$ and a nearly unchanged pH.



If a strong base is added to the HF/ F^- buffer, then the added base will react completely with the available acid, HF. This results in a nearly unchanged $[\text{H}_3\text{O}^+]$ and a nearly unchanged pH.



The slight shift in pH after challenge is governed by the hydrolysis equilibrium of HF, based on the new HF and F^- concentrations:



LO 6.20: The student can identify a solution as being a buffer solution and explain the buffer mechanism in terms of the reactions that would occur on addition of acid/base.

+ The Buffer Mechanism

A buffer is able to resist pH change because the two components (conjugate acid and conjugate base) are both present in appreciable amounts at equilibrium and are able to neutralize small amounts of other acids and bases (in the form of H_3O^+ and OH^-) when they are added to the solution. Take, for example, a fluoride buffer made from

equimolar amounts of weak acid and its conjugate base, respectively. Both are present

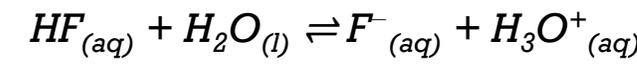
If a strong acid is added to a buffer, the conjugate base will react completely with the added acid. This results in a decrease in $[H_3O^+]$ and a corresponding increase in H_3O^+

13) A weak acid, HF, is in solution with dissolved sodium fluoride, NaF. If HCl is added, which ion will react with the extra hydrogen ions from the HCl to keep the pH from changing?

- a) OH^-
- b) Na^+
- c) F^-
- d) Na^-

weak acid, HF, will react with the added H_3O^+ to form $H_2O(l)$ and F^- . The pH will increase.

The slight shift in pH after challenge is governed by the hydrolysis equilibrium of HF, based on the new HF and F^- concentrations:



LO 6.20: The student can identify a solution as being a buffer solution and explain the buffer mechanism in terms of the reactions that would occur on addition of acid/base.



The Buffer Mechanism

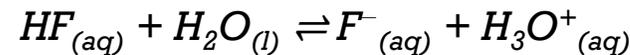
A buffer is able to resist pH change because the two components (conjugate acid and conjugate base) are both present in appreciable amounts at equilibrium and are able to neutralize small amounts of other acids and bases (in the form of H_3O^+ and OH^-) when they are added to the solution. Take, for example, a fluoride buffer made from a weak acid and its conjugate base, respectively.

If a strong acid is added to a buffer, the conjugate base will react with the H_3O^+ and form the weak acid. This results in a decrease in $[H_3O^+]$ and a slight increase in pH.

13) A weak acid, HF, is in solution with dissolved sodium fluoride, NaF. If HCl is added, which ion will react with the extra hydrogen ions from the HCl to keep the pH from changing?

- a) OH^-
- b) Na^+
- c) F^-
- d) Na^-

The slight shift in pH after challenge is governed by the hydrolysis equilibrium of HF, based on the new HF and F^- concentrations:



LO 6.20: The student can identify a solution as being a buffer solution and explain the buffer mechanism in terms of the reactions that would occur on addition of acid/base.

+

K_{sp} and Solubility Calculations

[Source](#)



14)

Question:

What is the maximum number of moles of AgBr that will fully dissolve in 1.0 L of water, if the K_{sp} value of silver bromide is 4.0×10^{-12} ?

[Video](#)

4.0×10^{-12}

b. 2.0×10^{-12}

c. 4.0×10^{-6}

d. 2.0×10^{-6}

Click reveals answer and explanation.

LO 6.21: The student can predict the solubility of a salt, or rank the solubility of salts, given the relevant K_{sp} values.

+

K_{sp} and Solubility Calculations

14)

Question:

What is the maximum number of moles of AgBr that will fully dissolve in 1.0 L of water, if the K_{sp} value of silver bromide is 4.0×10^{-12} ?

a. 4.0×10^{-12}

b. 2.0×10^{-12}

c. 4.0×10^{-6}

d. 2.0×10^{-6}

[Video](#)

Answer:

The correct answer is “d” 2.0×10^{-6} . To determine the solubility of AgBr we need to determine the maximum concentrations of Ag⁺ and Br⁻ that will equal the equilibrium constant. The K_{sp} equation is $K_{sp} = [Ag^+][Br^-]$ with $[Ag^+] = [Br^-] = X$, so to solve for X we need to take the square root of 4.0×10^{-12}

LO 6.21: The student can predict the solubility of a salt, or rank the solubility of salts, given the relevant K_{sp} values.

+

Find a K_{sp} from solubility data

[Source](#)

- 15) Using a conductivity meter and a pH probe, a student determined the molar solubility of saturated solutions made by dissolving hydroxide salts in water. The student's data is shown in the table below:

Compound	Solubility (mol/L)
Ba(OH) ₂	0.11 M
Ca(OH) ₂	0.011 M
Cr(OH) ₃	1.3 x 10 ⁻⁸ M
Sn(OH) ₂	1.1 x 10 ⁻⁹ M

[Video](#)

Which of the hydroxide salts above has the smallest K_{sp} value?

- a. Ba(OH)₂ b. Ca(OH)₂ c. Cr(OH)₃ d. Sn(OH)₂

Click reveals answer and explanation.

LO 6.22: The student can interpret data regarding solubility of salts to determine, or rank, the relevant K_{sp} values.

+

Find a K_{sp} from solubility data

[Source](#)

15)

Using a conductivity meter and a pH probe, a student determined the molar solubility of saturated solutions made by dissolving hydroxide salts in water. The student's data is shown in the table below:

Compound	Solubility (mol/L)
Ba(OH) ₂	0.11 M
Ca(OH) ₂	0.011 M
Cr(OH) ₃	1.3 x 10 ⁻⁸ M
Sn(OH) ₂	1.1 x 10 ⁻⁹ M

[Video](#)

Which of the hydroxide salts above has the smallest K_{sp} value?

- a. Ba(OH)₂ b. Ca(OH)₂ c. Cr(OH)₃ d. Sn(OH)₂

Option C is correct.

Ba(OH)₂, Ca(OH)₂ and Sn(OH)₂ all have the same K_{sp} expression $K = [X][2X]^2$ or $K_{sp} = 4[X]^3$
So the salt with the lowest solubility will have the smallest K_{sp} .

$$K_{sp} \text{ for Sn(OH)}_2 = (4)(1.1 \times 10^{-9})^3 = 5.4 \times 10^{-27}$$

Cr(OH)₃'s K_{sp} expression is different. It is $K_{sp} = [X][3X]^3 = 27[X]^4$

$$K_{sp} \text{ for Cr(OH)}_3 = (27)(1.3 \times 10^{-8})^4 = 6.7 \times 10^{-31}$$

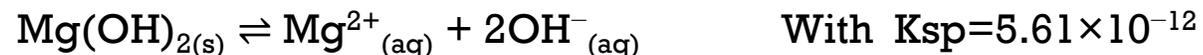
Solving without a calculator... Ba, Ca and Sn all have the same K_{sp} expression, so the smallest solubility will have the smallest K_{sp} .

The contest is now between Sn(OH)₂ and Cr(OH)₃. You need to quickly establish that since $(10^{-8})^4 = 10^{-32}$ and $(10^{-9})^3 = 10^{-27}$ the K_{sp} of Cr(OH)₃ will be smaller than that of Sn(OH)₂.

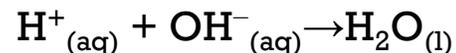
LO 6.22: The student can interpret data regarding solubility of salts to determine, or rank, the relevant K_{sp} values.

+ Common Ion Effect

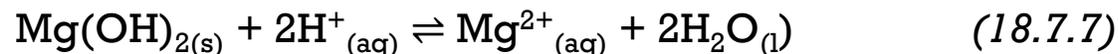
The solubility of a sparingly soluble hydroxide can be greatly increased by the addition of an acid. For example, the hydroxide salt $\text{Mg}(\text{OH})_2$ is relatively insoluble in water:



When acid is added to a saturated solution that contains excess solid $\text{Mg}(\text{OH})_2$, the following reaction occurs, removing OH^- from solution:



The overall equation for the reaction of $\text{Mg}(\text{OH})_2$ with acid is thus



As more acid is added to a suspension of $\text{Mg}(\text{OH})_2$, the equilibrium shown in Equation [18.7.7](#) is driven to the right, so more $\text{Mg}(\text{OH})_2$ dissolves.

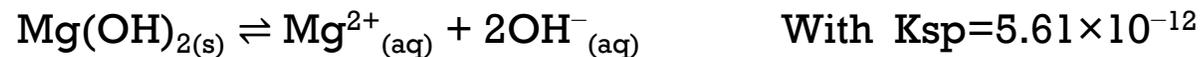
In contrast, the solubility of a sparingly soluble salt may be decreased greatly by the addition of a common ion. For example, if MgCl_2 is added to a saturated $\text{Mg}(\text{OH})_2$ solution, additional $\text{Mg}(\text{OH})_2$ will precipitate out. The additional Mg^{2+} ions will shift the original equilibrium to the left, thus reducing the solubility of the magnesium hydroxide.

LO 6.23: The student can interpret data regarding the relative solubility of salts in terms of factors (common ions, pH) that influence the solubility.



+ Common Ion Effect

The solubility of a sparingly soluble hydroxide can be greatly increased by the addition of an acid. For example, the hydroxide salt $\text{Mg}(\text{OH})_2$ is relatively insoluble in water:



When acid is added to the following

The overall

As more acid is added, Equation 18

In contrast

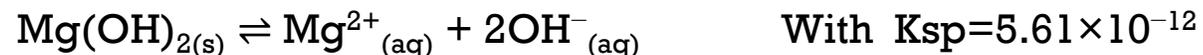
by the addition of a common ion. For example, if MgCl_2 is added to a saturated $\text{Mg}(\text{OH})_2$ solution, additional $\text{Mg}(\text{OH})_2$ will precipitate out. The additional Mg^{2+} ions will shift the original equilibrium to the left, thus reducing the solubility of the magnesium hydroxide.

16) For a solution equimolar in HCN and NaCN, which statement is false?

- a) This is an example of the common ion effect
- b) $[\text{H}^+]$ is larger than it would be if only the HCN was in solution
- c) $[\text{H}^+]$ is less than $[\text{CN}^-]$
- d) Additional NaOH will increase $[\text{CN}^-]$ and decrease $[\text{HCN}]$

+ Common Ion Effect

The solubility of a sparingly soluble hydroxide can be greatly increased by the addition of an acid. For example, the hydroxide salt $\text{Mg}(\text{OH})_2$ is relatively insoluble in water:



When acid is added to the following

The overall

As more acid is added, Equation 18

In contrast

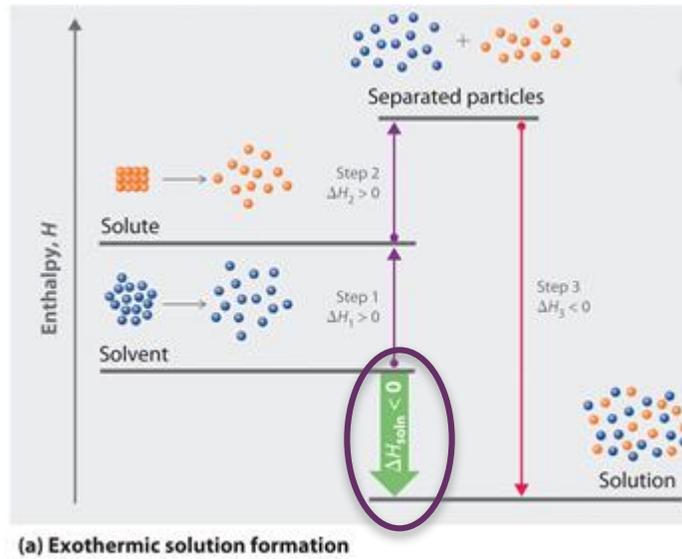
by the addition of a common ion. For example, if MgCl_2 is added to a saturated $\text{Mg}(\text{OH})_2$ solution, additional $\text{Mg}(\text{OH})_2$ will precipitate out. The additional Mg^{2+} ions will shift the original equilibrium to the left, thus reducing the solubility of the magnesium hydroxide.

16) For a solution equimolar in HCN and NaCN, which statement is false?

- a) This is an example of the common ion effect
- b) $[\text{H}^+]$ is larger than it would be if only the HCN was in solution
- c) $[\text{H}^+]$ is less than $[\text{CN}^-]$
- d) Additional NaOH will increase $[\text{CN}^-]$ and decrease $[\text{HCN}]$

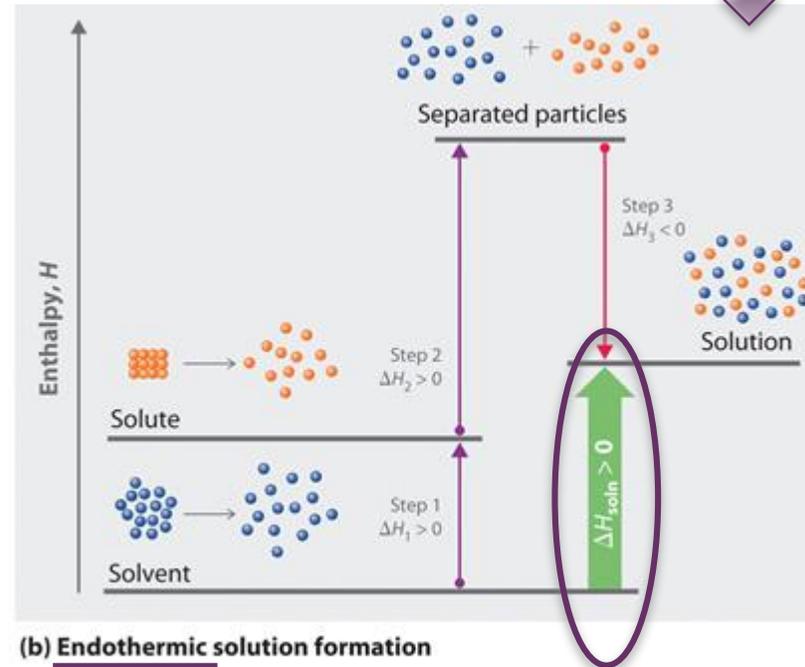
+

Salt dissolution: ΔH and ΔS

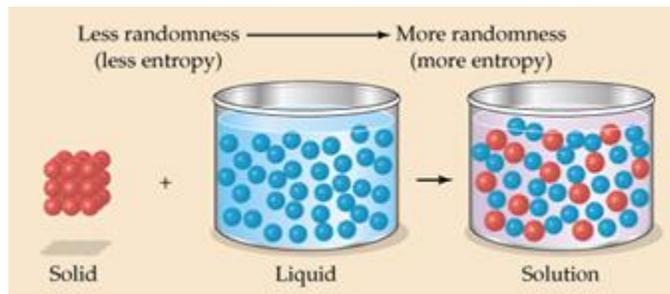


(a) Exothermic solution formation

- The enthalpy (ΔH_{soln}) of dissolution is dependent upon the intermolecular forces of the solute and solvent.



(b) Endothermic solution formation



- The entropy (ΔS_{soln}) of dissolution generally increases the disorder of the system.

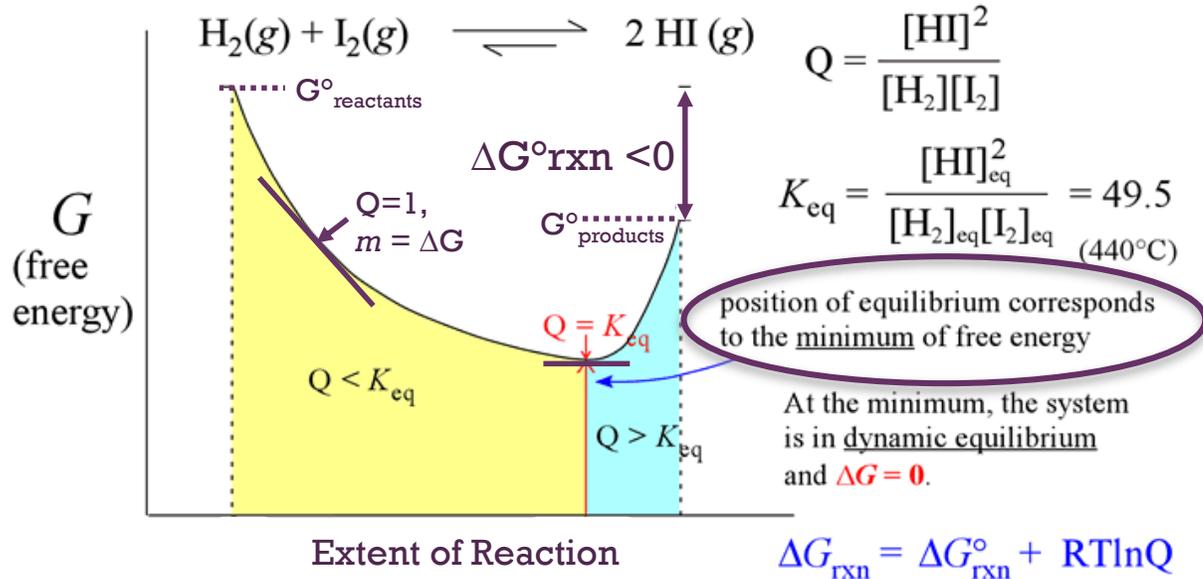
[Video](#)
[Java Tutorial](#)

LO 6.24: The student can analyze the enthalpic and entropic changes associated with the dissolution of a salt, using particulate level interactions and representations.

+

K, ΔG° and thermodynamic favorability

[Source](#)

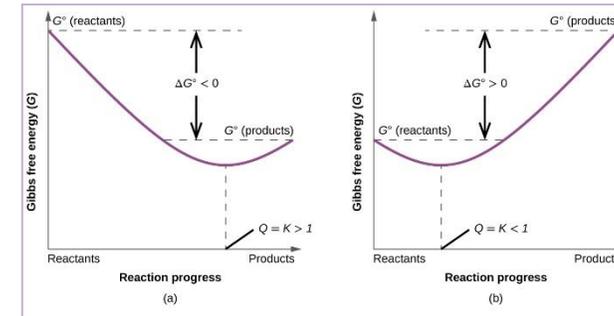


$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= -RT \ln K$$

$$= -nFE^\circ$$

[Video](#)



The key to understanding the relationship between ΔG° and K is recognizing that the magnitude of ΔG° tells us how far the standard-state is from equilibrium. The smaller the value of ΔG° , the closer the standard-state is to equilibrium. The larger the value of ΔG° , the further the reaction has to go to reach equilibrium.

LO 6.25: The student is able to express the equilibrium constant in terms of ΔG° and RT and use this relationship to estimate the magnitude of K and, consequently, the thermodynamic favorability of the process.