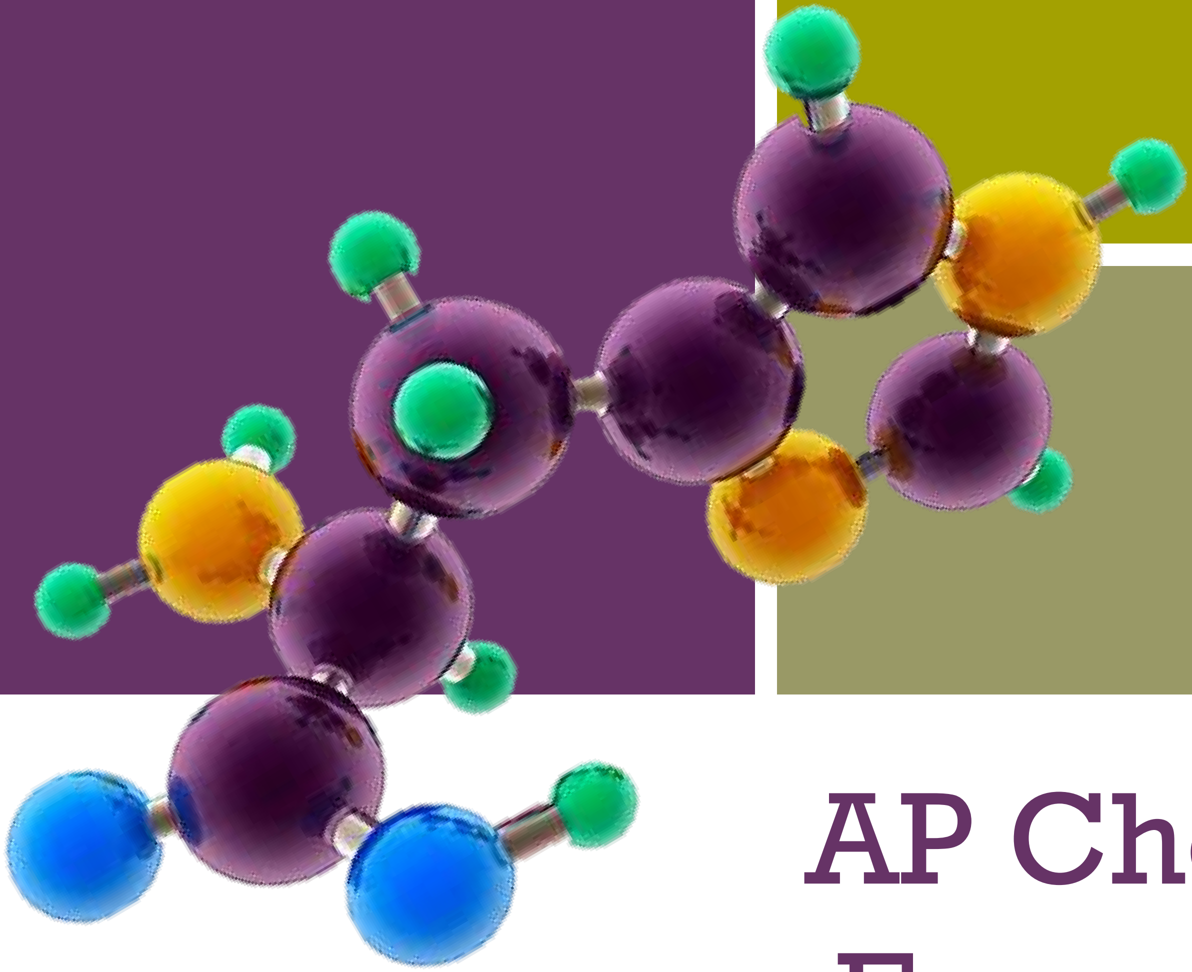
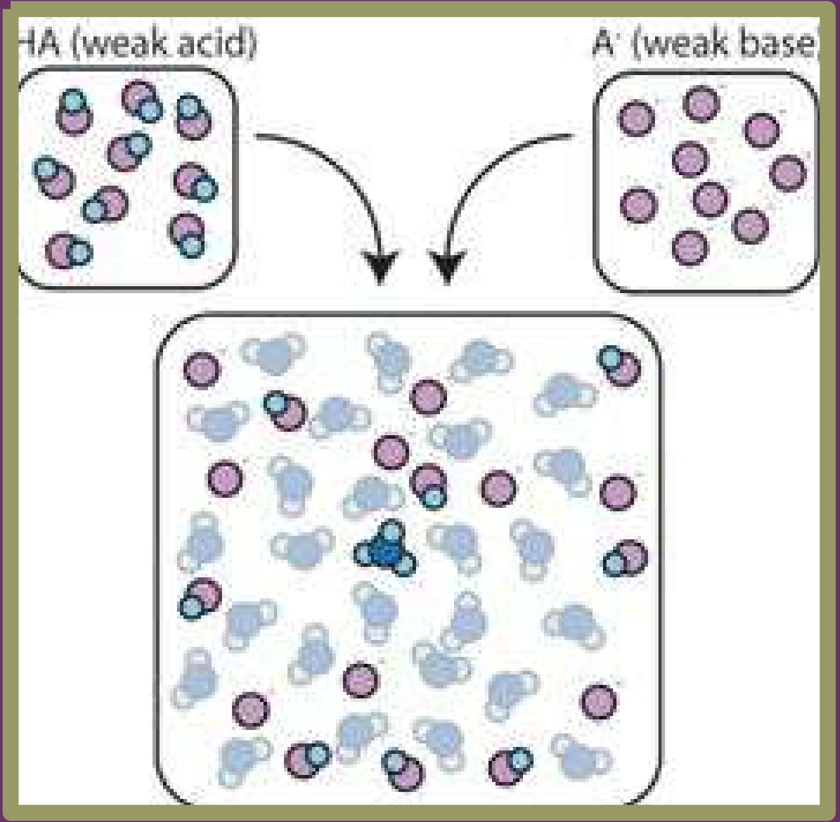


AP Chemistry Exam Review



+

Big Idea #6

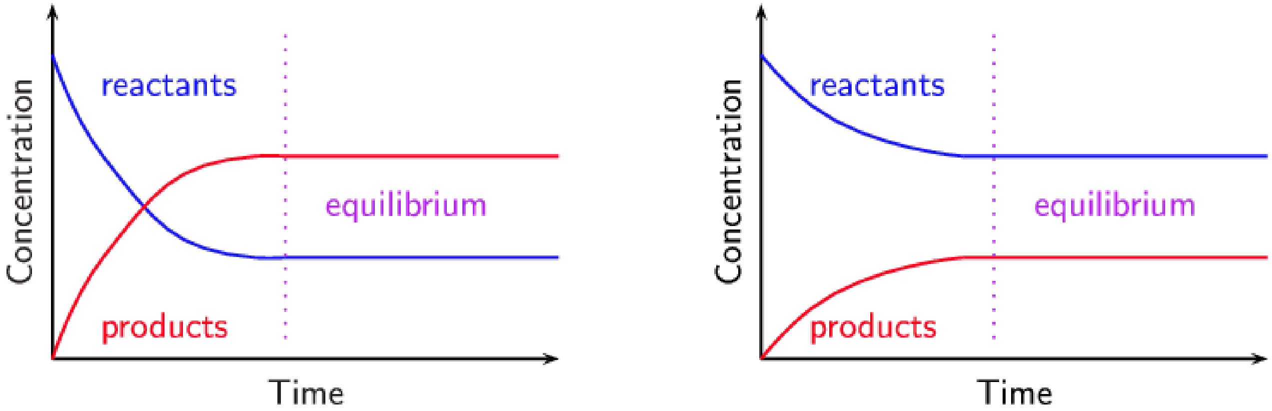
Equilibrium

Source

+ What is chemical equilibrium?

Video

- 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.



In the reaction:

$$aA + bB \rightleftharpoons cC + dD$$

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

[A] = concentration of A in $moldm^{-3}$

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.

Source

Video

+ 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.

$$P_4\,(s) + 6\,Cl_2\,(g) \leftrightarrows 4\,PCl_3\,(l)$$

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

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

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g) \qquad K_1 = 1.6 \times 10^{-11}$$

$$2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g) \qquad K_2 = 1.3 \times 10^{-10}$$

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g) \qquad K_3 = ?$$

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

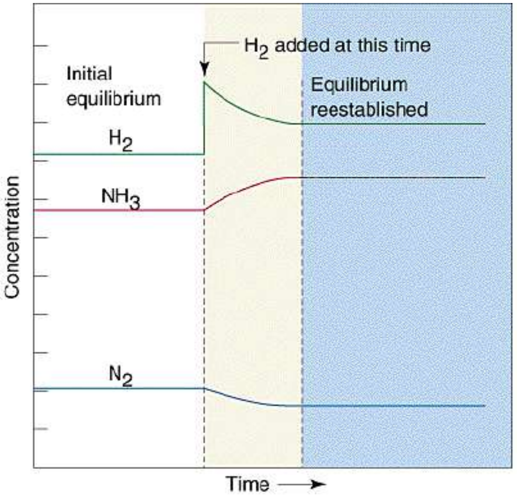
LO 6.2: The student can, given a manipulation of a chemical reaction or set of reactions (e.g., reversal of reaction or addition of two reactions), determine the effects of that manipulation on Q or K.

Source

+ Kinetics and Equilibrium

Video

- 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

Source

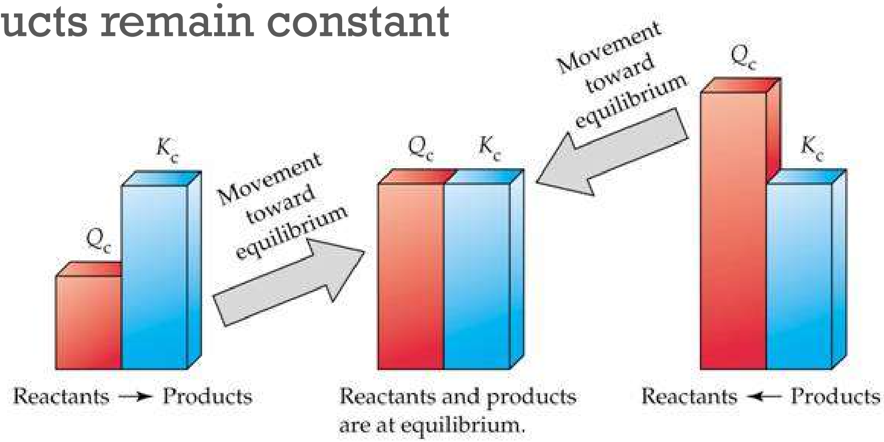
+ Q vs. K

Video

- Equilibrium is reacted 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

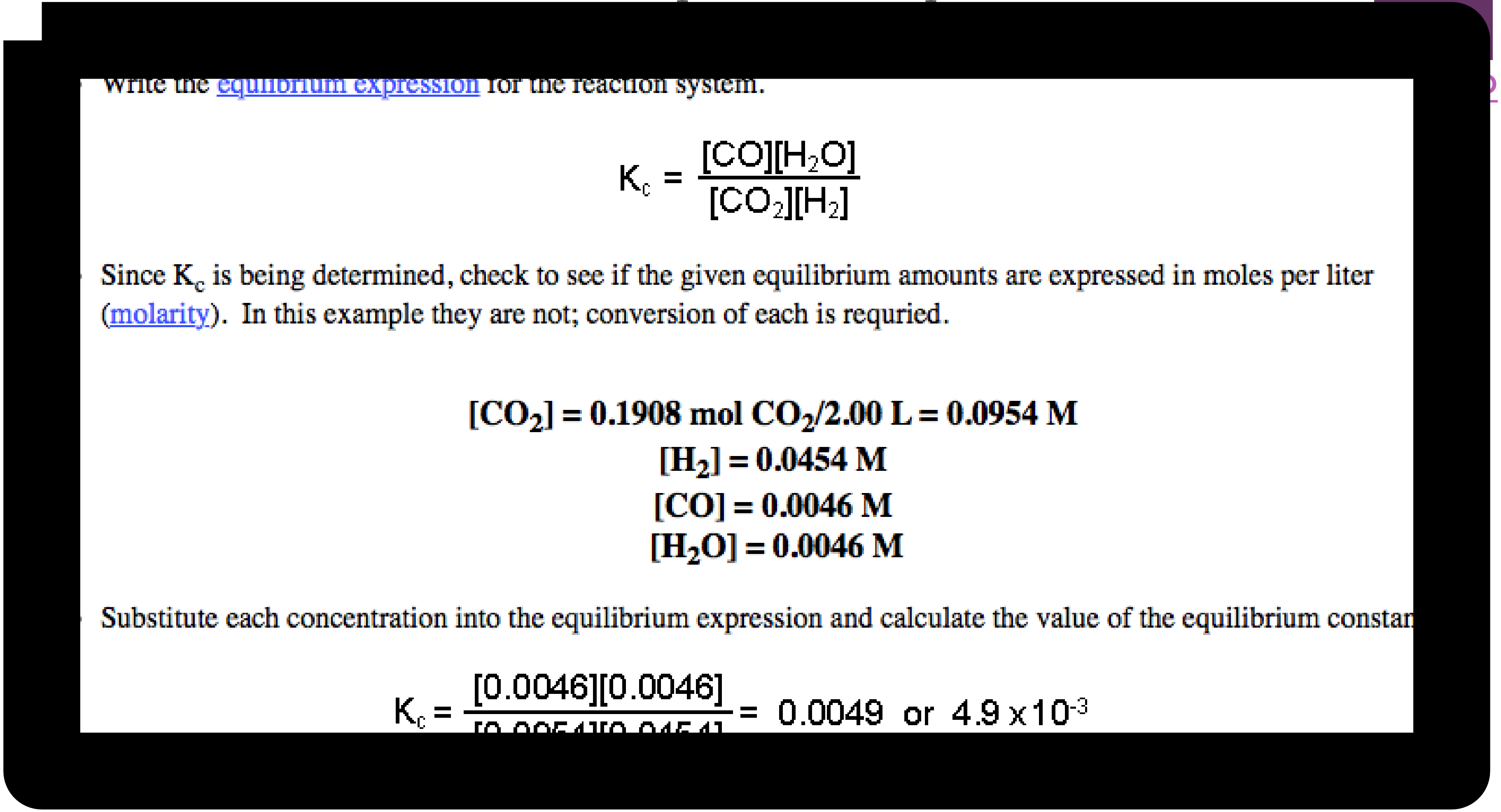


LO 6.4: Given a set of initial conditions and the equilibrium constant, K, use the tendency of Q to approach K to predict and justify the prediction as to whether the reaction will proceed toward products or reactants as equilibrium is approached.

Source Source

+ Calculating *K*

- Equilibrium constants can be determined using experimental concentrations of reactants and products at equilibrium.



- Write the equilibrium expression for the reaction system.

$$K_c = \frac{[CO][H_2O]}{[CO_2][H_2]}$$

- 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 requried.

$$[CO_2] = 0.1908 \text{ mol } CO_2/2.00 \text{ L} = 0.0954 \text{ M}$$
$$[H_2] = 0.0454 \text{ M}$$
$$[CO] = 0.0046 \text{ M}$$
$$[H_2O] = 0.0046 \text{ M}$$

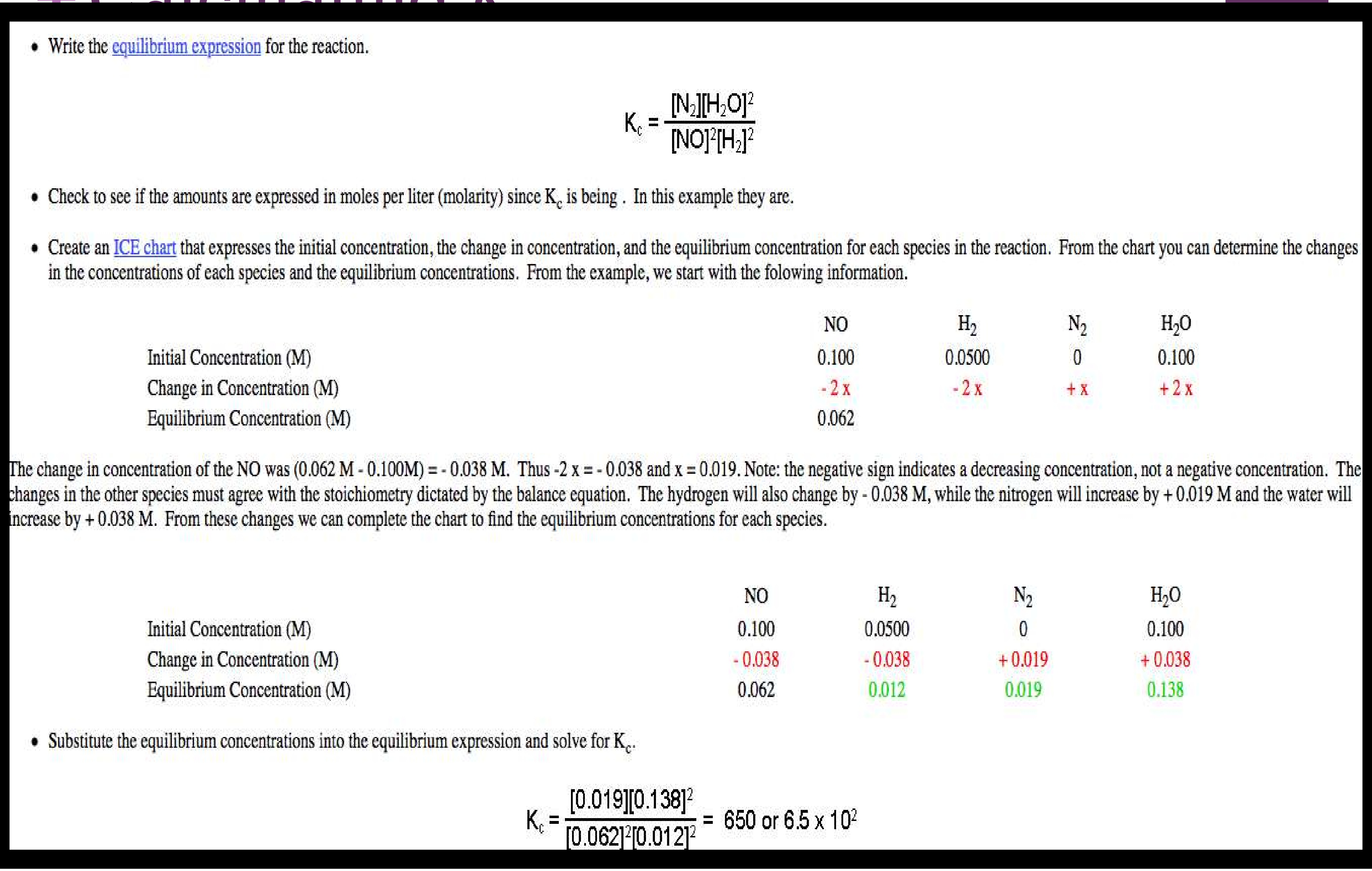
- Substitute each concentration into the equilibrium expression and calculate the value of the equilibrium constan

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

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.

Source Source

Calculating *K*



- Write the equilibrium expression for the reaction.

$$K_c = \frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2}$$

- 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 folowing information.

	NO	H_2	N_2	H_2O
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 M - 0.100M) = - 0.038 M. Thus -2 x = - 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 M and the water will increase by + 0.038 M. From these changes we can complete the chart to find the equilibrium concentrations for each species.

	NO	H_2	N_2	H_2O
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

Video

- 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

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

Click reveals answer.

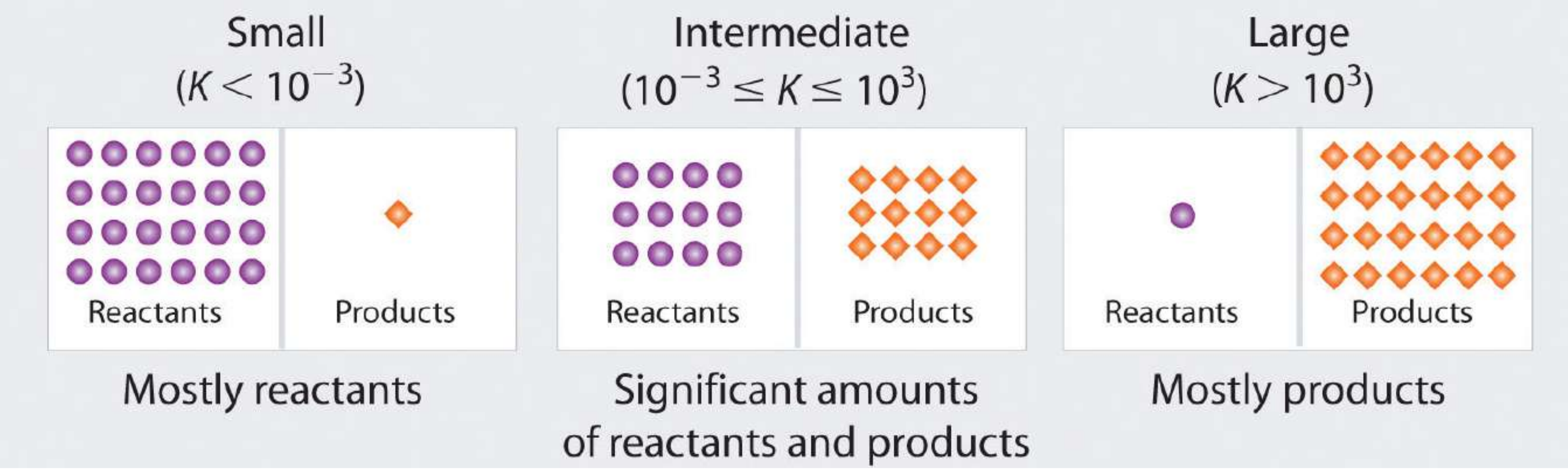
LO 6.6: Given a set of initial conditions (concentrations or partial pressures) and K, use stoichiometric relationships and the law of mass action (Q equals K at equilibrium) to determine qualitatively and/or quantitatively the conditions at equilibrium for a system involving a single reversible reaction.

Source

Magnitude of K

Video

- 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.



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.

Le Chatelier's Principle

Animation Source

Video

- 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.

Change	**Direction System Shifts to Reestablish Equilibrium**
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 products
Decreasing the temperature	Endothermic: shifts towards reactants Exothermic: shifts towards products

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.

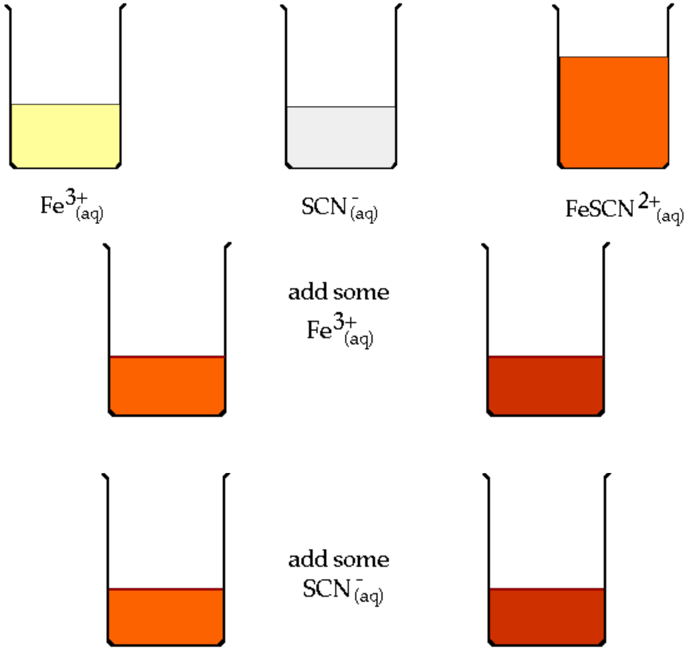
Animation Source

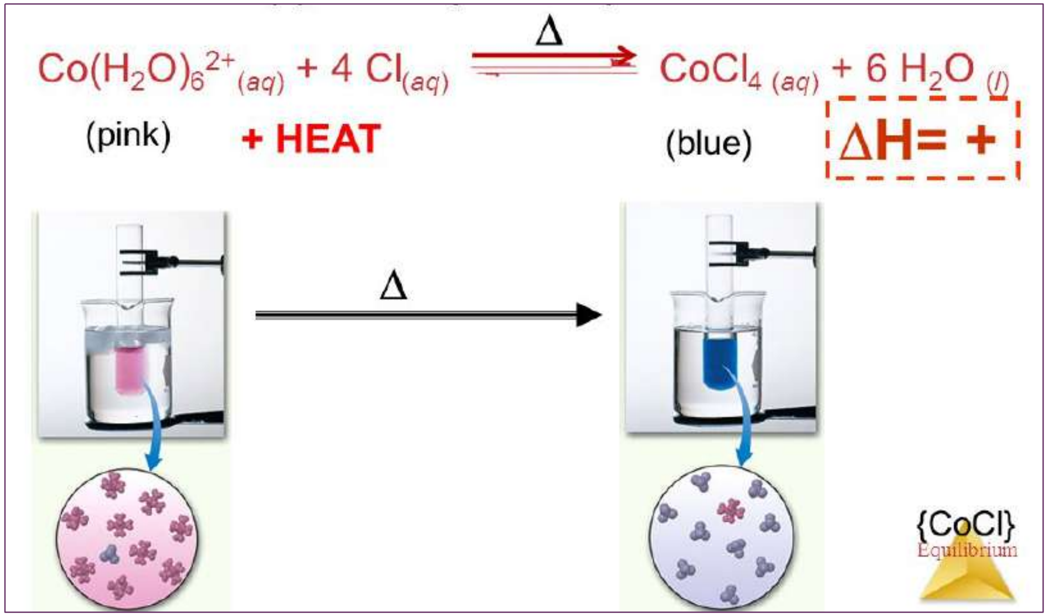
Video

+ 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)

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





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

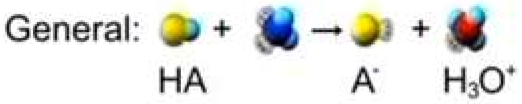
- Some changes that occur to a system at equilibrium will affect the reaction's current position (Q). Others will affect the value of K

Change	**Direction System Shifts to Reestablish Equilibrium**	**Effect on Q or K**
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

ining the

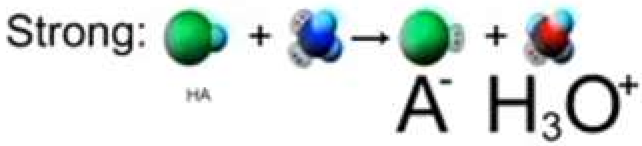
+ Acid/Base Particulates

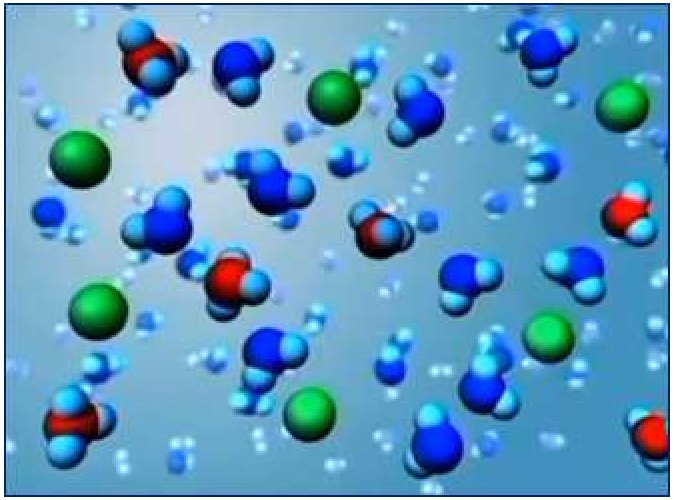


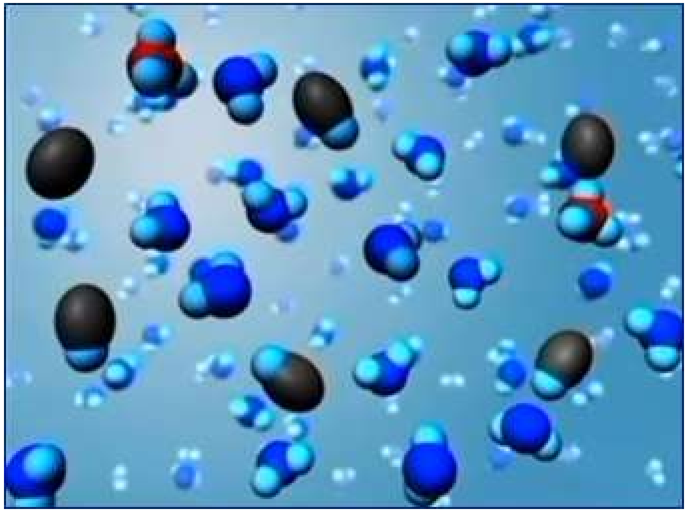
Source

Select Acid-Base Ionization

Video







Strong: Since $K_a = \frac{[\;\;_3\;\;^+][\;\;^-]}{[\;\;\;\;]} >> 1$, 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 $K_a = \frac{[\;\;_3\;\;^+][\;\;^-]}{[\;\;\;\;]} \ll 1$, at equilibrium weak acids are molecules that only partially ionize in aqueous solution, disassociating into few H_3O^+ ions and the additional anion.

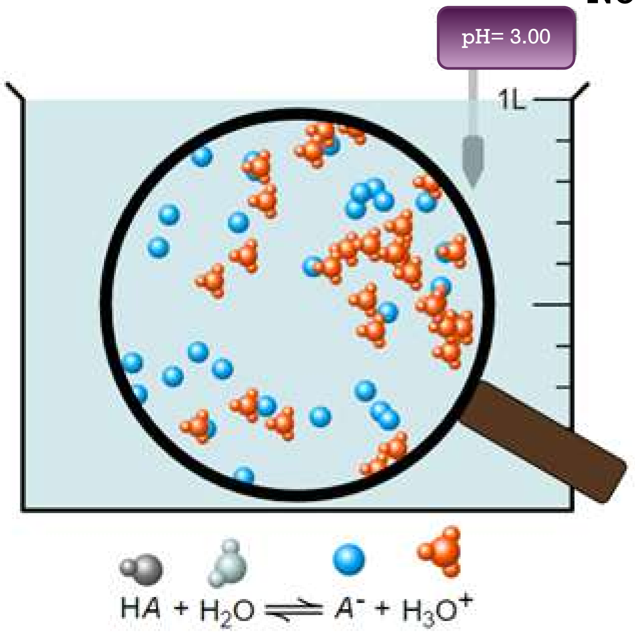
LO 6.11: The student can generate or use a particulate representation of an acid (strong or weak or polyprotic) and a strong base to explain the species that will have large versus small concentrations at equilibrium.

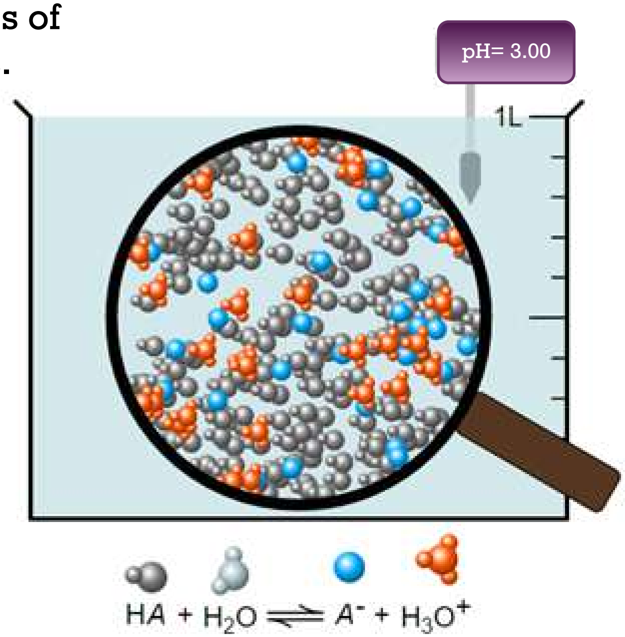
Source

Video

+ pH of Weak or Strong Acid

Note the similar pH values of both monoprotic acids.





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

- This is a particulate picture of a weak acid whose [HA] = 1.00*M* and $K_a = 1.00\text{x}10^{-6}$.
- 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.

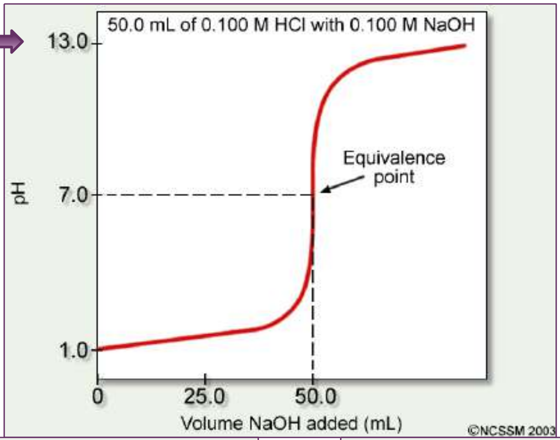
LO 6.12: Reason about the distinction between strong and weak acid solutions with similar values of pH, including the percent ionization of the acids, the concentrations needed to achieve the same pH, and the amount of base needed to reach the equivalence point in a titration.

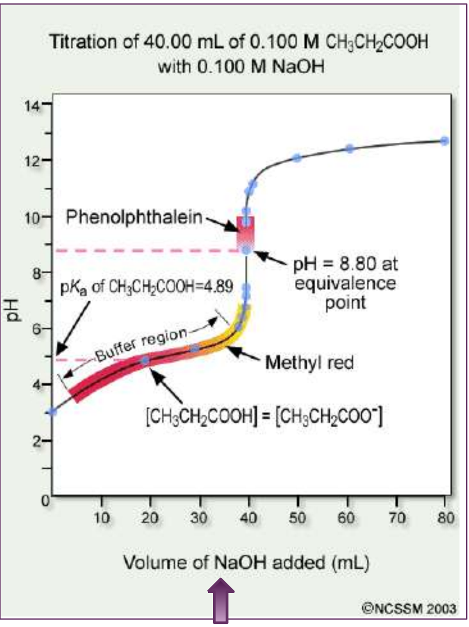
+ Titrations

Source

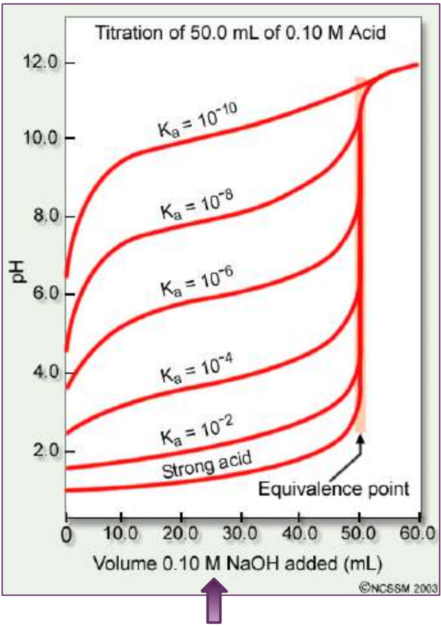
Video

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

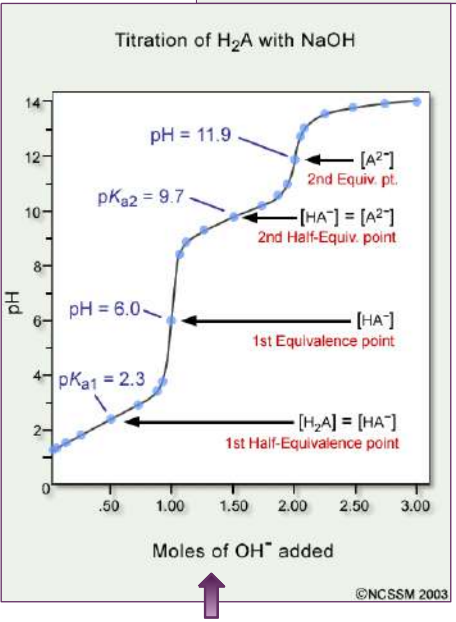




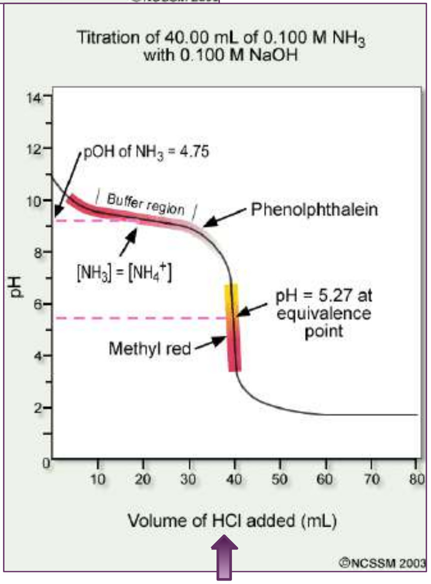
This illustration shows the titration curve of a weak acid with a strong base with indicator changes.



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.

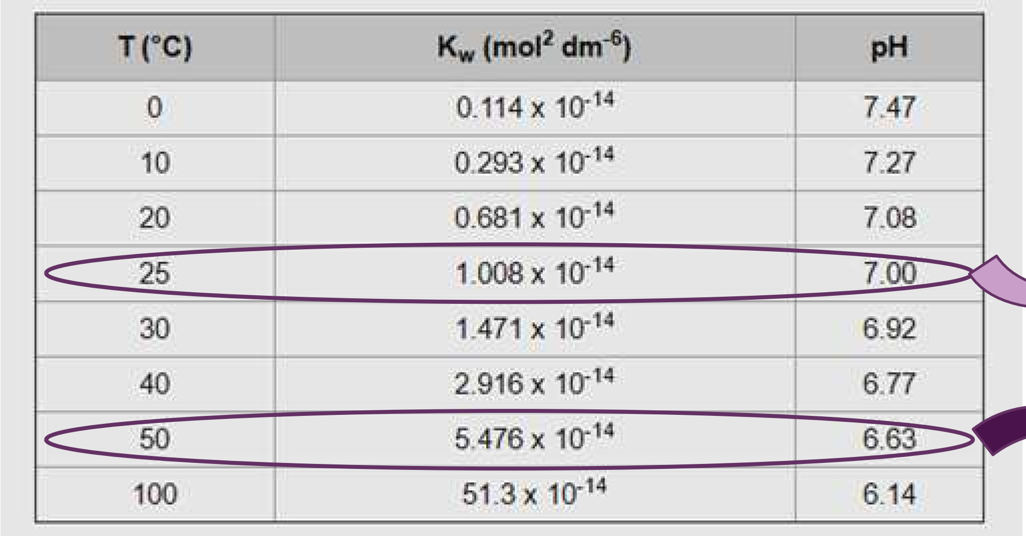
LO 6.13: The student can interpret titration data for monoprotic or polyprotic acids involving titration of a weak or strong acid by a strong base (or a weak or strong base by a strong acid) to determine the concentration of the titrant and the pK_a for a weak acid, or the pK_b for a weak base.

+ K_w and Temperature

Source

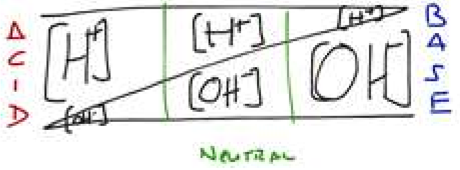
Video

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



T (°C)	K_w (mol² 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 [H+] > [OH-].
- At 50°C, the pH of pure water is 6.63, which is defined as "neutral", when $[H^+] = [OH^-]$. 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.



$$E + H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

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 $[H^+] = [OH^-]$ as opposed to requiring pH = 7, including especially the application to biological systems.

Source

Acid/Base Mixtures and its pH

Video

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. pH < 7
b. pH = 7
c. pH > 7
d. pH = pK_a

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.

+ pH and Acid/Base Equilibria

Source

Video

Source link for more calculations.

1. Vinegar is 0.50M acetic acid, $HC_2H_3O_2$, with a $K_a = 1.8 \times 10^{-5}$. What would be the pH of this solution?

1st click reveals answer and explanation.

2. Identify and compare the relative strengths of the two acids and the two bases in this neutralization reaction: $OH^-(aq) + NH_4^+(aq) \leftrightarrows H_2O(l) + NH_3(aq)$.

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.

Source

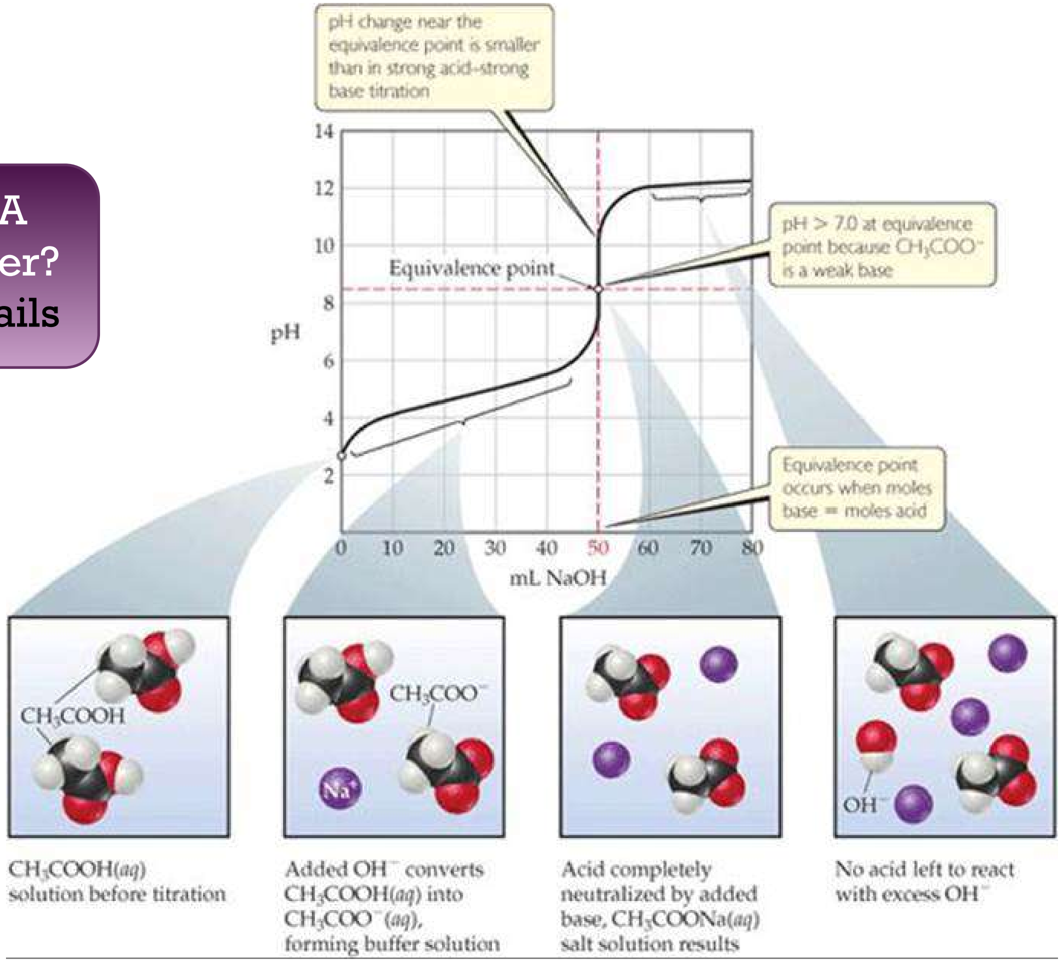
+ Acid/Base reaction species

Video

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.



CH$_3$COOH(*aq*) solution before titration

Added OH$^-$ converts CH$_3$COOH(*aq*) into CH$_3$COO$^-$(*aq*), forming buffer solution

Acid completely neutralized by added base, CH$_3$COONa(*aq*) salt solution results

No acid left to react with excess OH$^-$

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:

Source

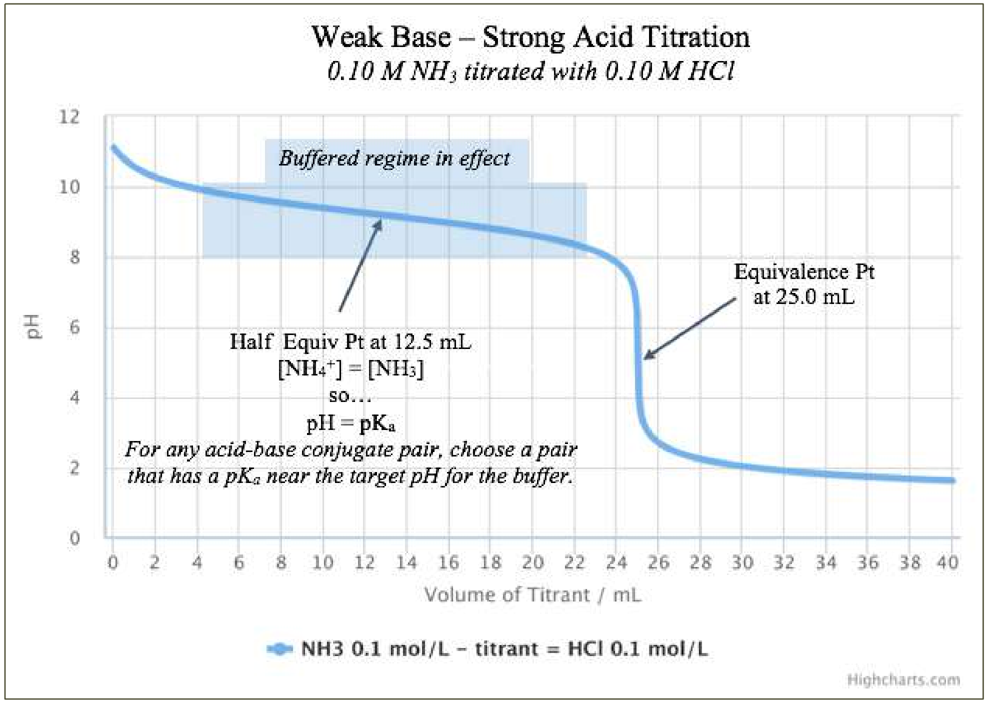
Video

Getting the pH correct:

- The pH of a buffer is primarily determined by the pKa 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 pKa.
- Choose a conjugate acid-base pair that has a pKa 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.



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.

Source

+ Finding the Major Species

Video

- A 50.0 mL sample of 0.50 M $HC_2H_3O_2$ 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 $HC_2H_3O_2$ is 4.7)

a. $[HC_2H_3O_2] > [C_2H_3O_2{}^{1-}] > [\,H^+\,] > [\,OH^-\,]$

b. $[HC_2H_3O_2] = [C_2H_3O_2{}^{1-}] > [\,H^+\,] > [\,OH^-\,]$

c. $[HC_2H_3O_2] > [C_2H_3O_2{}^{1-}] = [\,H^+\,] > [\,OH^-\,]$

d. $[C_2H_3O_2{}^{1-}] > [HC_2H_3O_2] > [\,OH^-\,] > [\,H^+\,]$

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.

Source

+ 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 hydrofluoric acid and NaF. A model 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 the added acid will react completely with the available base, F^-. This results in a nearly unchanged [H_3O^+] and a nearly unchanged pH.

$$H_3O^+_{(aq)} + F^-_{(aq)} \rightarrow HF_{(aq)} + H_2O_{(l)}$$

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 [H_3O^+] and a nearly unchanged pH.

$$HF_{(aq)} + OH^-_{(aq)} \rightarrow F^-_{(aq)} + H_2O_{(l)}$$

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

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$$

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

Video

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

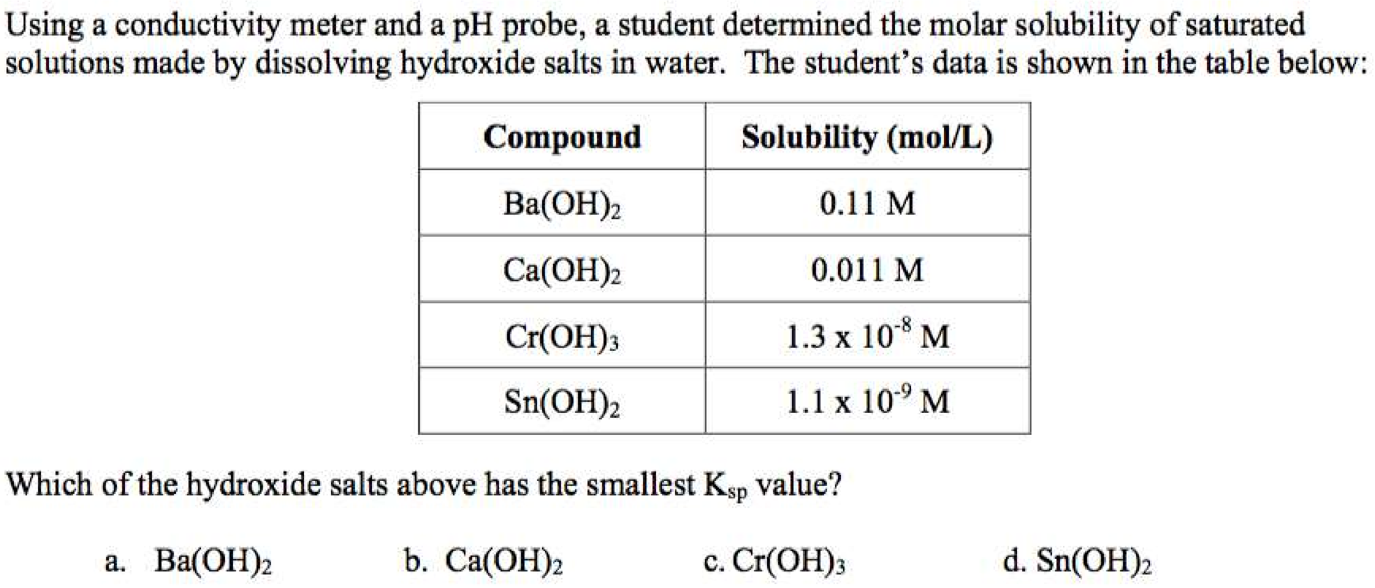
4.0×10^{-12} b. 2.0×10^{-12} c. 4.0×10^{-6} d. 2.0×10^{-6}

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.

Source

+ Find a K_{sp} from solubility data



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
$Cr(OH)_3$	1.3×10^{-8} M
$Sn(OH)_2$	1.1×10^{-9} M

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. $Sn(OH)_2$

Video

Click reveals answer and explanation.

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

Source

Common Ion Effect

Video

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

$$Mg(OH)_{2(s)} \rightleftharpoons Mg^{2+}{}_{(aq)} + 2OH^{-}{}_{(aq)} \qquad \text{With } Ksp=5.61\times10^{-12}$$

When acid is added to a saturated solution that contains excess solid $Mg(OH)_2$, the following reaction occurs, removing OH− from solution:

$$H^{+}{}_{(aq)} + OH^{-}{}_{(aq)} \rightarrow H_2O_{(l)}$$

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

$$Mg(OH)_{2(s)} + 2H^{+}{}_{(aq)} \rightleftharpoons Mg^{2+}{}_{(aq)} + 2H_2O_{(l)} \qquad (18.7.7)$$

As more acid is added to a suspension of $Mg(OH)_2$, the equilibrium shown in Equation 18.7.7 is driven to the right, so more $Mg(OH)_2$ 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 $Mg(OH)_2$ solution, additional $Mg(OH)_2$ will precipitate out. The additional Mg^{2+} ions will shift the original equilibrium to the left, thus reducing the solubility of the magnesium hydroxide.

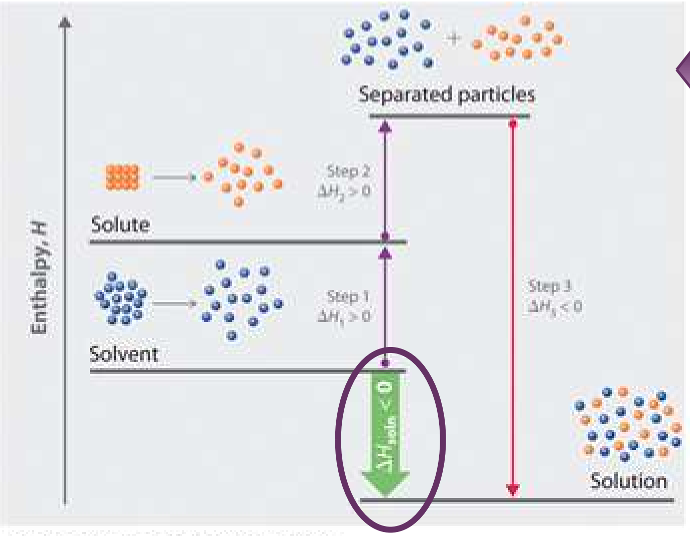
LO 6.23: The student can interpret data regarding the relative solubility of salts in terms of factors (common ions, pH) that influence the solubility.

Source

+ Salt dissolution: ΔH and ΔS

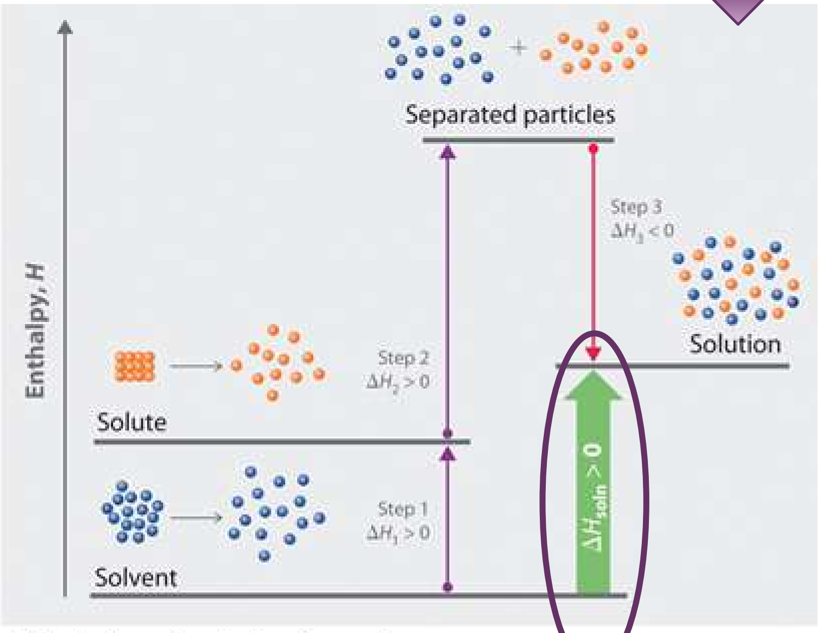
Video

Java Tutorial

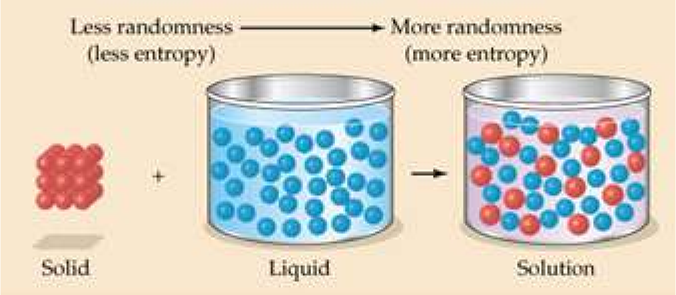


(a) **Exothermic solution formation**

- 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.

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

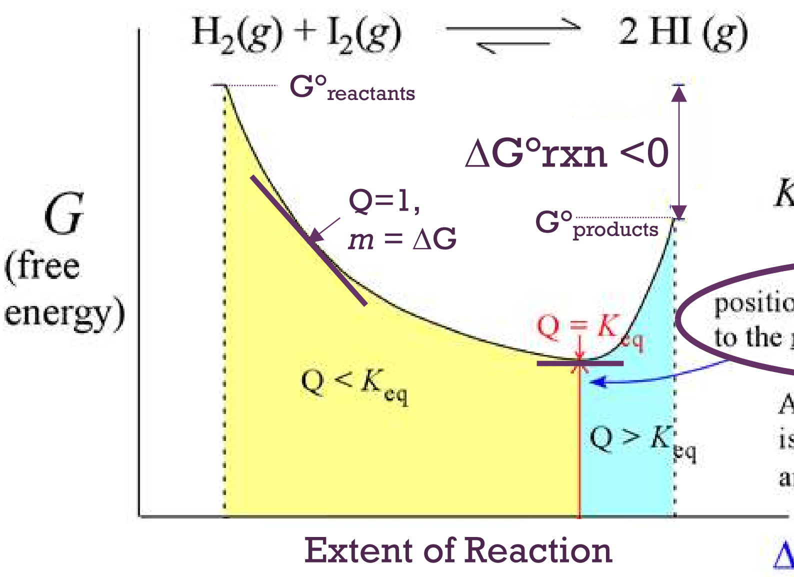
Source

Video

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

$$= -RT\ln K$$

$$= -nFE^\circ$$



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

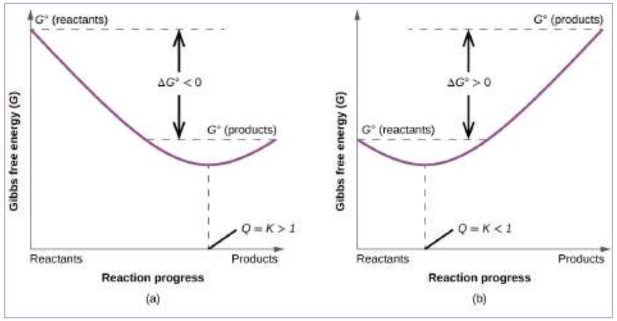
$$K_{eq} = \frac{[HI]_{eq}^2}{[H_2]_{eq}[I_2]_{eq}} = 49.5$$

(440°C)

position of equilibrium corresponds to the <u>minimum</u> of free energy

At the minimum, the system is in <u>dynamic equilibrium</u> and $\Delta G = 0$.

$$\Delta G_{rxn} = \Delta G^\circ_{rxn} + RT\ln Q$$



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.

Science Practices

Laboratory Exercises

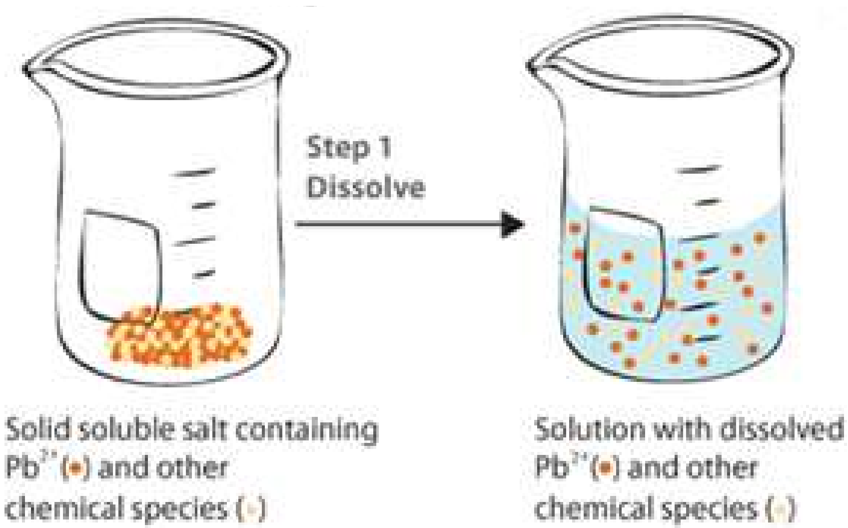
Gravimetric Analysis

Source

Video

Virtual Lab

- What It Determines:
  - amount of analyte by mass measurements
  - % composition
  - empirical formulas
- How It Can Be Done:
  - Dehydration of a hydrate
  - Forming a precipitate, which is then isolated and massed
- Analysis:
  - Remember to ALWAYS go to moles!
  - use mole ratios to convert between various components



Possible Source of Error	**Contamination in solid**	**Incomplete Precipitation**	**Solid not fully dehydrated**
Impact on Results	measured mass too large	Lose ion in filtrate – mass too small	measured mass too large

Science Practices

1.5 The student can *re-express key elements* of natural phenomena across multiple representations in the domain.

2.2 The student can *apply mathematical routines* to quantities that describe natural phenomena.

4.2 The student can *design a plan* for collecting data to answer a particular scientific question.

5.1 The student can *analyze data* to identify patterns or relationships.

6.1 The student can *justify claims with evidence*.

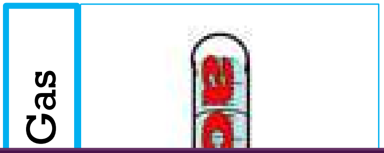
Learning Objectives

1.3: The student is able to select and apply mathematical relationships to mass data in order to justify a claim regarding the identity and/or estimated purity of a substance.

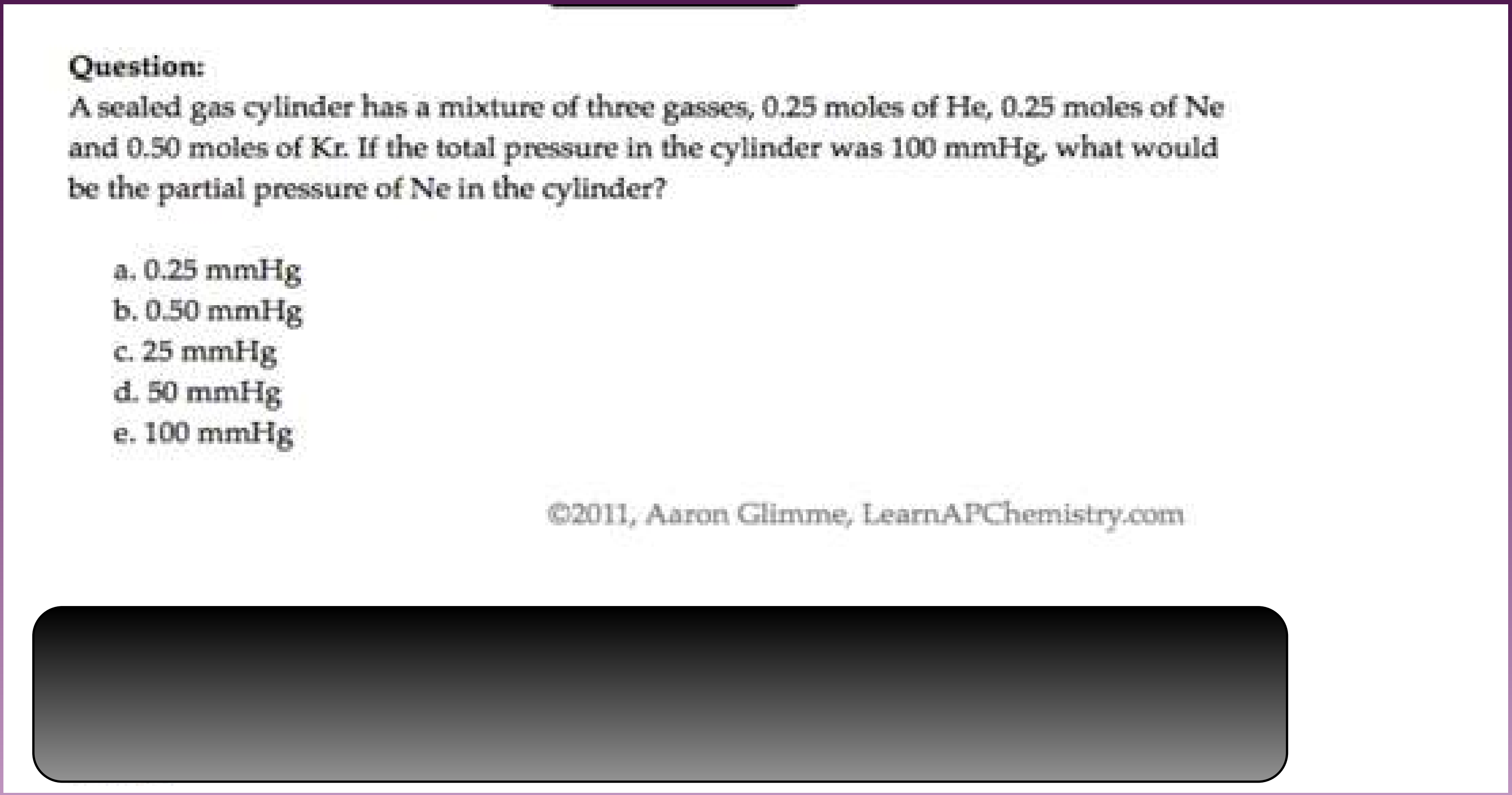
1.17: The student is able to express the law of conservation of mass quantitatively and qualitatively using symbolic representations and particulate drawings.

1.19: The student can design, and/or interpret data from, an experiment that uses gravimetric analysis to determine the concentration of an analyte in a solution.

+ Gas Laws Labs



Source 1



Question:

A sealed gas cylinder has a mixture of three gasses, 0.25 moles of He, 0.25 moles of Ne and 0.50 moles of Kr. If the total pressure in the cylinder was 100 mmHg, what would be the partial pressure of Ne in the cylinder?

a. 0.25 mmHg
b. 0.50 mmHg
c. 25 mmHg
d. 50 mmHg
e. 100 mmHg

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Science Practices

1.4 The student can *use representations and models* to analyze situations or solve problems qualitatively and quantitatively.
2.2 The student can *apply mathematical routines* to quantities that describe natural phenomena.
2.3 The student can *estimate numerically* quantities that describe natural phenomena.
6.4 The student can *make claims and predictions about natural phenomena* based on scientific theories and models.

Learning Objectives

LO 2.4: The student is able to use KMT and concepts of intermolecular forces to make predictions about the macroscopic properties of gases, including both ideal and nonideal behaviors.
LO 2.6: The student can apply mathematical relationships or estimation to determine macroscopic variables for ideal gases.

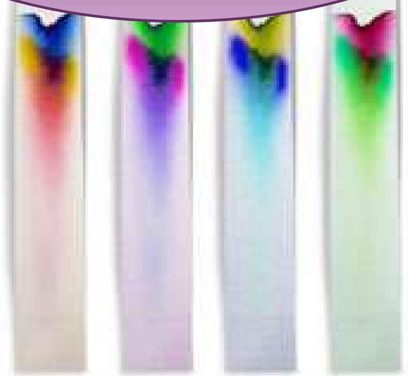
+ Chrom

Click here to watch animation explaining column chromatography

Chromatograph

All chromatographic techniques flow the mixture, that is to be separated, through a material that **retains some components more than others.**

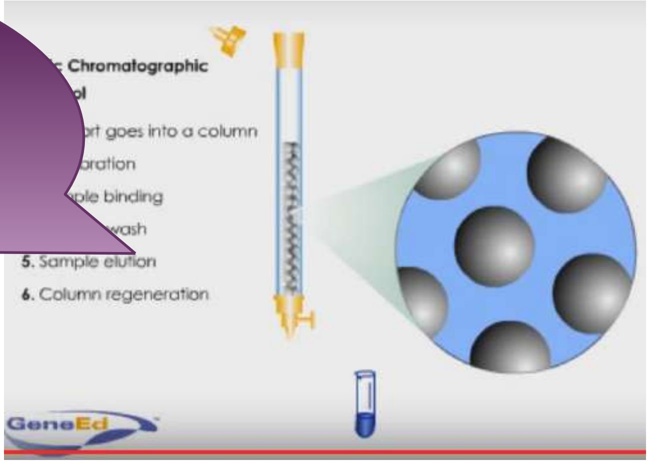
This causes different components to flow through the material at different speeds, so they separate.



Source (includes nice discussion of biological applications)

- Main Error:
  - Incomplete Separation
- Can be caused by:
  - Inconsistent spotting
  - Overloading
  - Improper packing (column)
  - Poor solvent choice

Advantage: We can collect and use the fractions



c Chromatographic
ol
rt goes into a column
aration
ple binding
wash
5. Sample elution
6. Column regeneration

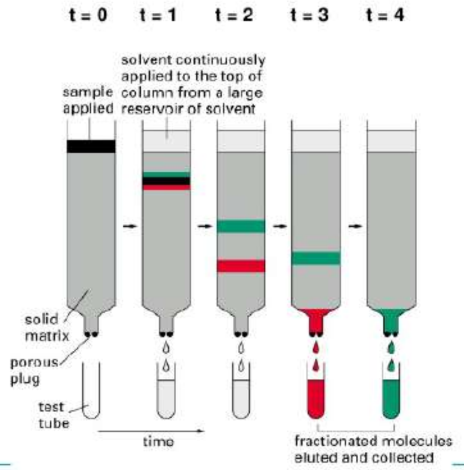
Source

Video

Virtual Lab

Column Chromatography

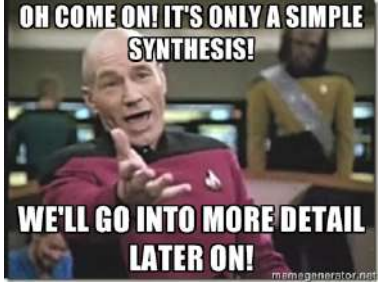
- Column chromatography: technique that uses an adsorbent packed in a glass column, and a solvent that moves down slowly through the packed column.



Science Practices
4.2 The student can *design a plan* for collecting data to answer a particular scientific question.
5.1The student can *analyze data* to identify patterns or relationships.
6.2 The student can *construct explanations of phenomena based on evidence* produced through scientific practices.

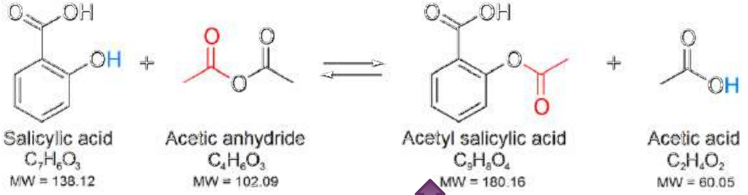
Learning Objectives
LO 2.7: The student is able to explain how solutes can be separated by chromatography based on intermolecular interactions.
LO 2.10: The student can design and/or interpret the results of a separation experiment (filtration, paper chromatography, column chromatography, or distillation) in terms of the relative strength of interactions among and between the components.

Synthesis



Source

- Inorganic Synthesis Examples:
  - Synthesis of Coordination Compound
  - Synthesis of Alum from aluminum
- Organic Synthesis Example:
  - Synthesis of Aspirin



Video

Virtual Lab

- Syntheses are done in solution, and require purification by filtration, recrystalization, column chromatography or a combination
- % Yield is calculated; analysis is done by melting point, NMR, IR

Synthesizing Alum Procedure:

Source of aluminum is reacted with KOH

Possible Error Sources: impurity in Al (such as plastic coating) prevents complete rxn

Science Practices

2.1 The student can *justify the selection of a mathematical routine* to solve problems.

2.2 The student can *apply mathematical routines* to quantities that describe natural phenomena.

4.2 The student can *design a plan* for collecting data to answer a particular scientific question.

6.1 The student can *justify claims with evidence*.

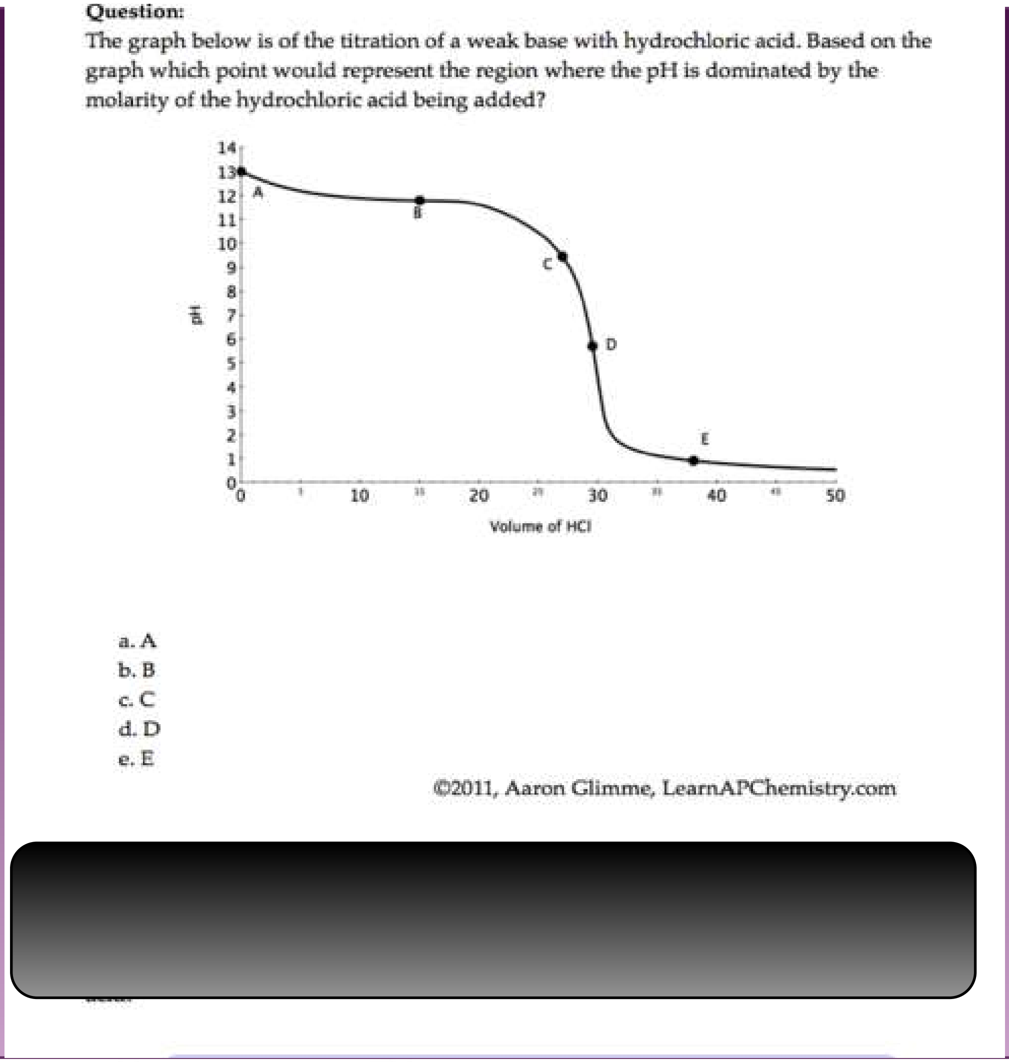
Learning Objectives

LO 3.5: The student is able to design a plan in order to collect data on the synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.

LO 3.6: The student is able to use data from synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.

Source 1

+ Titration- Acid/Base



Question:

The graph below is of the titration of a weak base with hydrochloric acid. Based on the graph which point would represent the region where the pH is dominated by the molarity of the hydrochloric acid being added?

a. A
b. B
c. C
d. D
e. E

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[redacted]

Science Practices

1.4 The student can *use representations and models* to analyze situations or solve problems qualitatively and quantitatively.

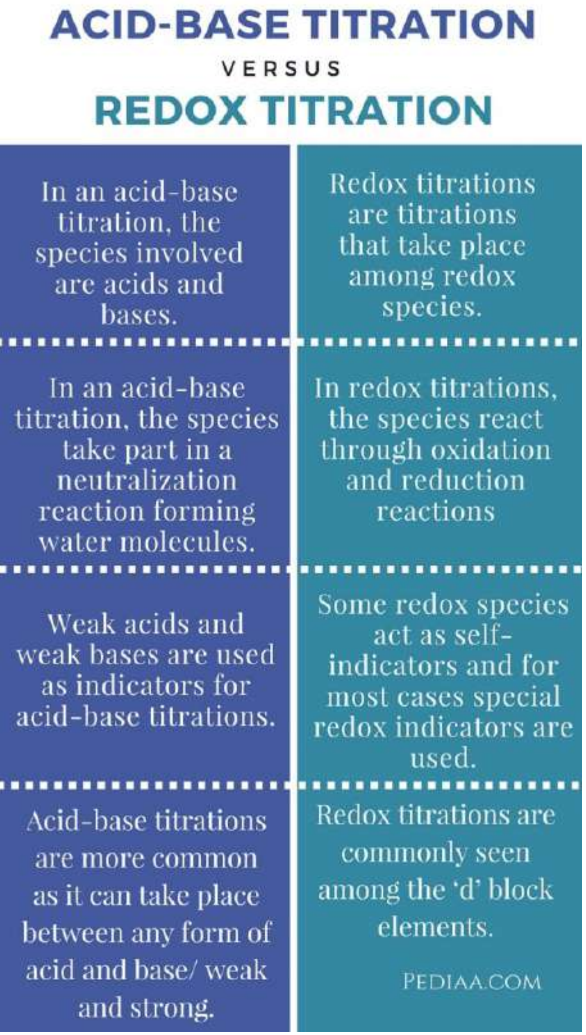
5.1 The student can *analyze data* to identify patterns or relationships.

Learning Objectives

LO 6.12: The student can 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.

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 pKa for a weak acid, or the pKb for a weak base.

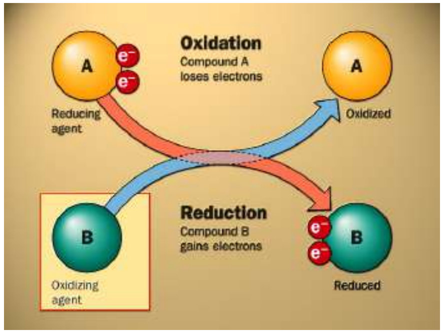
REDOX Titration



ACID-BASE TITRATION
VERSUS
REDOX TITRATION

In an acid-base titration, the species involved are acids and bases.	Redox titrations are titrations that take place among redox species.
In an acid-base titration, the species take part in a neutralization reaction forming water molecules.	In redox titrations, the species react through oxidation and reduction reactions
Weak acids and weak bases are used as indicators for acid-base titrations.	Some redox species act as self-indicators and for most cases special redox indicators are used.
Acid-base titrations are more common as it can take place between any form of acid and base/ weak acid and base/ weak and strong.	Redox titrations are commonly seen among the 'd' block elements.

PEDIAA.COM

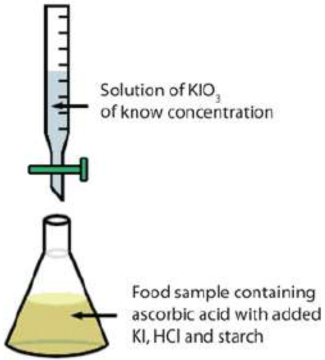
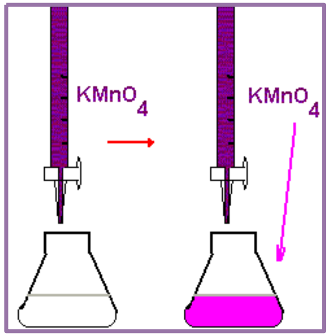
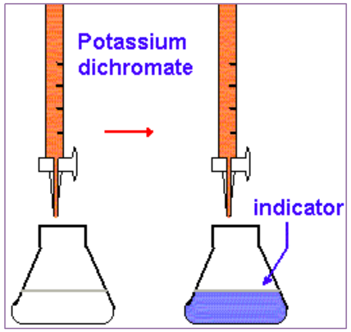


Source 1

Video

Virtual Lab

Applications/Methods of REDOX Titrations:

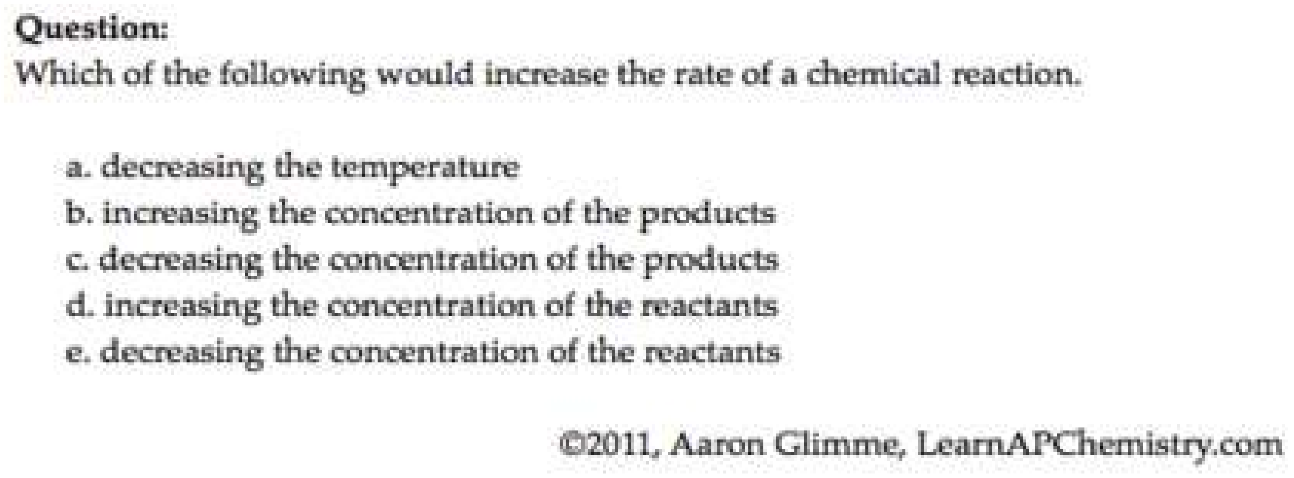
Vitamin C Content	 Source
Concentration of a redox species	  Source
Back Titrations	Done when the endpoint can be hard to ID – add an excess, and then back titrate to determine how much is in excess

Science Practices
4.2 The student can *design a plan* for collecting data to answer a particular scientific question.
5.1 The student can *analyze data* to identify patterns or relationships.

Learning Objectives
LO 1.20: The student can design, and/or interpret data from, an experiment that uses titration to determine the concentration of an analyte in a solution.
LO 3.9: The student is able to design and/or interpret the results of an experiment involving a redox titration.

Virtual Lab 1 Virtual Lab 2 Video Source 1

Kinetics



Question:
Which of the following would increase the rate of a chemical reaction.

a. decreasing the temperature
b. increasing the concentration of the products
c. decreasing the concentration of the products
d. increasing the concentration of the reactants
e. decreasing the concentration of the reactants

Science Practices

2.1 The student can *justify the selection of a mathematical routine* to solve problems.
2.2 The student can *apply mathematical routines* to quantities that describe natural phenomena.
5.1 The student can *analyze data* to identify patterns or relationships.

Learning Objectives

LO 4.2: The student is able to analyze concentration vs. time data to determine the rate law for a zeroth-, first-, or second-order reaction.
LO 4.3: The student is able to connect the half-life of a reaction to the rate constant of a first-order reaction and justify the use of this relation in terms of the reaction being a first-order reaction.

+ Calorimetry

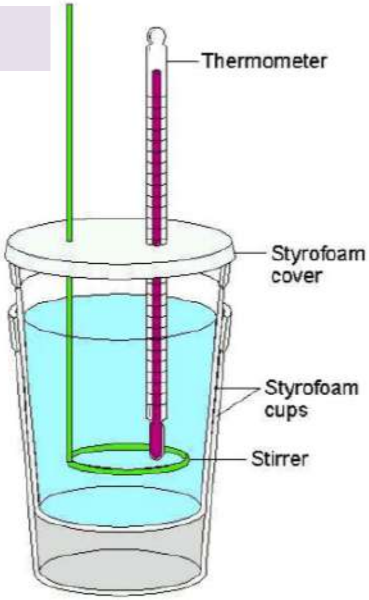
Application of <u>Law of Conservation of Energy</u>: All heat produced/consumed during reaction (system) is exchanged with the surroundings.

Source 1

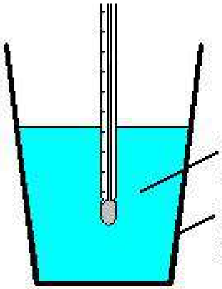
Coffee Cup Calorimeter

Applications

A calorimeter is used to measure the heat absorbed or released in a chemical (or other) process by measuring the temperature change of an insulated mass of water.



Constant P



Reaction occurs in solution.
Temperature increases or decreases.
Insulating cup traps heat or keeps out heat from the surroundings.

n:

Source

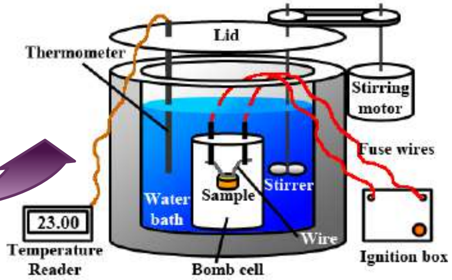
Applications:

- Specific Heat of Material
  - Can be used to ID unknown
- Heat of reaction
- Energy content of food
  - This is usually done in a bomb calorimeter (constant V)

Video

Virtual Lab

Source



<u>Science Practices</u>

1.4 The student can *use representations and models* to analyze situations or solve problems qualitatively and quantitatively.

2.2 The student can *apply mathematical routines* to quantities that describe natural phenomena.

4.2 The student can *design a plan* for collecting data to answer a particular scientific question.

5.1 The student can *analyze data* to identify patterns or relationships.

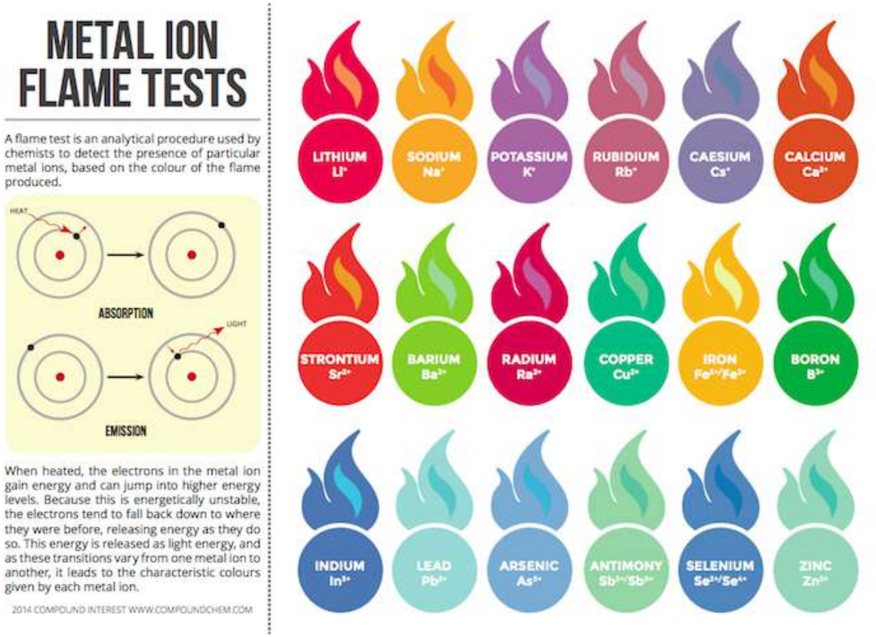
<u>Learning Objectives</u>

LO 5.4: The student is able to use conservation of energy to relate the magnitudes of the energy changes occurring in two or more interacting systems, including identification of the systems, the type (heat versus work), or the direction of energy flow.

LO 5.7: The student is able to design and/or interpret the results of an experiment in which calorimetry is used to determine the change in enthalpy of a chemical process (heating/cooling, phase transition, or chemical reaction) at constant pressure.

Virtual Lab 1 Virtual Lab 2 Video Source

Qualitative Analysis



METAL ION FLAME TESTS

A flame test is an analytical procedure used by chemists to detect the presence of particular metal ions, based on the colour of the flame produced.

When heated, the electrons in the metal ion gain energy and can jump into higher energy levels. Because this is energetically unstable, the electrons tend to fall back down to where they were before, releasing energy as they do so. This energy is released as light energy, and as these transitions vary from one metal ion to another, it leads to the characteristic colours given by each metal ion.

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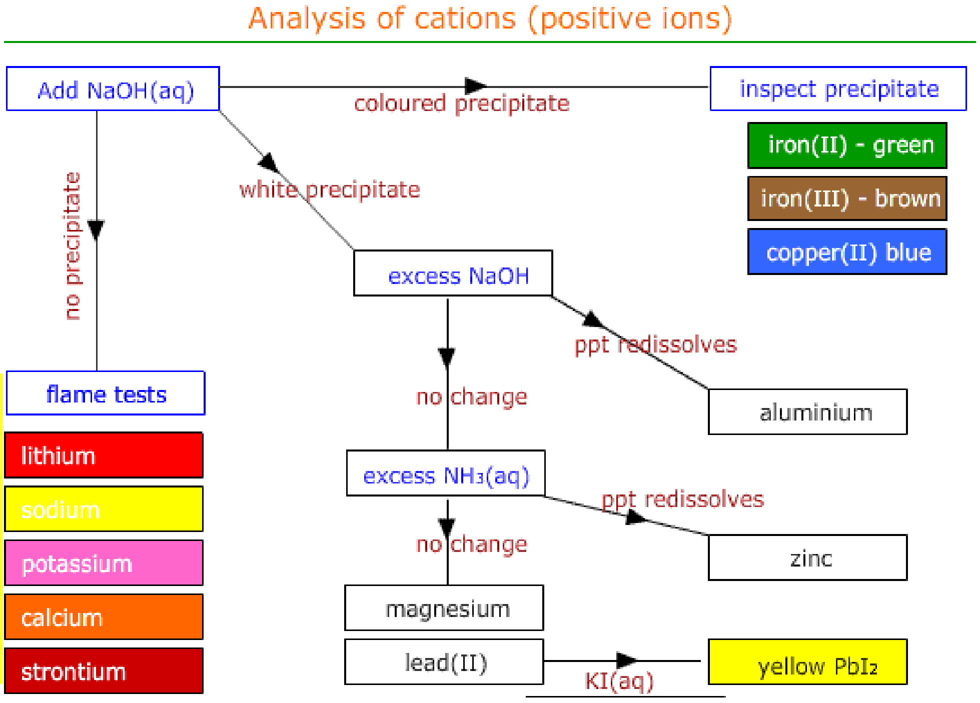


Analysis of cations (positive ions)

Remember that the lower the K_{sp}, the less soluble the substance. Qualitative analysis then also allows a ranking of salts by K_{sp}, and demonstrates the effect of pH on solubility.

Science Practices

2.2 The student can *apply mathematical routines* to quantities that describe natural phenomena.

2.3 The student can *estimate numerically* quantities that describe natural phenomena.

5.1 The student can *analyze data* to identify patterns or relationships.

6.4 The student can *make claims and predictions about natural phenomena* based on scientific theories and models.

Learning Objectives

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

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

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.

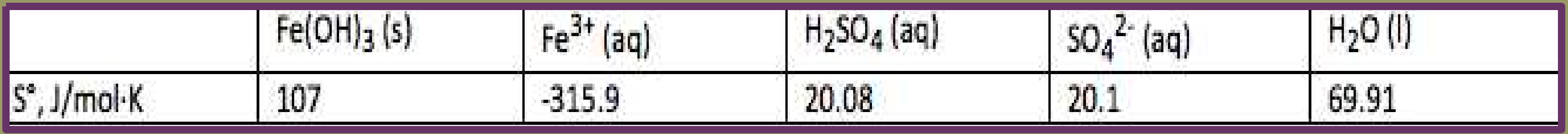
+

Now…

Just for Fun: Some "Harry" Problems

$2Fe(OH)_3$ (s) + $3H_2SO_4$ (aq) à $Fe_2(SO_4)_3$ (aq) + $6H_2O$ (l)

- a. Ron Weasley had to perform the above reaction in Professor Snape's potions class. Ron mixed 45.0 g iron III hydroxide with 45.0 ml 2.00 M sulfuric acid in a cauldron, leaving the lid off. The heat capacity of his cauldron was 645 J/K.
- i. What other information would Ron need to know or measure experimentally to determine the molar change in enthalpy of the reaction, $\Delta H°rxn$?
- ii. Unbeknownst to Ron, Draco Malfoy dropped a 100 g sneakoscope (assume this is unreactive and at room temperature) into Ron's cauldron before Ron measured his temperature but after the reaction was initiated. Explain in which direction this would alter Ron's experimental value of $\Delta H°rxn$.
- b. Hermione Granger also performed this reaction. She mixed 45.0 g iron III hydroxide with 450 ml 2.00 M sulfuric acid in her cauldron, and kept the lid on, using a spell to enable her to measure temperature (revelio thermo).
- i. How would you expect her temperature change to differ from Ron's, if at all? For this comparison, assume that Ron's reaction was performed with the lid on and with no sneakoscope.
- ii. Hermione repeated the experiment using all the same measurements, but used aluminum hydroxide instead of iron III hydroxide. Assuming that both reactions have the same $\Delta H°rxn$, would her temperature change be smaller, the same, or greater in her second experiment.
- c. Given the following values for absolute entropies, what would $\Delta S°rxn$ be?



	$Fe(OH)_3$ (s)	Fe^{3+} (aq)	H_2SO_4 (aq)	SO_4^{2-} (aq)	H_2O (l)
S°, J/mol·K	107	-315.9	20.08	20.1	69.91

Just for Fun… Answers

- ai. $q_{rxn} = -(C_{cal}\Delta T + c_{(mixture)}m_{(mixture)}\Delta T)$; $\Delta H°rxn = q_{rxn}$ / moles rxn
- Ron would need to measure the starting and final temperatures, know (or measure) the specific heat capacity and mass of the reaction mixture, and the number of moles of reaction. In this case,
- 45.0 g $Fe(OH)_3$ / 106.8 g/mole = 0.421 moles; 0.045 L x 2 M = 0.09 moles H_2SO_4 (limiting), so 0.03 moles reaction (to those who dislike this concept, I apologize).

- aii. Because the sneakoscope would absorb heat, ΔT would be lower, so the heat of the reaction would appear to be lower as would $\Delta H°rxn$.

- bi. For Hermione, the $Fe(OH)_3$ is now limiting, as there are now 0.9 moles H_2SO_4. So, there are more moles of reaction, about 0.21, a seven-fold increase. This would increase the temperature change. However, there is also a significant increase in mass of the reaction mixture, from about 90 g to 500 g. This would decrease the temperature change. So the exact ΔT outcome, increase or decrease, depends on the specific heat capacity of the reaction mixture.

- bii. 45.0 g $Al(OH)_3$ / 78 g/mole = 0.577 mole. This is still limiting, but now there are about 0.288 moles of reaction, so more heat is produced. As the total mass doesn't change, and the specific heat capacity is unlikely to change dramatically, it is reasonable to predict that ΔT would be larger.

- c. $\Delta S°rxn$ = sum product S° - sum reactant S°
- $\Delta S°rxn = 6(69.91) + 2(-315.9) + 3(20.1) – 2(107) – 3(20.08) = -426.3$ J/mole*K.