CH 302 – Exam Review

CHEMICAL EQUILIBRIUM, ACIDS & BASES

"THINKING LIKE A CHEMIST"

Exam Two Learning Outcomes

Learning Outcomes for Chemical Equilibrium

Describe the relationship between free energy and equilibrium

Convert ΔG to Q*, as well as* $\Delta G°$ to K and vice versa.

Know the importance of the activity of a species and how it relates to concentration, pressure, and equilibrium.

Write the mass action expression for homogeneous and heterogeneous equilibria.

Determine new values for K when combining multiple reactions.

Determine if a system is at equilibrium and if not which direction the reaction will shift to achieve equilibrium.

Know the difference between K_p and K_c and be able to convert between the two.

Set up and solve a RICE table for a multitude of various reaction types.

Calculate the concentration/pressure of all species at equilibrium.

Show a complete understanding of Le Chatelier's principle.

Predict the response of a reaction to an applied stress (concentration, pressure, volume, temperature) both qualitatively and quantitatively.

Calculate the new value of K when the temperature changes to a new value.

Exam Two Learning Outcomes

Learning Outcomes for Acids and Bases

Understand the strength of an acid (or base) as determined by the percent of ionization in solution.

Identify strong and weak acids and bases.

Identify acid/base conjugate pairs and their relative strengths.

Understand the process of auto-ionization of water and what is meant by acidic, basic, and neutral.

Know the value of K_w at 25°C, and the relationship between K_a and K_b for a conjugate pair.

Convert between hydronium ion concentration, hydroxide ion concentration, pH and pOH for a given solution.

Determine the pH of a strong acid or base solution.

Determine the pH of a weak acid or weak base solution.

Determine the pH of the solution made from the salt of a weak acid or the salt of a weak base.

Recognize and predict the components of a buffer solution.

Calculate the pH of a buffer solution, and a buffer solution after the addition of strong acid or strong base.

Determine the majority species for acid/base solutions as well as the pH following neutralization.

Interpret a titration curve plot including calculating the concentration and K_a or K_b for the analyte.

Understand the concept of an acid/base indicator, and determine which indicators are appropriate for a given titration.

Determine the protonation state (or overall charge) for a polyprotic species at a particular pH.

Apply concepts from equilibria to acid/base problems

Doing Well on Exam 2

Attacking Problems on Exam 2

1. Identify the type of problem (Type 1, Type 2, Chemical Equilibria, buffer, titration, etc.)
2. Think about the path to the correct answer: I will have to use the Henderson-Hasselbalch equation, I will need to write out the neutralization reaction, I will need to make a RICE table, etc.
3. Reflect on your answer: I expect the pH of $NaNO_2$ to be basic, $[OH^-]$ in a HCl solution should be very small, a ratio of 0.6M $NaNO_2$: 0.5M HNO_2 should have pH > pKa, etc.

See how the whole unit is one continuous ”story”

1. Students often get overwhelmed by the content in Unit 2 because we introduce acids/ bases and various experiments that might be hard to visualize (buffers, titrations, indications, etc.)
2. Acid/Base problems are just applied chemical equilibria problems. They are simplified to make the math easier, which gives rise to a bunch of new equations.
3. K follows the same rules, regardless of if it is Ka, Kc, Kp, Kw, Kb, etc. In other words, all are measurements of how much a reaction proceeds forward, all are temperature dependent, etc

Chemical Equilibrium

K, REACTIONS AT EQUILIBRIUM, STRESSING EQUILIBRIUM

Brief overview of Equilibrium Terminology

- We use K and ΔG° to express which side of a reaction is "favored"
  - If K > 1 , the products are favored (ΔG° is negative)
  - If K < 1 , the reactants are favored (ΔG° is positive)
- We use the relationship between Q and K to explain which way a reaction will progress toward equilibrium
  - If the starting point of a reaction is Q < K, the reaction moves forward toward equilibrium
  - If the starting point of a reaction is Q > K, the reaction moves backward toward equilibrium.
- We stress a system at equilibrium to examine how a reaction will "shift" to oppose the stress
  - Stressing equilibrium can either change Q or K, resulting in more products or reactants formed

Chemical Equilibrium Fundamentals

- The basis of chemical equilibrium is that reactions tend to progress to the lowest free energy. In many reactions, this is not 100% reactants or 100% products, rather a combination of the two.

$$\Delta G_r = \Delta G_r{}^\circ + RT \ln Q \longrightarrow$$

The free energy of a reaction under any measurable initial conditions

$$\Delta G_r{}^\circ = -RT \ln K \longrightarrow$$

At equilibrium, $\Delta G_r = 0$. Also, Q = K. Therefore, we get a new equation for the relationship between standard free energy and K.

$$K = e^{\frac{-\Delta G_r{}^\circ}{RT}} \longrightarrow$$

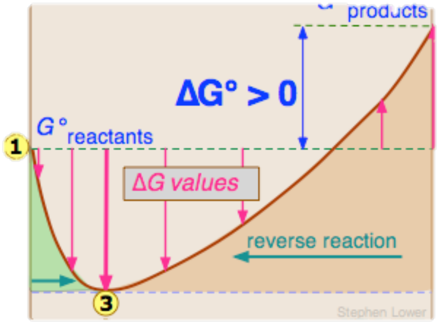
We can rearrange this equation to solve directly for a K value at a given temperature given the standard free energy change.

Free Energy (ΔG) vs K

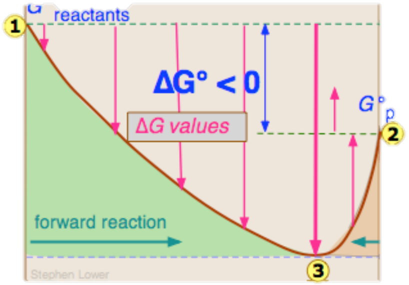
- When K is small (less than 1), only a small fraction of your reactants becomes products. A reaction with a small K reacts to a small extent.

- When K is large (greater than 1), a greater amount of products are formed than reactants remain. A reaction with a large K reacts to a large extent.

When the reactant is favored, the ΔG° is positive.



When the product is favored, the ΔG° is negative.



Note: these graphs are an exaggeration of K values. Very large K (think combustion) and very small K (think solubility examples) do not really look like this on a graph

Chemical Equilibrium Fundamentals

- We know that many chemical reactions do not go 100%. These equilibrium concentrations are determined by the equilibrium constant, K, and the mass action equation:

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

$$K = \frac{a_C^c \, a_D^d}{a_A^a \, a_B^b}$$

Mathematically, K is equal to the ratio of the action of the products raised to the power of their coefficients divided by the action of the reactants raised to the power of their coefficients. This is “Mass Action.”

$$\boxed{aA + bB \rightleftharpoons cC + dD}$$

Mass Action Expression

- We can directly correlate "action" to pressure and concentration to create a more sensible relationship.

$$K = \frac{a_C^c\, a_D^d}{a_A^a\, a_B^b}$$

$$a_i = \frac{[i]}{[i]^\circ} \longrightarrow \boxed{K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

$$a_i = \frac{P_i}{P_i^\circ} \longrightarrow \boxed{K_p = \frac{P_C^c P_D^d}{P_A^a P_B^b}}$$

REMEMBER: Assume for this example that all species in the K_c example are aqueous. All the species in the K_p example are gases. Remember: the action of any LIQUID or SOLID is 1. These terms will drop out of the mass action expression.

3 Ways to Manipulate K

1. Reverse the reaction

$$2A + B \rightleftharpoons 3C + 2D \text{ , K = 4}$$

$$3C + 2D \rightleftharpoons 2A + B \text{ , K = 0.25}$$

$$K' = K^{-1}$$

2. Multiply the coefficients by a factor, x

$$2A + B \rightleftharpoons 3C + 2D \text{ , K = 4}$$

$$4A + 2B \rightleftharpoons 6C + 4D \text{ , K = 16}$$

$$K' = K^{x}$$

3. Modify the temperature

$$\ln\left(\frac{K_2}{K_1}\right) = \frac{\Delta H_{rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Temperature Dependence of K

- By changing temperature, you are actually able to change K. We will discuss this more in La Chatelier’s Principle

$$\ln\left(\frac{K_2}{K_1}\right) = \frac{\Delta H_{rxn}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

By increasing T, you are creating more products for an endothermic reaction.

By increasing T, you are creating more reactants for an exothermic reaction

By increasing K, you are shifting toward the products.

By decreasing K, you are shifting toward the reactants.

Q vs K: Chemical Equilibrium Terminology

- The purpose of K is to tell you the concentrations at equilibrium.
- The purpose of Q is to tell you the concentrations at any given starting point.
- The purpose of comparing Q to K is to tell you how the reaction will proceed at your Q-concentrations.

1. $Q < K$: reaction moves forward toward equilibrium
2. $Q = K$: reaction is at equilibrium (just right)
3. $Q > K$: reaction moves backward toward equilibrium

Q vs K: Chemical Equilibrium Terminology

Consider the familiar reaction below:

$$H_2\ (g) + I_2\ (g) \rightleftharpoons 2HI\ (g)$$

- At equilibrium, you measure 1atm H_2, 1atm I_2, and 4atm HI.
- Show the calculation for K_p when:
  - You are at equilibrium
  - You add 2 atm HI
  - You add 1 atm I_2 and H_2

1. Q < K : "-x" goes on reactant side ; "+x" goes on product side
2. Q = K : reaction is at equilibrium (just right)
3. Q > K : "+x" goes on reactant side ; "-x" goes on reactant side

Q vs K: Chemical Equilibrium Terminology

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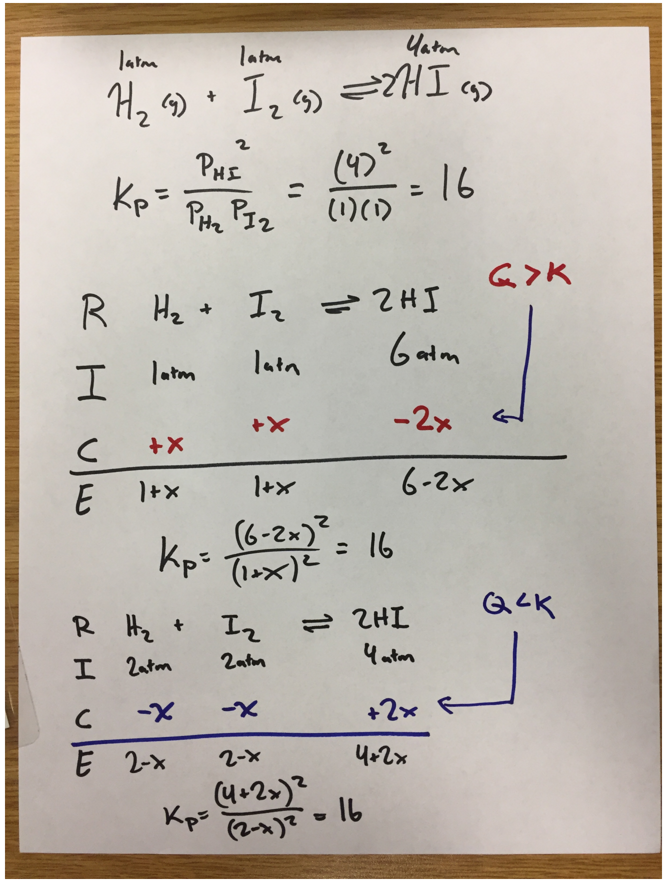


$$\overset{1atm}{H_{2\,(g)}} + \overset{1atm}{I_{2\,(g)}} \rightleftharpoons \overset{4atm}{2HI_{(g)}}$$

$$K_p = \frac{P_{HI}^{\;2}}{P_{H_2}\,P_{I_2}} = \frac{(4)^2}{(1)(1)} = 16$$

Q>K

R	H_2 +	I_2	$\rightleftharpoons$ 2HI
I	1atm	1atm	6atm
C	+x	+x	−2x
E	1+x	1+x	6-2x

$$K_p = \frac{(6-2x)^2}{(1+x)^2} = 16$$

Q<K

R	H_2 +	I_2	$\rightleftharpoons$ 2HI
I	2atm	2atm	4atm
C	−x	−x	+2x
E	2-x	2-x	4+2x

$$K_p = \frac{(4+2x)^2}{(2-x)^2} = 16$$

Le Chatelier’s Principle

1. Adding or removing species:
   - Adding: reaction shifts toward the opposite side of the addition
   - Removing: reaction shifts toward the same side as the removal
2. Increasing or decreasing the volume (gases):
   - Increasing volume (decreasing pressure): shifts toward the side with the most gas moles
   - Decreasing volume (increasing pressure): shifts toward the side with the least gas moles
3. Diluting or concentration a solution (aqueous):
   - Diluting: shifts toward the side of the most aqueous moles
   - Concentration: shifts toward the side of least aqueous moles
4. Changing the temperature:
   - Endothermic: increasing T shifts toward products ; decreasing T shifts toward reactants
   - Exothermic: increasing T shifts toward reactants ; increasing T shifts toward products

Le Chatelier's Principle: Temperature

- To simplify the relationship between K and temperature, we can think of temperature like a product or a reactant of a chemical reaction depending on whether the reaction is exothermic or endothermic.
- Endothermic reactions are driven by an input of heat; therefore, heat is a reactant. Increasing the heat is like adding a reactant. This shifts the equilibrium toward the products.

$$\boxed{\text{Heat + Reactants} \rightleftharpoons \text{Products}}$$

- Exothermic reactions have an output of heat; therefore, heat is a product. Increasing the heat is like adding a product. This shifts the equilibrium toward the reactants.

$$\boxed{\text{Reactants} \rightleftharpoons \text{Products + Heat}}$$

Acids and Bases

TYPE 1, TYPE 2 PROBLEMS; BUFFERS, TITRATIONS, POLYPROTICS, INDICATORS

Acids and Bases Fundamentals

- The study of acids and bases revolves around understanding the chemical environment of aqueous solutions associated with proton and hydroxide concentrations.
- The standard units of measurement for acids and bases are pH and pOH

$$pH = -\log[H^+] \qquad pOH = -\log[OH^-]$$

- Some things to note about this relationship:
  - Because this relationship is based on the negative log, **a high value of [H⁺] will have a low pH value**.
  - By using a logarithmic scale, you should understand that a difference between pH = 2 and pH = 7 is not a difference of 5, **but 5 orders of magnitude**.

A particularly important K value

K_w represents the auto-ionization of water; that is, it is the equilibrium constant for the following reaction at 298.15K:

- $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$
- $K_w = 1.0 \times 10^{-14} =$ **$[H^+][OH^-]$**
- **K_w represents the standard for our pH scale**
  - For a neutral solution, pH = 7
    - **Acidic solutions** have pH < 7
    - **Basic solutions** have pH > 7
  - Therefore, at K_w for a neutral solution the H^+ and OH^- concentrations are equal to 1.0×10^{-7}
  - K_w for a basic or acidic solution will have different values of H^+ and OH^-, but their product will still = 1.0×10^{-14}
- **K_w is important to our discussion because it holds its value no matter the pH**
  - **If we want to switch between pH and pOH, we can simply insert values into our K_w expression.**

Acid and Base Question Types (Simplified)

In addition to real world applications of acid/base problems, you should be able to put yourself into the mindset of a chemist doing these experiments in the following scenarios:

- Strong acid, strong base questions
  - Simple relationships converting $[H_3O^+]$ and $[OH^-]$ to pH, pOH
- Weak acid, weak base questions
  - Approximations or quadratic formula (if necessary) with your general formula:
  - $HA \rightleftharpoons H^+ + A^-$ OR $B \rightleftharpoons BH^+ + OH^-$

Type 1: one species dissolving in water

Acid and Base Question Types (Simplified)

In addition to real world applications of acid/base problems, you should be able to put yourself into the mindset of a chemist doing these experiments in the following scenarios:

Type 2: two species go into solution

- Buffer Questions (partial neutralizations)
  - Mixtures of a weak acid and its salt (conjugate base) ; weak base and its salt (conjugate acid)
  - Solve for pH, pOH using Henderson-Hasselbalch equation
- Neutralization reactions (Titration experiments ; real world application)
  - By adding a titrant to an analyte solution, you can neutralize your original species. Depending where you are in the neutralization experiment, you will make different calculations.

Quantifying Acids and Bases

- The standard units of measurement for acids and bases are pH and pOH

$$\mathrm{pH} = -\log[H^{+}] \qquad \mathrm{pOH} = -\log[OH^{-}]$$

- In a strong acid or strong base solution, we can use this relationship:

$$\mathrm{pH} = -\log[C_{A}] \qquad \mathrm{pOH} = -\log[C_{B}]^{*}$$

- pH and pOH can be interconverted using the relationship based on $\mathrm{pK_w}$ (14):

$$\mathrm{pH} = 14 - pOH \qquad \mathrm{pOH} = 14 - pH$$

Quantifying Weak Acids and Bases

- For a weak acid, we are dealing with a more interesting equilibrium (additional steps). The approximation formulas are below:

$$[H^+] = \sqrt{C_{HA} \cdot K_a} \qquad [OH^-] = \sqrt{C_B \cdot K_b}$$

- Don’t forget that at any time you can convert between different terms:

$$K_w = 1 \cdot 10^{-14} = [H^+][OH^-]$$

$$\begin{tikzcd} & \frac{K_w}{K_a} = K_b \\ K_w = K_a K_b \arrow[ur] \arrow[dr] & \\ & \frac{K_w}{K_b} = K_a \end{tikzcd}$$

Quantifying Weak Acids and Bases

- Important Reminder: K_a will get you $[H^+]$, K_b will get you $[OH^-]$. Therefore, K_a corresponds to a weak acid reaction and K_a corresponds to a weak base reaction.

$$HA \rightleftharpoons H^+ + A^- \longrightarrow K_a = \frac{[H^+][A^-]}{[HA]} \longrightarrow \mathrm{pH} = -\log[H^+]$$

(Solve using approximation or quadratic)

$$B \rightleftharpoons BH^+ + OH^- \longrightarrow K_b = \frac{[OH^-][BH^+]}{[B]} \longrightarrow \mathrm{pOH} = -\log[OH^-]$$

Another helpful relationship:

- Another way of explaining how much a weak acid or base dissolves is percent ionization. You can memorize the formula:

$$[H^{+}] = (\%\ \text{ionization})(\text{C}_{HA})$$

[H⁺] can be solved for using the approximation or quadratic formula. Rearrange to solve for % ionization.

If you have the % ionization, you can multiply by the concentration of your weak acid to get [H⁺].

$$[OH^{-}] = (\%\ \text{ionization})(\text{C}_{B})$$

Note that these are fractions, 0-1, so you need to multiply by 100 to get the “percent” value

Neutralization Reactions: Salts

- The product of a neutralization reaction is a salt. In acid/base chemistry, your salt can be neutral, acidic, or basic depending on the reaction:

1. GENERIC REACTION (very helpful):

   Acid(aq) + Base(aq) $\rightleftharpoons$ Salt(aq) + Water(l)

2. Strong Acid, Strong Base: results in a neutral salt

   $HCl(aq) + NaOH(aq) \rightleftharpoons NaCl(aq) + H_2O(l)$

3. Strong Acid to weak base: results in an acidic salt

   $HCl(aq) + NH_3(aq) \rightleftharpoons NH_4Cl\ (aq) + H_2O(l)$*

4. Strong base to weak acid: results in a basic salt

   $CH_3COOH(aq) + NaOH(aq) \rightleftharpoons NaCH_3COO(aq) + H_2O(l)$

Buffers: Conceptual, Calculation

These values can be in Molarity or moles.

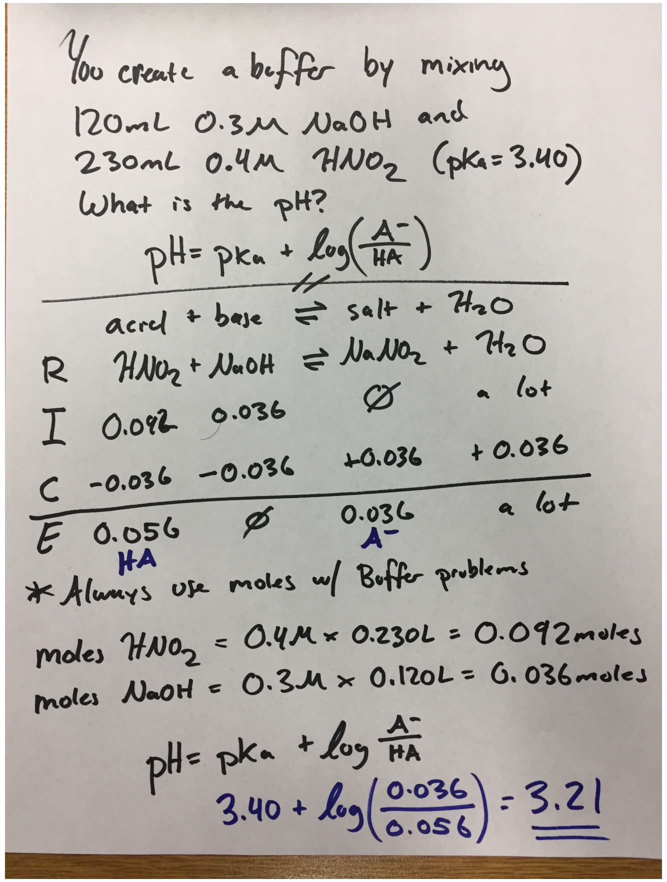
- **The purpose of a buffer is to resist changes in pH.**
- Here's the idea:
  - If you add one mole of NaOH to pure water, you are adding one mole of OH^-. This results in a pretty big change in pH
  - If you add one mole of NaOH to a solution of acetic acid and sodium acetate, you are just creating one more mole of sodium acetate. **This *barely* increases the pH.**
- **By definition, a buffer is made of:**
  - **A weak acid and its salt (conjugate base)**
  - **A weak base and its salt (conjugate acid)**
- By definition, a buffer is effective when:
  - The pH is within about +/- 1 of the pK_a or pK_b
  - This favorable ratio of the weak acid : salt (or weak base : salt) is known as the "Buffer Zone

$$\mathrm{pH} = \mathrm{pK}_a + \log\left(\frac{A^-}{HA}\right)$$

$$\mathrm{pOH} = \mathrm{pK}_b + \log\left(\frac{BH^+}{B}\right)$$

If you can identify that you have a buffer based on this definition, you know to use the Henderson-Hasselbalch equation (above)

Buffers: Conceptual, Calculation



You create a buffer by mixing
120mL 0.3M NaOH and
230mL 0.4M HNO_2 (pKa=3.40)
What is the pH?

$$pH = pka + \log\left(\frac{A^-}{HA}\right)$$

	acid	+ base	⇌	salt	+ H_2O
R	HNO_2	+ NaOH	⇌	$NaNO_2$	+ H_2O
I	0.092	0.036		∅	a lot
C	−0.036	−0.036		+0.036	+0.036
E	0.056 HA	∅		0.036 A^-	a lot

* Always use moles w/ Buffer problems

moles HNO_2 = 0.4M × 0.230L = 0.092 moles
moles NaOH = 0.3M × 0.120L = 0.036 moles

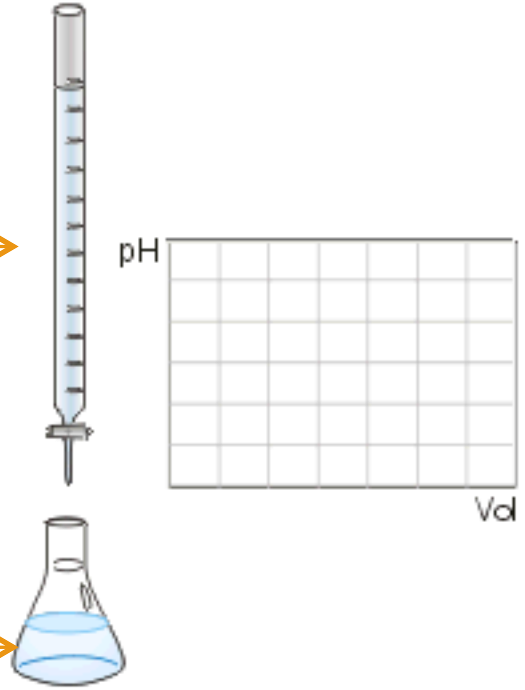
$$pH = pka + \log\frac{A^-}{HA}$$

$$3.40 + \log\left(\frac{0.036}{0.056}\right) = \underline{\underline{3.21}}$$

Titration Experiments: NaOH added to HCl

This is the titrant. For this experiment, the titrant is NaOH (strong base). It is added drop-wise to the analyte. We know the concentration of the titrant.

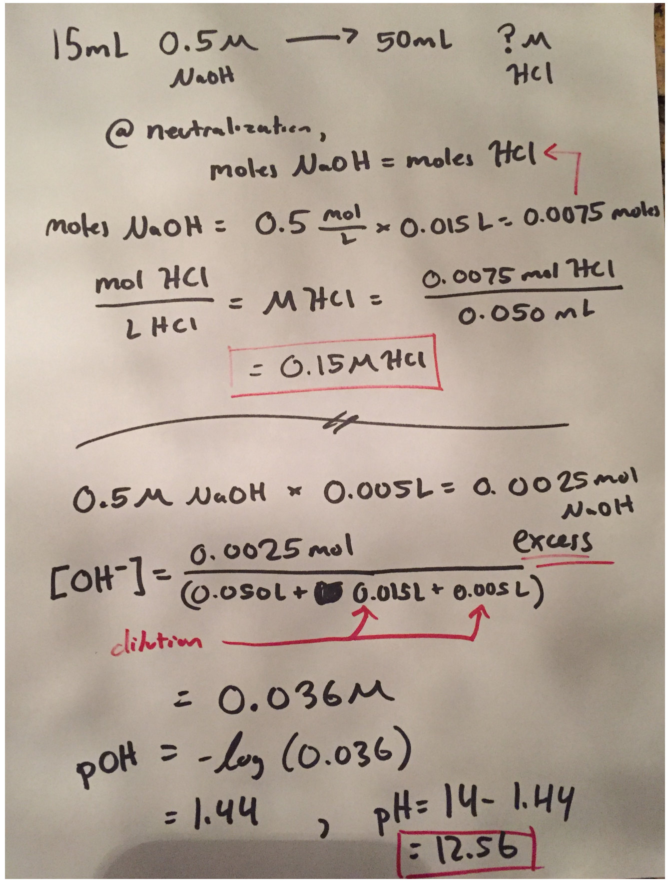
This is the analyte. For this experiment, the analyte is HCl (strong acid). We don't always know the concentration of the analyte. There is an indicator in the analyte that changes color at a specific pH.



Question: Suppose it took 15mL of 0.5M NaOH to completely neutralize 50mL of HCl analyte. What was your original HCl concentration? What is the pH of your sample if you then add 5mL MORE of NaOH?

The objective of a titration experiment is to fully neutralize the analyte solution. This will form a neutral, basic, or acidic salt, depending on the experiment.

Titration Experiments: NaOH added to HCl



$15\text{mL}\ \underset{\text{NaOH}}{0.5\text{M}} \longrightarrow 50\text{mL}\ \underset{\text{HCl}}{?\text{M}}$

@ neutralization,

moles NaOH = moles HCl

$\text{moles NaOH} = 0.5\,\frac{\text{mol}}{\text{L}} \times 0.015\,\text{L} = 0.0075\ \text{moles}$

$$\frac{\text{mol HCl}}{\text{L HCl}} = \text{M HCl} = \frac{0.0075\ \text{mol HCl}}{0.050\ \text{mL}}$$

$$\boxed{= 0.15\text{M HCl}}$$

$0.5\text{M NaOH} \times 0.005\text{L} = 0.0025\ \text{mol NaOH}$

$$[OH^-] = \frac{0.0025\ \text{mol}}{(0.050\text{L} + 0.015\text{L} + 0.005\text{L})}$$

excess

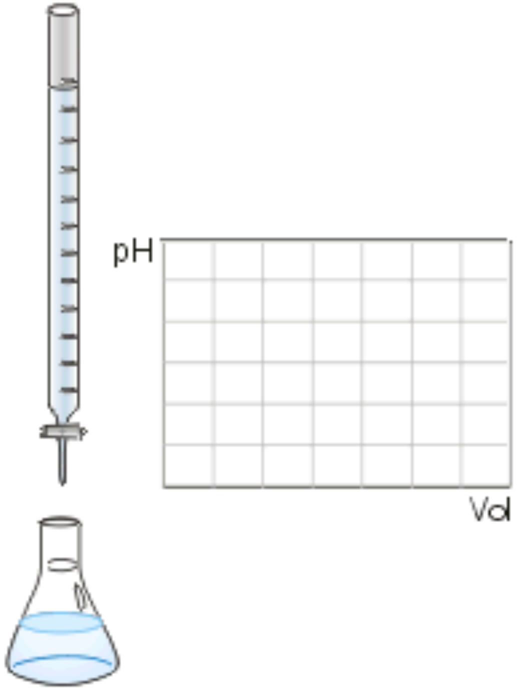
dilution

$= 0.036\text{M}$

$\text{pOH} = -\log(0.036)$

$= 1.44$, $\text{pH} = 14 - 1.44$

$\boxed{= 12.56}$



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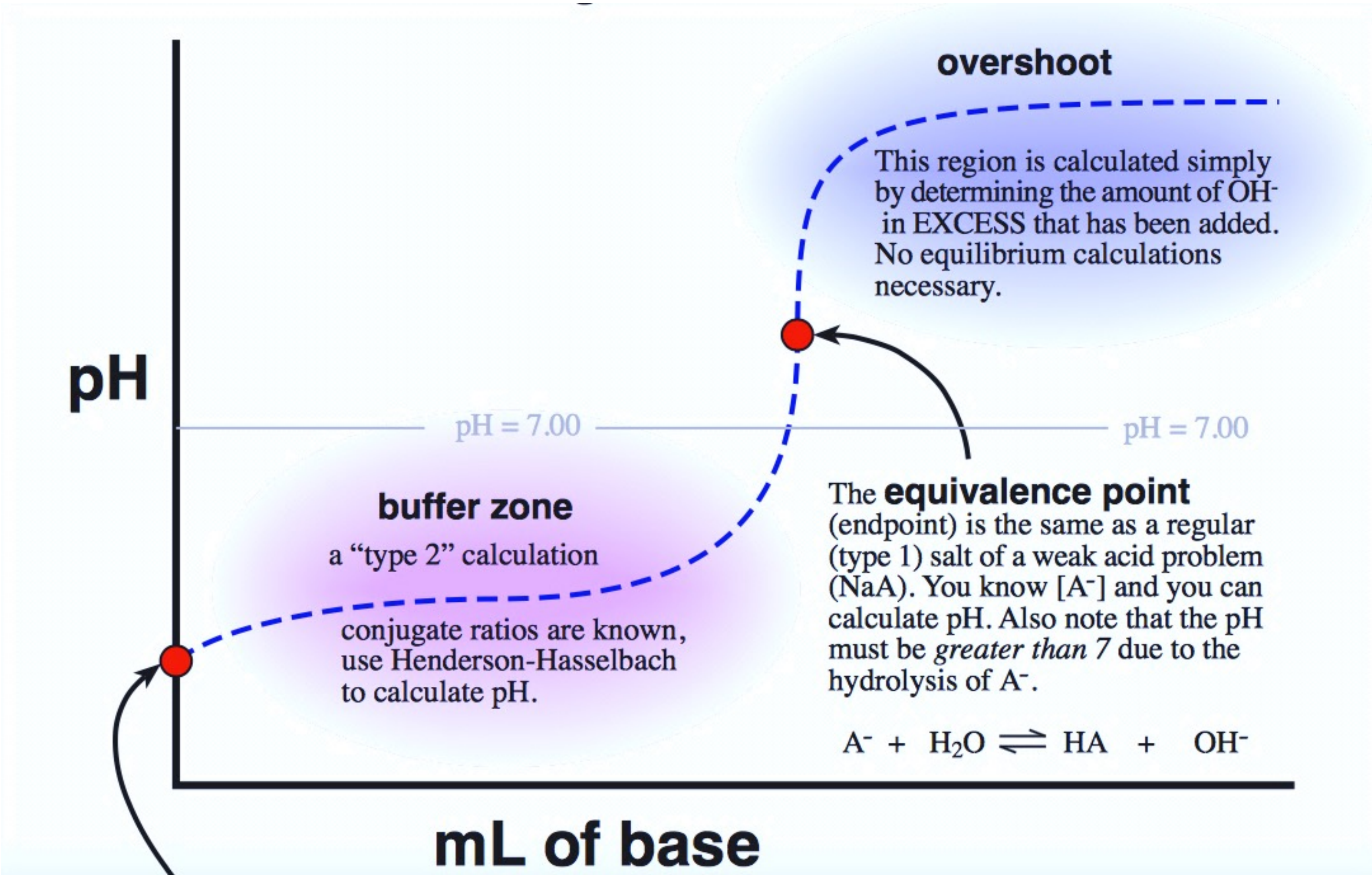
Titrations of Buffered Solutions



This region is calculated simply by determining the amount of OH^- in EXCESS that has been added. No equilibrium calculations necessary.

conjugate ratios are known, use Henderson-Hasselbach to calculate pH.

The **equivalence point** (endpoint) is the same as a regular (type 1) salt of a weak acid problem (NaA). You know $[A^-]$ and you can calculate pH. Also note that the pH must be *greater than 7* due to the hydrolysis of A^-.

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

A) Before you have started:

Calculate pH using the approximation formula or using $K_a = [H^+][A^-] / [HA]$.

B) Within the buffer zone:

$pH = pKa + \log([A^-]/[HA])$

C) At the equivalence point:

You only have the salt of your weak acid in solution. Use $K_w = K_aK_b$, then solve for $[OH^-]$, pOH, then convert to pH

D) After the equivalence point:

Find the pH using the excess base in solution

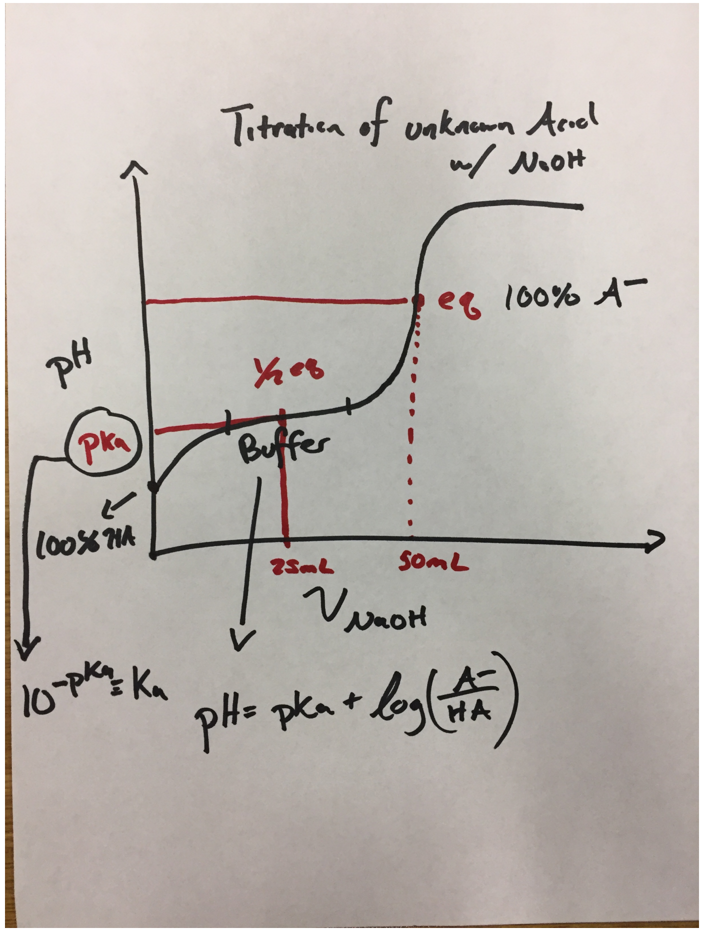
$pOH = -\log([\text{excess base}])$

$pH = 14 - pOH$

Points to know:

- Equivalence point
- End Point
- Half-equivalence point
- Buffer Zone

Titrations of Buffered Solutions



A) Before you have started:
Calculate pH using the approximation formula or using $K_a = [H^+][A^-] / [HA]$.

B) Within the buffer zone:
$pH = pKa + \log([A^-]/[HA])$

C) At the equivalence point:
You only have the salt of your weak acid in solution. Use $K_w = K_aK_b$, then solve for $[OH^-]$, pOH, then convert to pH

D) After the equivalence point:
Find the pH using the excess base in solution
pOH = -log ([excess base])
pH = 14- pOH

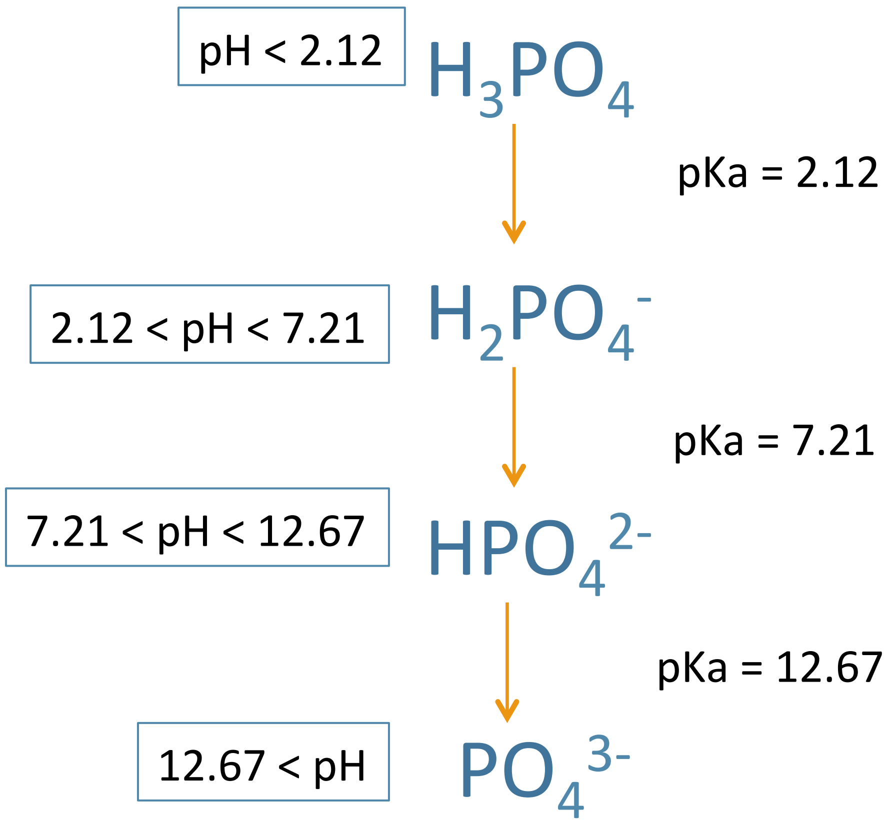
Points to know:

- Equivalence point
- End Point
- Half-equivalence point
- Buffer Zone

Fraction of Species is an important tool here

Dominant Species Fundamentals

- **For polyprotic acids, two simple rules go a long way when identifying the dominant species in solution:**
  1. **If pH < pK_a** : you have too many hydrogen ions in solution for it to be favorable for a hydrogen ion on your molecule to "pop off"
     - Think about it: when the environment is more acidic, it is more "proton-donating,"
     - Your molecule will be protonated
  2. **If pH > pK_a** : the environment around your molecule is thirsty for hydrogen ions, so it is favorable for a hydrogen ion on your molecule to pop off
     - Think about it: when the environment is more basic, it is more "proton-accepting."
     - Your molecule will be deprotonated



pH < 2.12 H_3PO_4

pKa = 2.12

2.12 < pH < 7.21 $H_2PO_4^-$

pKa = 7.21

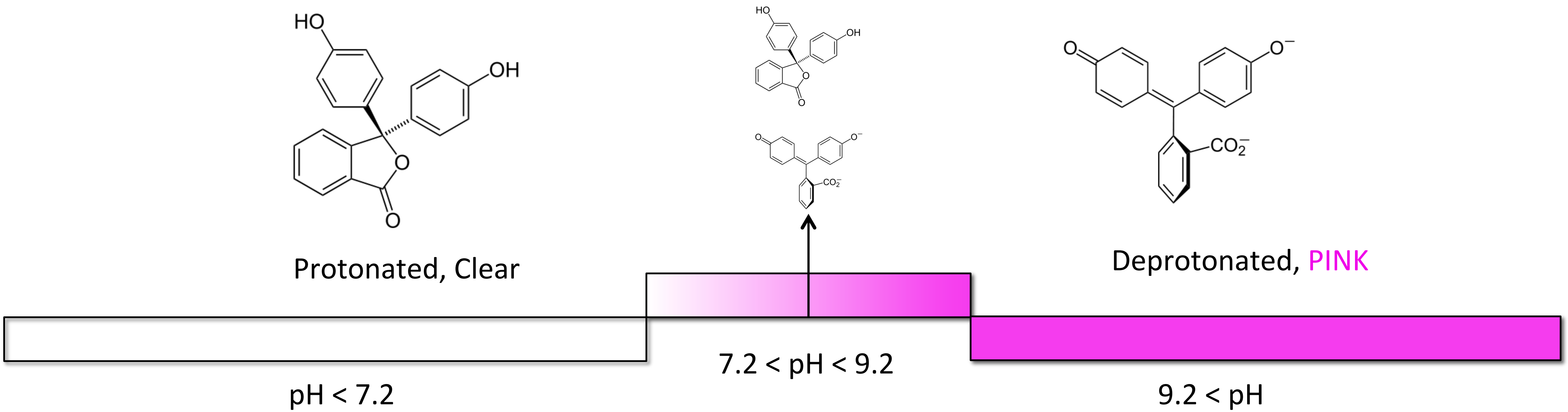
7.21 < pH < 12.67 HPO_4^{2-}

pKa = 12.67

12.67 < pH PO_4^{3-}

Dominant Species: Indicators

- The purpose of an indicator is to change color in a particular pH range (color change occurs +/- 1 pH unit from the pKa). The protonated and deprotonated forms of the molecule will have different colors. **Therefore, the color of solution will represent the dominant species in solution.**
- Consider phenolphthalein, pKa = 8.2:



Protonated, Clear

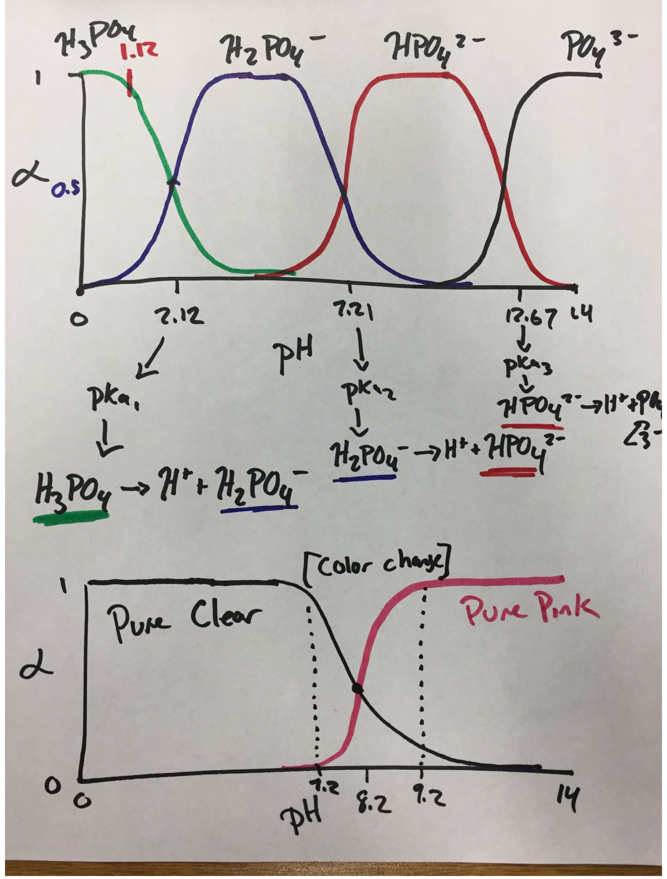
Deprotonated, PINK

pH < 7.2

7.2 < pH < 9.2

9.2 < pH

Dominant Species: Indicators



CH 302 – Exam Review

ADDITIONAL SLIDES

Le Chatelier’s Principle: Adding/Removing Stuff

- By adding or removing product or reactant, you are manipulating Q
- When you manipulate Q, the reaction “shifts” to get you back to K

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$$

$$K_p = \frac{(P_{HCl})^2}{P_{H_2} P_{Cl_2}}$$

Add product →

Q becomes greater than K, reaction shifts toward the reactants

Add reactant →

Q becomes less than K, reaction shifts toward the products

Le Chatelier’s Principle: Adding/Removing Stuff

- By adding or removing product or reactant, you are manipulating Q
- When you manipulate Q, the reaction “shifts” to get you back to K

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$$

$$K_p = \frac{(P_{HCl})^2}{P_{H_2} P_{Cl_2}}$$

Remove product → Q becomes less than K, reaction shifts toward the products

Remove reactant → Q becomes greater than K, reaction shifts toward the reactants

Le Chatelier’s Principle: Volume and Pressure

- By changing the volume or pressure, you can manipulate Q for a gaseous system.
- The reaction shifts based on the number of gas species in the products or reactants

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$$

$$K_p = \frac{(P_{NH_3})^2}{(P_{H_2})^3 P_{N_2}} \xrightarrow[\text{Increase the pressure}]{\text{Reduce the volume}}$$

If you increase the pressure, you are causing the most stress on the side of the reaction with the most gas species. **The reaction will shift toward the side of the least gas species.**

Le Chatelier’s Principle: Volume and Pressure

- By changing the volume or pressure, you can manipulate Q for a gaseous system.
- The reaction shifts based on the number of gas species in the products or reactants

$$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$$

$$K_p = \frac{(P_{NH_3})^2}{(P_{H_2})^3 P_{N_2}}$$

Increase the volume

$$\longrightarrow$$

Decrease the pressure

If you decrease the pressure, you are causing the most stress on the side of the reaction with the least gas species. **The reaction will shift toward the side of the most gas species.**

Le Chatelier's Principle: Concentration

- Based on the same principle, changing the concentration is changing the volume of the solvent (as opposed to the volume of the container in the previous example).
- The reaction shifts based on the number of aqueous species in the products or reactants

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

$$K_A = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

Concentrate the solution →

If you concentrate the solution, you are causing the most stress on the side with the most aqueous species. **The reaction will shift in toward the side with the least aqueous species.**

Le Chatelier's Principle: Concentration

- Based on the same principle, changing the concentration is changing the volume of the solvent (as opposed to the volume of the container in the previous example).
- The reaction shifts based on the number of aqueous species in the products or reactants

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

$$K_A = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

Dilute the solution →

If you dilute the solution, you are causing the most stress on the side with the least aqueous species. **The reaction will shift toward the side of the most aqueous species.**