

CHEM 152 ALEKs Scratch Work

Create ICE Table				
X	$[CH_4] + [H_2O] \rightleftharpoons [CO] + [3H_2]$			
I	0	2.2	0.8	3.6
C		x		
E				

X	$[H_2] + [F_2] \rightleftharpoons [2HF]$		
I	1	2	0
C	-x	-x	2x
E	1 - x	2 - x	2x

Interconverting pH and hydronium ion concentration

Given $[H_3O^+]$ or pH, you can convert between the two using:

$$pH = -\log\left(\frac{[H_3O^+]}{1\text{mol/L}}\right) \quad [H_3O^+] = (10^{-pH}) \frac{\text{mol}}{L}$$

$$A_2 + 3 B_2 \rightleftharpoons 2 AB_3 \quad K_1^2 = \frac{[AB_3]^2}{[A_2][B_2]^3}$$

$$2 AB_3 \rightleftharpoons 2 AB + 2 B_2 \quad K_2 = \frac{[AB]^2[B_2]^2}{[AB_3]^2}$$

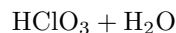
$$A_2 + B_2 \rightleftharpoons 2 AB \quad K = \frac{[AB]^2}{[A_2][B_2]} = K_1 K_2$$

$$K = \frac{[AB_3]^2}{[A_2][B_2]^3} * \frac{[AB]^2[B_2]^2}{[AB_3]^2}$$

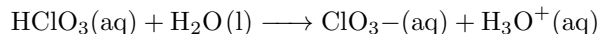
$$K = \frac{[AB]^2}{[A_2][B_2]}$$

Predicting reaction products of strong acid with water

Given a reaction:



1. Firstly, recognize that $HClO_3$ is an acid (which the H at the front in . indictative that this is an acid).
2. You know that acids donate an hydrogen cations H^+ to water, producing hydronium cations H_3O^+ and anions. In this case, chloric acid reacts with water like this:



3. The reaction of acids is sometimes called a **proton transfer reaction**, since the hydrogen cation that gets tranfered, a hydrogen atom without its electron, is usually just a proton.

Identifying Bronsted-Lowry Acids and Bases

A compound acts as a Bronsted-Lowry acid when it **donates** hydrogen cation during a chemical reaction, and acts as a Bronsted-Lowry base when it **accepts** a hydrogen cation.

Definition 1 *Bronsted Lowry Acid/Bases*

1. *Acids* - Compounds that **donates** hydrogen cations (H^+) during a chemical reaction.
2. *Bases* - Compounds that **accepts** hydrogen cations (H^+) during a chemical reaction.
3. *NOTE* - Bronsted Lowry acids and bases always come in pairs.
4. *ALSO* - Compounds are only defined as Bronsted Lowry acids or bases in the context of a specific chemical reaction. A compound can be a Bronsted-Lowry acid in 1 reaction, and act as a Bronsted-Lowry base in another reaction.

Calculating an equilibrium constant from an equilibrium composition

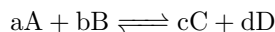
In order to find the equilibrium constant K_P , you'll need to write the equilibrium constant expression. Remember for K_P :

$$K_P = \frac{\left(\frac{P_C}{P_0}\right)^c \left(\frac{P_D}{P_0}\right)^d}{\left(\frac{P_A}{P_0}\right)^a \left(\frac{P_B}{P_0}\right)^b}$$

Plugging in $P_0 = 1$ atm and simplifying:

$$K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

Then for a reaction of form:



You would get an equilibrium constant of form:

$$K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

Example: Find K_P for a reaction $2SO_3 \longrightarrow 2SO_2 + O_2$:

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

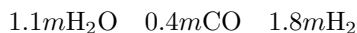
Writing a pressure equilibrium constant expression

You know how to do this normally from the notes you've taken during class. Doing examples:

$$K_P = \frac{[PBr_3]^4}{[P_4][Br_2]^6}$$

Setting up a reaction table

Example - Given a reaction $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$, in a 500ml flask filled with:



Complete the table below, using x as the unknown change in molarity of H_2O :

X	$[CH_4] + [H_2O] \rightleftharpoons [CO] + [3H_2]$
I	
C	x
E	

In order to solve this, you can use the following five-step plan to write a reaction table for any chemical reaction:

- Be sure the chemical equation is balanced.
 - The chemical reaction must be balanced since you need to know the right stoichiometric coefficients to set up the reaction table. In this case, the reaction is already balanced.
- Find the initial molarity of each reactant.
 - In this case, you're not given the molarity, but the moles of each reactant and the volume of the reaction flask, allowing you to find the initial molarity of each reactant. Like for finding H_2O :

$$[\text{H}_2\text{O}] = \frac{1.1 \text{ mol}}{500. \text{ mL}} = \frac{1.1 \text{ mol}}{0.5 \text{ L}} = 2.2 \frac{\text{mol}}{\text{L}} = 2.2 \text{ M}$$

Applying this to the rest of the compounds:

X	$[\text{CH}_4] + [\text{H}_2\text{O}] \rightleftharpoons [\text{CO}] + [3\text{H}_2]$			
I	0	2.2	0.8	3.6
C		x		
E				

- Write a variable expression for the unknown change in molarity of each reactant.
 - Since you don't know the what the change in molarity of each reactant will be, you must write them as algebra expression using a variable (like x). Remember that you must pick one of the changes to be x (ex change in H_2O), and then write all other changes to be in terms of x.
 - From this, you can write the change of molarity for CH_4 in terms of x like:

$$\frac{\text{moles/liter CH}_4 \text{ consumed}}{\text{moles/liter H}_2\text{O consumed}} = \frac{1}{1}$$

- Also to note, if x represents the change in the molarity of a **product**, then all the changes of **reactants** should have a minus since. Since if x is positive (the molarity of products is increasing) the forward reaction must be dominating, and that must mean the molarity of the reactants must be decreasing. Applying these previous two points:

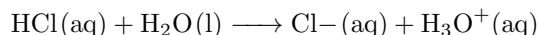
X	$[\text{CH}_4] + [\text{H}_2\text{O}] \rightleftharpoons [\text{CO}] + [3\text{H}_2]$			
I	0	2.2	0.8	3.6
C	x	x	-x	-3x
E				

- Add the first and second row to get the third.
 - The final molarity will be the sum of the initial molarity and the changes in molarity.

X	$[\text{CH}_4] + [\text{H}_2\text{O}] \rightleftharpoons [\text{CO}] + [3\text{H}_2]$			
I	0	2.2	0.8	3.6
C	x	x	-x	-3x
E	x	2.2 + x	0.8-x	3.6 - 3x

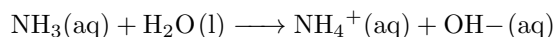
Finding the conjugate of an acid or base

The conjugate base of a Bronsted-Lowry acid is the product of the chemical reaction of the acid with water. For example, here's the reaction of hydrochloric acid HCl^- with water:



In this reaction, HCl acts as a Bronsted-Lowry acid by donating an H^+ to H_2O . As a result, the HCl turns into Cl^- , meaning Cl^- is the conjugate base of HCl .

Similarly, the conjugate acid of a Bronsted-Lowry base is the product of the chemical reaction of the base with water. For example, here is the reaction of ammonia NH_3 with water:



In this reaction NH_3 acts as a Bronsted-Lowry base by accepting an H^+ from H_2O . As a result, the NH_3 turns into NH_4^+ , meaning NH_4^+ is the conjugate acid of NH_3 .

Another way to look at conjugate acids and bases:

1. To find the conjugate base of an acid, remove one H^+ from its chemical formula. (AKA remove 1 H and subtract 1 from its charge)
2. To find the conjugate acid of a base, add one H^+ to its chemical formula. (AKA add 1 H and add 1 to its charge)

Using Le Chatelier's Principle to predict the result of changing concentration

There are two ways a mixture in chemical equilibrium can respond to a perturbation ("disturbance"):

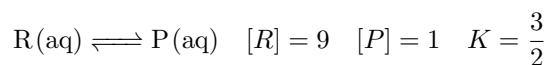
1. The rate of the forward reaction can temporarily become higher than the rate of the reverse reaction. That would turn a small amount of reactants into products. When this happens, chemists say the equilibrium has shifted to the right.
2. The rate of the reverse reaction can temporarily become higher than the rate of the forward reaction. That would turn a small amount of reactants into products. When this happens, chemists say the equilibrium has shifted to the left.

Le Chatelier's Principle Rules of Thumb - Generally,

1. If the reaction becomes BIASED towards the forward reaction, the reaction is **"shifted to the right"**.
2. If the reaction becomes BIASED towards the reverse reaction, the reaction is **"shifted to the left"**.

Predicting Equilibrium Composition From a Sketch

Given:



Find the number of R and P molecules in the sample when the reaction reaches equilibrium.

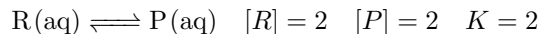
$$K = \frac{3}{2} = \frac{P}{R}$$

Using an ICE table to find P and R and getting an equation for K:

$$\begin{aligned} K &= \frac{3}{2} = \frac{P}{R} = \frac{1-x}{9+x} \\ \frac{3}{2} &= \frac{1-x}{9+x} \\ \frac{3}{2} * (9+x) &= 1-x \\ 13.5 + 1.5x &= 1-x \\ 13.5 - 1 &= -x - 1.5x \\ 12.5 &= -2.5x \\ x &= -5 \end{aligned}$$

In this case, I set x to be in terms of R, so since $x = -5$, 5 R molecules must turn into P molecules to be in equilibrium

Given:



Find the number of R and P molecules in the sample when the reaction reaches equilibrium.

$$\begin{aligned} K = 2 &= \frac{P}{R} = \frac{10 - x}{2 + x} \\ 2 &= \frac{10 - x}{2 + x} \\ 2 * (2 + x) &= 10 - x \\ 4 + 2x &= 10 - x \\ 4 - 10 &= -3x \\ -6 &= -3x \\ x &= 2 \end{aligned}$$

In this case, I set x to be in terms of R, so since $x = 2$, 2 P molecules must turn into R molecules to be in equilibrium

Identifying Strong and Weak acids/bases

Chemical reactions that happen when an acid or base dissolves in water:

1. An acid reacts in water to produce hydronium H_3O^+ cations and the acids conjugate base A^-

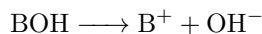


Strong acids react to completion, **so there is no HA left after the acid reacts.** For weak acids the reaction kinda looks like this:

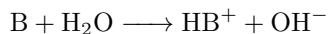


2. An base reacts in water to produce hydroxide OH^- anions and the bases conjugate acid B^+ .

- (a) A strong base BOH dissociates completely into its conjugate acid B^+ and hydroxide OH^- anions:



- (b) A weak base B reacts with water to form its conjugate acid HB^+ and hydroxide anions:



- 3.

Using Le Chatelier's Principle to predict the result of changing temperature

Following up on the previous section regarding Le Chatelier's Principle:

When the perturbation is a change in temperature, you must look at the feat of reaction to correctly decide how the system will respond. Keep in mind:

1. A **negative** heat of reaction AKA exothermic means the forward reaction releases heat,
2. A **positive** heat of reaction AKA endothermic means the forward reaction absorbs heat.

3. Continuing with this logic, if the forward reaction releases heat, the reverse reaction must absorb it and vice versa.

For example with an exothermic reaction where perturbation lowers the temperature, heat is removed from the reaction vessel. In order to counteract this, the effect of the perturbation can be reduced by changing reactants into products to create heat. In other words, if the forward reaction takes place in an exothermic reaction, some of the removed thermal energy will be replaced by the reaction.

Where a perturbation causes thermal energy to change, the forward or reverse reaction will occur in order to replace some of the removed/added thermal energy AKA change in thermal energy.

For example, with an exothermic reaction - if a perturbation raises the temperature, you are biasing the reverse reaction. AKA change products into reactants to remove heat.

For example, with an exothermic reaction - if a perturbation lowers the temperature, you are biasing the forward reaction. AKA change reactants into products to add heat.

Ordering acids by their strength

(I didn't have time to write down the ALEKS explanation) You can order them by finding the K_a value for each acid, where:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

Interconverting hydronium and hydroxide concentration at 25°C

The key to this problem is understanding the close connection between hydronium cation (H_3O^+) molarity and hydroxide anion (OH^-) molarity in an aqueous solution. The connection between the two can be written mathematically, like this:

$$[\text{H}_3\text{O}^+] \cdot [\text{OH}^-] = K_w$$

This equation is the ion product of water, and the constant K_w is called the *ion product constant* or *dissociation constant* of water. At temperature of about 25°C,

$$K_w = 1.0 \cdot 10^{-14} \text{M}^2$$

However, at temperatures higher than 25°C, K_w is higher, and at temperatures lower than 25 °C, K_w is lower. Remember that K_w is a measurement - and carries measurement uncertainty (AKA the significant figures present in K_w should be reflected in your final answer)

Finally, in order to find either $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$, use the ion product of water to solve for your unknowns. For example, to solve for $[\text{H}_3\text{O}^+]$:

$$\begin{aligned} [\text{H}_3\text{O}^+] \cdot [\text{OH}^-] &= K_w \\ [\text{H}_3\text{O}^+] &= \frac{K_w}{[\text{OH}^-]} \\ [\text{H}_3\text{O}^+] &= \frac{1.0 \cdot 10^{-14} \text{M}^2}{[\text{OH}^-]} \end{aligned}$$

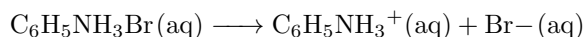
Predicting the qualitative acid-base properties of salts

...That you're asked about pH should suggest that some of these compounds may act as acids or bases when dissolved in water. In fact, some do.

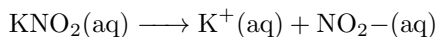
All of the compounds in this problem are salts, and will therefore separate into cations and anions as soon as they dissolve.

For example:

1. A compound $\text{C}_6\text{H}_5\text{NH}_3\text{Br}$ breaks up like this:



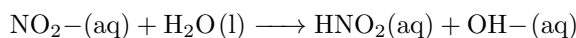
2. And another compound KNO_2 breaks up like this:



What you must ask yourself in each case is whether the cations or anions might act as a Bronsted-Lowry acid, Bronsted-Lowry base, or neither.

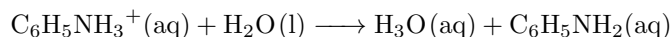
There are 4 important facts to keep in mind:

1. The conjugate base of a weak acid acts as a weak base.
For example, the data given in the question tells you that NO_2^- is the conjugate base of a weak acid HNO_2 , and will therefore act as a weak Bronsted-Lowry base:



In this reaction, NO_2^- accepts protons from water, producing hydroxide (OH^-) anions and raising the pH of the solution - AKA making the solution basic.

2. The weaker a weak acid, the stronger its conjugate base.
You might say that the more reluctantly an acid donates its proton to water, the more eagerly its conjugate base tries to get the proton back. This means for example, since the acid dissociation constant K_a of HClO is lower than the K_a of HNO_2 , ClO^- is a stronger base than NO_2^- . That is, a solution containing ClO^- will have a higher pH than a solution containing the same molarity of NO_2^- .¹
3. The conjugate acid of a weak base acts as a weak acid.
For example, the data given in the question that $\text{C}_6\text{H}_5\text{NH}_3^+$ is the conjugate acid of a weak base ($\text{C}_6\text{H}_5\text{NH}_2$) and will therefore act as a weak Bronsted-Lowry acid:



In this reaction $\text{C}_6\text{H}_5\text{NH}_3^+$ donates protons to water, producing hydronium cations and lowering the pH of the solution - AKA making the solution acidic

4. The weaker a weak base, the stronger its conjugate acid.
You might say the more reluctantly a base accepts a proton from water, the more eagerly its conjugate acid tries to give the proton back.

Calculating the pH of a weak acid solution

To calculate the pH of a weak acid solution - there are two key points to solving the problem:

1. Like all acids, when acids dissolve, they react with water by transferring H^+
 $\text{HCN} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CN}^-$
2. HCN is a weak acid, its reaction with water will **not** go to completion. Instead, the equilibrium molarities of reactants and products will be determined by the acid dissociation constant equation.
3. Combining these previous two points

¹I'm having a little bit of trouble conceptualizing and connecting how one chemical can have a higher pH than another. Am I supposed to take your word that one chemical is weaker than another, then based off of this "tier list" we can say that this is true?

Writing the dissociation reactions of a polyprotic acid

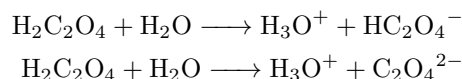
Example 1 Oxalic acid $H_2C_2O_4$ is a polyprotic acid. Write balanced chemical equations that oxalic acid can undergo when its dissolved in water.

A polyprotic acid is an acid that has more than 1 acidic hydrogen. In this case, carbonic acid has two, and they are written in the front of the chemical formula.

If you're told a compound is a polyprotic acid and you see that the formula starts off with several H's, it's a reasonable assumption at the general chemistry level that these are the acidic hydrogens.

However, you should be aware that with organic polyprotic acids in particular the acidic hydrogens are usually **not** written at the front of the chemical formula. You will need a more advanced chemical intuition to deduce the number of acidic hydrogens in such cases.

Each acidic hydrogen can be donated to a water molecule to form a hydronium cation. Since there are two acidic hydrogens in oxalic acid, this can happen twice:

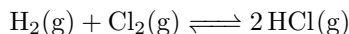


To predict the result from each reaction, you take an H^+ molecule away from the acid molecule and add it to an H_2O , forming hydronium. That's one product. The other product, the conjugate base of the acid, is the acid molecule without its H^+ .

Each time you take away an H^+ , be careful to adjust the charge on the conjugate base left behind. Remember that removing a charge of +1 from a neutral object leaves behind a -1 charge.

Calculating equilibrium composition from an equilibrium constant

Example 2 Suppose a 250. mL flask is filled with 1.4 mol of H_2 and 1.3 mol of Cl_2 . The following reaction becomes possible:



The equilibrium constant K for this reaction is 8.50 at the temperature of the flask. Calculate the equilibrium molarity of H_2 .

The key to solving an equilibrium composition problem is the connection between the equilibrium molarities of each reactant and the equilibrium constant K :

$$\frac{[HCl]^2}{[H_2][Cl_2]} = K$$

You can use this equation to find K from the equilibrium molarities. But if you know K instead, you can use the equation "backwards" to calculate the equilibrium molarities.

Setting up an ICE table:

X	$[H_2]$	$+ [Cl_2]$	$\rightleftharpoons [2HCl]$
I	5.6	5.2	0
C	-x	-x	2x
E	5.6 - x	5.2 - x	2x

Substituting last row expressions into molarities for equilibrium constant expression:

$$\begin{aligned} \frac{(2x)^2}{(5.6 - x)(5.2 - x)} &= 8.50 \\ (2x)^2 &= 8.50 * (5.6 - x)(5.2 - x) \\ 4x^2 &= 247.52 - 91.8x + 8.5x^2 \\ -4.5x^2 + 91.8x - 247.52 &= 0 \\ x &= \frac{-91.8 \pm \sqrt{91.8^2 - 4 * (-4.5) * (-247.52)}}{2 * (-4.5)} \\ x &= 3.1975, 17.2025 \end{aligned}$$

Although the quadratic formula yields two values for x , only one is physically reasonable - only molarities can give positive values can be used as the true value for x . This means that x can only equal 3.1975. Going back and calculating H_2 :

$$\begin{aligned}[H_2] &= 5.6 - x = 5.6 - 3.1975 \\ [H_2] &= 2.40\text{ M}\end{aligned}$$

Diluting a strong acid solution to a given pH

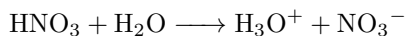
Example 3 *A chemist must prepare 200.0 mL of nitric acid solution with a pH of 1.10 at 25°C. He must do this in three steps:*

1. *Fill 200mL flask halfway with distilled water.*
2. *Measure out a small volume of concentrated (7.0M) stock nitric acid solution and add it to the flask.*
3. *Fill the flask to the mark with distilled water.*

Calculate the volume of concentrated nitric acid that the chemist must measure out in the second step.

The key to solving this problem is two facts:

1. First, since nitric acid (HNO_3) is an acid, it chemically reacts with water as soon as it dissolves like this.



Also, since HNO_3 is a strong acid, all of the HNO_3 reacted with the water. That is, the 7.0M stock solution of nitric acid doesn't contain any nitric acid molecules. It's really a 7.0M solution of H_3O^+ cations and NO_3^- anions.

2. Second, the pH of an aqueous solution is just another way to express $[H_3O^+]$ the molarity of hydronium. So when you're told the final pH of the solution should be 1.10, you're actually being told the final molarity of H_3O^+ should be:

$$[H_3O^+] = 10^{-pH} M = 10^{-1.10} M = 0.07943 M$$

What these two facts tell you is that this problem is really a dilution problem. In essence, you're being asked to figure out how much 7.0M H_3O^+ solution must be poured out so that when it's diluted to 200mL, the molarity of H_3O^+ falls to 0.07943M.

You can solve this dilution problem by using the key fact about all dilution problems: the number of moles of solute always stays the same. That is, the number of moles of H_3O^+ in the final solution must equal the number of moles of H_3O^+ in the concentrated solution the chemist pours out.

In the final solution, the molarity of H_3O^+ is supposed to be 0.07943M. You can find the total moles of H_3O^+ in the final solution by multiplying this molarity by the total volume of the final solution:

$$0.07943 M * 200 mL = 0.07943 M * 0.2 L = 0.01589 mol$$

To get the amount of solution needed to get the 0.01589mol of H_3O^+ , divide the number of moles needed by the molarity of H_3O^+ in the concentrated solution:

$$\begin{aligned}\frac{0.01589 mol}{7.0 M} &= 0.002270 L \\ 0.002270 L &= 2.270 * 10^{-3} L \\ &= 2.270 mL\end{aligned}$$

Preparing a strong base solution with a given pH

Example 4 A chemist must prepare 600.0 mL of sodium hydroxide solution with a pH of 13.70 at 25°C. He must do this in three steps:

1. Fill 600mL flask halfway with distilled water.
2. Weigh out a small amount of solid sodium hydroxide and add it to the flask.
3. Fill the flask to the mark with distilled water.

Calculate the mass of solid sodium hydroxide that the chemist must measure out in the second step.

The key to solving this problem is two facts:

1. First, the pH of an aqueous solution is just another way to express $[\text{H}_3\text{O}^+]$ the molarity of hydronium. So when you're told the final pH of the solution should be 13.70, you're actually being told the final molarity of H_3O^+ should be:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} M = 10^{-13.70} M = 1.995 \times 10^{-14} M$$

2. Secondly, sodium hydroxide (NaOH) is an electrolyte. That means that is NaOH dissolves it breaks up into Na^+ cations and OH^- anions. So by adding the correct amount of NaOH, the chemist can set $[\text{OH}^-]$ to whatever he wants. Why is this helpful? Because in an aqueous solution, the concentration of hydronium and hydroxide are related by the ion product of water like this:

$$[\text{H}_3\text{O}^+] \cdot [\text{OH}^-] = K_w$$

In other words, once $[\text{H}_3\text{O}^+]$ is found, the chemist can find the exact $[\text{OH}^-]$ needed to make $[\text{H}_3\text{O}^+]$ equal to $1.995 \times 10^{-14} M$.

$$\begin{aligned} [\text{H}_3\text{O}^+] \cdot [\text{OH}^-] &= K_w \\ [\text{OH}^-] &= \frac{K_w}{[\text{H}_3\text{O}^+]} \\ [\text{OH}^-] &= \frac{1.01 \cdot 10^{-14}}{1.995 \cdot 10^{-14}} \\ [\text{OH}^-] &= 0.5063 M \end{aligned}$$

In the final solution, the molarity of OH^- is supposed to be 0.5063M. You can find the total moles of OH^- in the final solution by multiplying this molarity by the total volume of the final solution:

$$0.5063 M \cdot 600 \text{ mL} = 0.5063 M \cdot 0.6 L = 0.3038 \text{ mol}$$

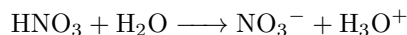
Since each formula unit of NaOH releases one OH^- when it dissolves, the moles of NaOH the chemist needs is equal to the number of moles of OH^- he needs. **To find the mass of NaOH he needs, you need to multiply the moles needed by the molar mass of NaOH:**

$$\begin{aligned} 0.3038 \text{ mol} \cdot 39.9970 \frac{\text{g}}{\text{mol}} \\ = 12.15 \text{ g} \end{aligned}$$

Calculating the pH of a strong acid solution

Example 5 A chemist dissolves 159. mg of pure nitric acid in enough water to make up 90. mL of solution. Calculate the pH of the solution.

As soon as the nitric acid HNO_3 , it will chemically react with the water to make hydronium cations and nitric anions like this:



Since HNO_3 is a strong acid, all of the HNO_3 will react. From this you can find the final molarity of HNO_3 in 4 steps:

1. Calculate the initial moles of acid:

Divide the mass by the molar mass of HNO_3 :

$$\frac{159.\text{mg}}{63.0125} = \frac{0.159}{63.0125} = 0.0025233\text{molHNO}_3$$

2. Calculate the moles of H_3O^+ produced:

Firstly, use the key fact that all of the HNO_3 will be consumed by the reaction with water. Next, use the balanced chemical equation (above) to find how much H_3O^+ the reaction will produce.

$$0.0025233\text{HNO}_3 * \frac{1\text{H}_3\text{O}^+}{1\text{HNO}_3} = 0.0025233\text{H}_3\text{O}^+$$

3. Find the final molarity of H_3O^+ : Divide the moles of H_3O^+ produced by the volume of the solution in liters:

$$\frac{0.0025233\text{mol}}{0.090\text{L}} = 0.02804\text{M}$$

4. Calculate the pH:

$$\text{pH} = -\log_{10}(\text{H}_3\text{O}) = -\log_{10}(0.02804) = 1.55222\dots$$

Scratch Work

$$0.521 = \frac{(4 + 2x)(x)}{(1.2 - x)(2.4 - x)}$$

$$0.521 * (1.2 - x)(2.4 - x) = 4x + 2x^2$$

$$K * (I_a - x)(I_b - x) = (I_c + 2x)(I_d + x)$$

$$K * (I_a * I_b - (I_a + I_b)x + x^2) = I_c * I_d - 2(I_c + I_d)x + 2x^2$$

$$K(I_a * I_b) - K(I_a + I_b)x + Kx^2 = (I_c * I_d) - 2(I_c + I_d)x + 2x^2$$

$$K(I_a * I_b) - (I_c * I_d) + K(I_a + I_b)x - 2(I_c + I_d)x + Kx^2 - 2x^2 = 0$$

$$0.521(1.2 * 2.4) - (4 * 0) + 0.521(1.2 + 2.4)x - 2(4 + 0)x + 0.521x^2 - 2x^2 = 0$$

$$1.50048 + 1.87560x - 8x - (0.521 - 8)x^2 = 0$$

X	$[\text{HBrO}] + [\text{H}_2\text{O}] \rightleftharpoons [\text{BrO}^-] + [\text{H}_3\text{O}^+]$		
I	0.88	0	0
C	-x	x	x
E	0.88-x	x	x

$$K_a = \frac{x^2}{0.88 - x}$$

$$2.3 * 10^{-9} * 0.88 = x^2$$

$$x = 4.49889 * 10^{-5}$$

$$\text{pH} = -\log_{10}(x)$$

For any weak acid with K_a and initial concentration M assuming small x approximation:

$$\text{pH} = -\log_{10}(K_a * M) \quad (3)$$

X	$[HA] + [H_2O] \rightleftharpoons [H_3O^+] + [A^-]$
I	1.9 0 0
C	-x x x
E	1.9-x x x

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$K_a = \frac{x^2}{1.9 - x}$$

$$K_a = \frac{x^2}{1.9}$$

$$6.0 * 10^{-6} * (1.9) = x^2$$

$$x = \sqrt{6.0 * 10^{-6} * (1.9)}$$

$$[H^+] = 0.00337...$$

$$pH = -\log_{10} \left(\sqrt{6.0 * 10^{-6} * (1.9)} \right)$$

$$pH = 2.47154757433$$

If K = 5, then the ratio of products to reactants will be 5 - I'm assuming that we can ignore the number of atoms in each molecule so we can make it easier to determine whether the solution is in equilibrium.

$$\text{Product} = 8$$

$$\text{Reactant} = 2$$

$$\frac{P}{R} = \frac{8}{2} = 4$$

Thus the solution isn't in equilibrium.

...For K = 11

$$\text{Product} = 11$$

$$\text{Reactant} = 1$$

$$\frac{P}{R} = \frac{11}{1} = 11$$

Finding pOH - done by subtracting pH from 14:

$$pOH = 14 - pH$$

Finding % evaporated since isolated:

1. Parsons Concentration is $69 \frac{g}{L}$

2. Regular Concentration is $2.2 \frac{g}{L}$

3. Setting up equation:

$$C_0 = C_i = 2.2 \frac{g}{L}$$

4. Presently,

$$xC_0 = 69$$

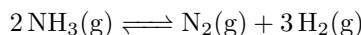
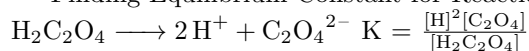
$$C_0 = 2.2$$

$$x(2.2) = 69$$

$$x = \frac{69}{2.2} = 31.36 \text{ times more salty}$$

$$\frac{100}{31.36} = 2.75\% \text{ Percent Evaporated} = 100 - 2.75 = 97.25$$

Finding Equilibrium Constant for Reaction:



Finding initial concentration for ammonia:

$$M = \frac{[NH_3]}{V} = \frac{2.2}{0.2}$$

$$M = 11M$$

X	$[2NH_3] \rightleftharpoons [N_2] + [3H_2]$		
I	11	0	0
C	-2x	x	3x
E	11 - 2x	x	3x

Creating K_c :

$$K_c = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$= \frac{[x][3x]^3}{[11 - 2x]^2}$$

Plugging in $[H_2] = 1.7/0.2 = 8.5M$ and applying small-x approximation:

$$K_c = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$= \frac{[x][8.5]^3}{[11]^2}$$

$$K = \frac{[H_2][Cl_2]}{[HCl]^2}$$

Finding M/L for HCl:

$$M = \frac{0.39\text{mol}}{17L} = 2.3 * 10^{-2}M$$

Important for finding K - since HCl is a strong acid, there will be no HCl at equilibrium ideally. And since we have 2M HCl for H2 and Cl2 respectively, the final K will be $(2.3 * 10^{-2})^2$

X	$[2HCl] \rightleftharpoons [H_2] + [Cl_2]$		
I	0.023	0	0
C	-2x	x	x
E	0.023 - 2x	x	x

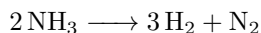
$$K = \frac{[H_2][Cl_2]}{[HCl]^2}$$

$$K = \frac{x * x}{0.023 - 2x}$$

$$(0.023 - 2x)K = x^2$$

$$0.023K -$$

Ammonia decomposition:

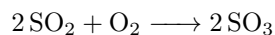


if we have 4.5 atm of ammonia gas, and have 1.4 atm of nitrogen gas in a 5L

Create ICE Table			
X	$[2NH_3]$	\rightleftharpoons	$[N_2] + [3H_2]$
I	4.5	0	0
C	-2x	x	3x
E	4.5 - 2x	x	3x

$$\begin{aligned}
 K_P &= \frac{[N_2][H_2]^3}{[NH_3]^2} \\
 &= \frac{x(3x)}{(4.5 - 2x)^3} \\
 &= \frac{1.4(3 * 1.4)}{(4.5 - 2 * 1.4)^3}
 \end{aligned}$$

Sulfur dioxide and oxygen gas react too make sulfur trioxide reaction:



Create ICE Table			
X	$[2SO_2]$	+	$[O_2] \rightleftharpoons [2SO_3]$
I	0.3066	0.4133	0
C	-2x	-x	2x
E	0.3066 - 2x	0.4133 - x	2x

$$K_b = \frac{[(CH_3)_2NH^+][OH^-]}{(CH_3)_2NH}$$

$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{(CH_3)_3N}$$

Original

$$O = 4$$

$$d = 6$$

Box C

$$O = 4$$

$$d = 6$$

Box E

$$O = 5$$

$$d = 5$$