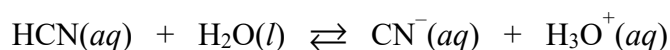## 8.1 Introduction to Acids and Bases

Essential knowledge statements from the AP Chemistry CED:
- The concentrations of hydronium ion and hydroxide ion are often reported as pH and pOH, respectively.
  - pH = $-\log[H_3O^+]$
  - pOH = $-\log[OH^-]$
- The terms "hydrogen ion" and "hydronium ion" and the symbols $H^+(aq)$ and $H_3O^+(aq)$ are often used interchangeably for the aqueous ion of hydrogen. Hydronium ion and $H_3O^+(aq)$ are preferred, but $H^+(aq)$ is also accepted on the AP Exam.
- Water autoionizes with an equilibrium constant K_w.
  - $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C
- In pure water, pH = pOH. This is called a neutral solution.
- At 25°C, $pK_w = 14.0$ and thus pH = pOH = 7.0.
  - $pK_w = 14 = pH + pOH$ at 25°C
- The value of K_w is temperature dependent, so the pH of pure, neutral water will deviate from 7.0 at temperatures other than 25°C.

4.8 Introduction to Acid-Base Reactions (Review)

Essential knowledge statements from the AP Chemistry CED:
- By definition, a Brønsted-Lowry acid is a proton donor and a Brønsted-Lowry base is a proton acceptor.
- Only in aqueous solutions, water plays an important role in many acid-base reactions, as its molecular structure allows it to accept protons from and donate protons to dissolved species.
- When an acid or base ionizes in water, the conjugate acid-base pairs can be identified and their relative strengths compared.

$$HCN(aq) \ + \ H_2O(l) \ \rightleftarrows \ CN^-(aq) \ + \ H_3O^+(aq)$$

1. Identify both sets of conjugate acid-base pairs in the reaction above.

Acid (Reactant)	Conjugate Base (Product)

Base (Reactant)	Conjugate Acid (Product)

$$CH_3NH_2(aq) + H_2O(l) \rightleftarrows CH_3NH_3^+(aq) + OH^-(aq)$$
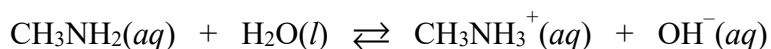
2. Identify both sets of conjugate acid-base pairs in the reaction above.
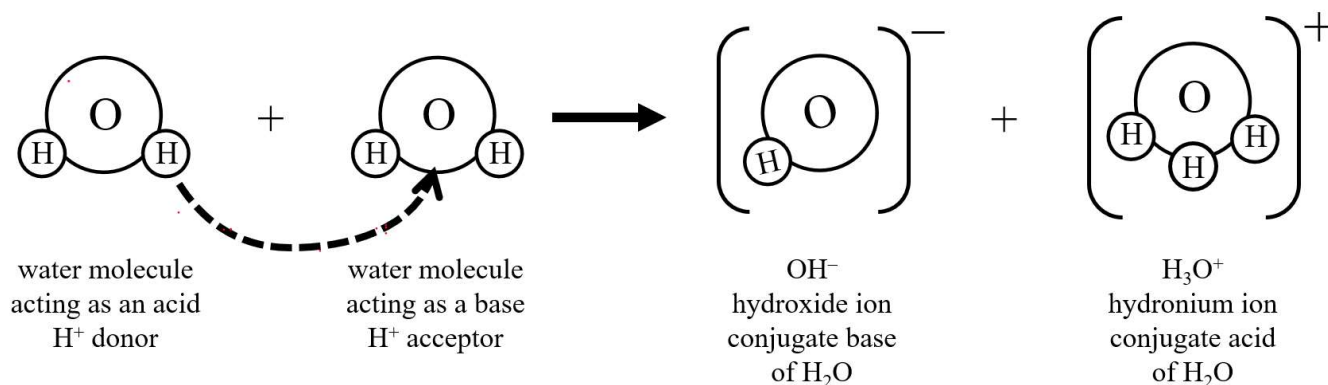
Acid (Reactant)	Conjugate Base (Product)

Base (Reactant)	Conjugate Acid (Product)

In Question #1, you observed that molecules of H_2O can behave as H^+ acceptors (bases).

In Question #2, you observed that molecules of H_2O can behave as H^+ donors (acids).

The following particle diagram illustrates the process that occurs in pure water, known as the **autoionization of water**.

water molecule	water molecule	OH^-	H_3O^+
acting as an acid	acting as a base	hydroxide ion	hydronium ion
H^+ donor	H^+ acceptor	conjugate base	conjugate acid
		of H_2O	of H_2O

- One water molecule behaves as an acid, donating a H^+ ion.
- Another water molecule behaves as a base, accepting a H^+ ion.
- The conjugate base of H_2O is the hydroxide ion, OH^-.
- The conjugate acid of H_2O is the hydronium ion, H_3O^+.

The chemical equation for the autoionization of water is shown below, along with the value of the equilibrium constant, K_w, at 25°C. An alternate form of the equation is also shown in which the H_3O^+ ion is replaced with the H^+ ion.

$$2\ H_2O(l) \rightleftarrows H_3O^+(aq) + OH^-(aq) \qquad K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°C$$

$$H_2O(l) \rightleftarrows H^+(aq) + OH^-(aq) \qquad K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°C$$

The letter "p" in symbols such as pH, pOH, and pK_w tells us to "take the negative log" of a quantity.

The concentrations for ions such as H^+ or OH^- in aqueous solution have units of mol/L or M.

Since the pH (or pOH) is calculated from a logarithm function, the pH (or pOH) does not have any units.

Remember that $[H_3O^+]$ and $[H^+]$ can be used interchangeably.

If you know the value of $[H_3O^+]$, you can calculate the pH of the solution as follows.

$$pH = -\log[H_3O^+]$$

If you know the pH of the solution, you can calculate the value of $[H_3O^+]$ as follows.

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

If you know the value of $[OH^-]$, you can calculate the pOH of the solution as follows.

$$pOH = -\log[OH^-]$$

If you know the pOH of the solution, you can calculate the value of $[OH^-]$ as follows.

$$[OH^-] = 10^{-pOH}$$

At a temperature of 25°C, the following relationships are true.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \qquad pK_w = pH + pOH = 14$$

$[H_3O^+]$	Significant Digits	pH
$1 \times 10^{-3}\ M$	1	3.0
$1.2 \times 10^{-3}\ M$	2	2.92
$1.22 \times 10^{-3}\ M$	3	2.914

With pH (or pOH), only the digits **to the right** of the decimal point are considered to be significant.

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3. Fill in the missing information in the tables below.

$[H_3O^+]$	pH
	11.65
	6.14
$4.3 \times 10^{-5}\ M$	
$0.015\ M$	

$[OH^-]$	pOH	pH
	10.39	
	8.04	
$6.8 \times 10^{-6}\ M$		
$2.8 \times 10^{-4}\ M$		

4. As the value of $[H_3O^+]$ increases, the value of pH   (decreases   increases   ).

As the value of $[OH^-]$ increases, the value of pH   (decreases   increases   ).

5. Fill in the missing information in the table below. Assume that all solutions are at 25°C.

$[H^+]$	pH	$[OH^-]$	pOH	Acidic or Basic?
0.075 M				
	5.63			
		$3.8 \times 10^{-7} M$		
			4.54	

$$H_2O(l) \rightleftarrows H^+(aq) \ + \ OH^-(aq) \qquad K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°C$$

The autoionization of water is an endothermic reaction. The value of the equilibrium constant K_w is affected by changes in temperature, as shown in the table below.

Temperature (°C)	K_w	pK_w	pH	pOH
0	1.14×10^{-15}	14.943		
10	2.93×10^{-15}	14.533		
20	6.81×10^{-15}	14.167		
25	1.01×10^{-14}	13.996		
30	1.47×10^{-14}	13.833		
40	2.92×10^{-14}	13.535		
50	5.48×10^{-14}	13.261		

6. In the table above, calculate the values of pH and pOH for pure water at various temperatures.

7. Decide if each of the following statements is true or false.

_____ A sample of pure H_2O has a pH of 7.0 at any temperature.

_____ $[H^+] = [OH^-]$ in a sample of pure H_2O at any temperature.

_____ A sample of pure H_2O is neutral, in which pH = pOH.

_____ If the pH of a sample of pure H_2O is less than 7, then the temperature of the sample must be less than 25°C.

_____ H_2O ionizes to a greater extent at 50°C than it does at 25°C.

_____ As the temperature increases, the pH of pure H_2O decreases.

_____ If a sample of pure H_2O has a pH of 6.8, then the pOH of pure water must be 7.2.

8.2 pH and pOH of Strong Acids and Bases

Essential knowledge statements from the AP Chemistry CED:
- Molecules of a strong acid (e.g., HCl, HBr, HI, $HClO_4$, H_2SO_4, and HNO_3) will completely ionize in aqueous solution to produce hydronium ions. As such, the concentration of H_3O^+ in a strong acid solution is equal to the initial concentration of the strong acid, and thus the pH of the strong acid solution is easily calculated.
- When dissolved in solution, strong bases (e.g., group I and II hydroxides) completely dissociate to produce hydroxide ions. As such, the concentration of OH^- in a strong base solution is equal to the initial concentration of the strong base, and thus the pOH (and pH) of the strong base solution is easily calculated.

Acid Name	Acid Formula	How the acid ionizes in aqueous solution
hydrochloric acid	HCl	$HCl(aq) + H_2O(l) \rightleftarrows Cl^-(aq) + H_3O^+(aq)$
hydrobromic acid	HBr	$HBr(aq) + H_2O(l) \rightleftarrows Br^-(aq) + H_3O^+(aq)$
hydroiodic acid	HI	$HI(aq) + H_2O(l) \rightleftarrows I^-(aq) + H_3O^+(aq)$
perchloric acid	$HClO_4$	$HClO_4(aq) + H_2O(l) \rightleftarrows ClO_4^-(aq) + H_3O^+(aq)$
nitric acid	HNO_3	$HNO_3(aq) + H_2O(l) \rightleftarrows NO_3^-(aq) + H_3O^+(aq)$
sulfuric acid	H_2SO_4	$H_2SO_4(aq) + H_2O(l) \rightleftarrows HSO_4^-(aq) + H_3O^+(aq)$

The table above shows the six strong acids that you should memorize for the AP Chemistry exam. A strong acid ionizes completely in aqueous solution to produce hydronium (H_3O^+) ions.

8. The following questions are related to strong acids. Assume that all solutions are at 25°C.

   (a) What is the pH of 0.050 M HCl(aq)? _____

   (b) What is the pH of 2.5×10^{-4} M HBr(aq)? _____

   (c) What is the concentration of a solution of $HClO_4(aq)$ if the pH is 4.25? _____

   (d) What is the concentration of a solution of $HNO_3(aq)$ if the pH is 6.33? _____

Base Name	Base Formula	How the base dissociates in aqueous solution
sodium hydroxide	NaOH	$NaOH(s) \rightleftharpoons Na^+(aq) + OH^-(aq)$
potassium hydroxide	KOH	$KOH(s) \rightleftharpoons K^+(aq) + OH^-(aq)$
calcium hydroxide	$Ca(OH)_2$	$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\ OH^-(aq)$
strontium hydroxide	$Sr(OH)_2$	$Sr(OH)_2(s) \rightleftharpoons Sr^{2+}(aq) + 2\ OH^-(aq)$

The table above shows examples of group I and group II hydroxides. A strong base is an ionic solid that completely dissociates in aqueous solution to produce hydroxide (OH^-) ions.

9. Fill in the missing information in the table below.

Base	Concentration of Base	$[OH^-]$	pOH	pH
NaOH	$0.15\ M$			
KOH				11.51
$Ca(OH)_2$	$4.7 \times 10^{-5}\ M$			
$Ba(OH)_2$				8.68

12.00

11.00

diluted with water

100.0 mL
NaOH(aq)

? mL
NaOH(aq)
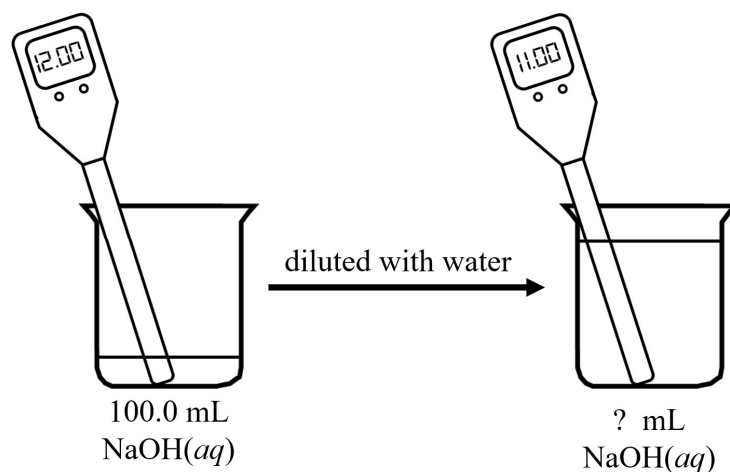
10. A student measures the pH of a 100.0 mL sample of NaOH(aq) and records the pH as 12.00. Calculate the volume of distilled water that should be added to this solution so that the pH of the diluted solution is equal to 11.00.

500.0 mL
0.050 *M* HCl(*aq*)   500.0 mL
0.10 *M* NaOH(*aq*)   1.0 L of combined
solution
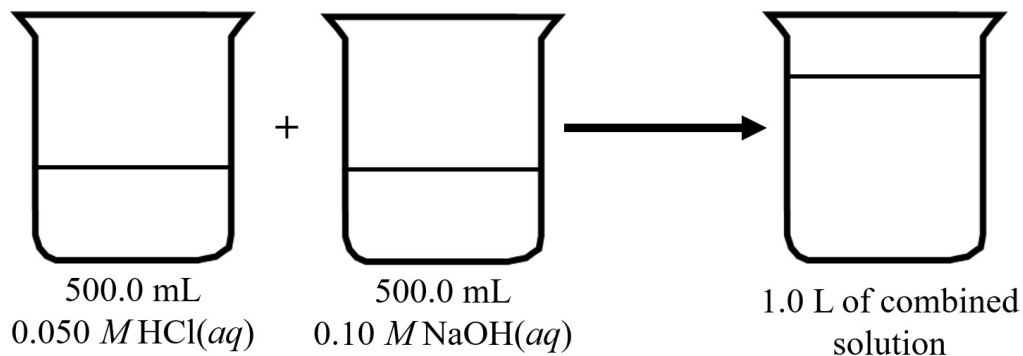
11. A student combined 500.0 mL of 0.050 *M* HCl(*aq*) with 500.0 mL of 0.10 *M* NaOH(*aq*) and mixed the final solution thoroughly.

(a) Write the net ionic equation for the reaction that occurs when aqueous solutions of a strong acid and a strong base are combined.

(b) Calculate the number of moles of HCl and the number of moles of NaOH present at the beginning of the experiment before the two solutions are combined.

(c) Calculate the number of moles of the excess reactant that remain left over in the combined solution. Assume that the acid-base reaction goes to completion.

(d) Calculate the pH of the combined solution at the end of the experiment.

12. The table below contains information about two solutions with the same initial concentration of acid.

Acid	Initial Concentration of Acid, HA	$[H_3O^+]$ at Equilibrium	pH
HNO_3	0.036 M	0.036 M	1.44
HNO_2	0.036 M	0.0036 M	2.44

$$\text{percent ionization} = \frac{[H_3O^+]_{\text{equilibrium}}}{[HA]_{\text{initial}}} \times 100\%$$

(a) Use the formula above and the data in the table to calculate the percent ionization of the acid in a solution of 0.036 M $HNO_3(aq)$. _____

(b) Use the formula above and the data in the table to calculate the percent ionization of the acid in a solution of 0.036 M $HNO_2(aq)$. _____

(c) Based on your answer to part (b), explain why HNO_2 is classified as a weak acid.

The particle diagram shown below should help you to compare the percent ionization for a strong acid and a weak acid.

0.036 M $HNO_3(aq)$
pH = 1.44

0.036 M $HNO_2(aq)$
pH = 2.44
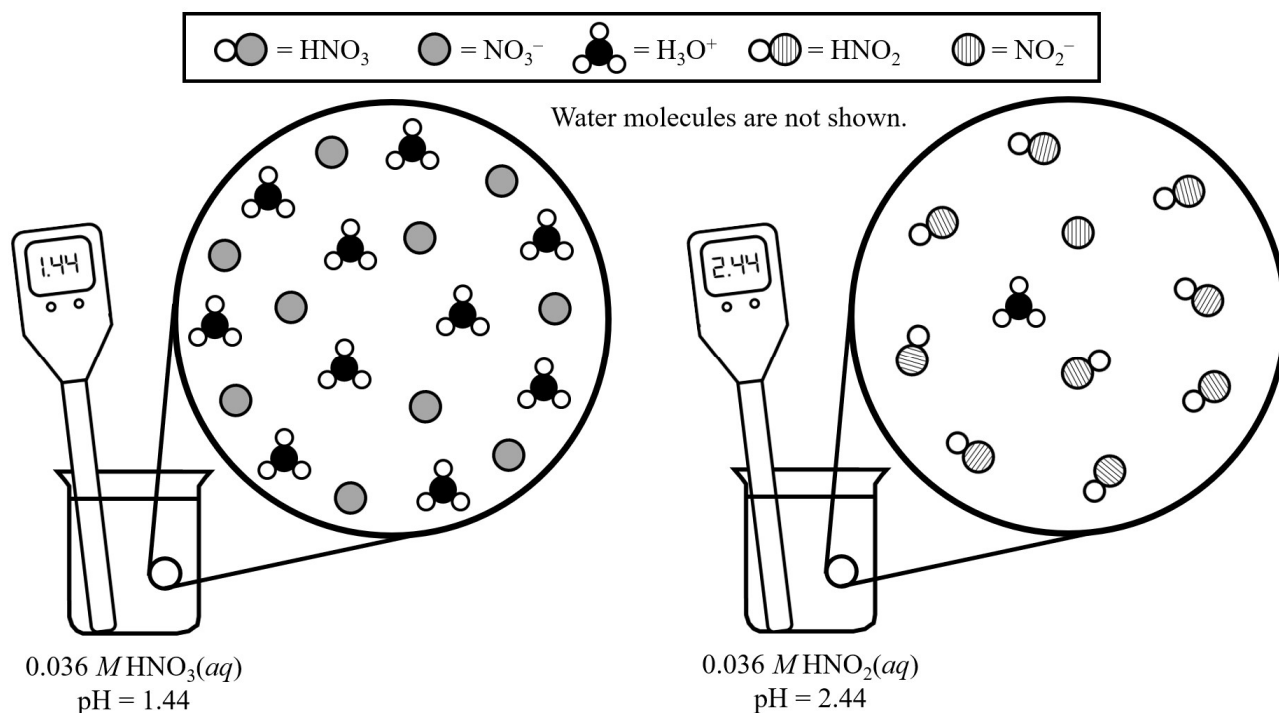
13. The table below contains information about two solutions with the same pH. Fill in the missing information in the table below.

Acid	Initial Concentration of Acid, HA	pH	$[H_3O^+]$ at Equilibrium	% Ionization of the Acid
HNO_3		1.86		
HNO_2	0.50 M	1.86		

8.3 Weak Acid and Base Equilibria

Essential knowledge statements from the AP Chemistry CED:
- Weak acids react with water to produce hydronium ions. However, molecules of a weak acid will only partially ionize in this way. In other words, only a small percentage of the molecules of a weak acid are ionized in a solution. Thus, the concentration of H_3O^+ is much less than the initial concentration of the molecular acid, and the vast majority of the acid molecules remain un-ionized.
- A solution of a weak acid involves equilibrium between an un-ionized acid and its conjugate base. The equilibrium constant for this reaction is K_a, often reported as pK_a. The pH of a weak acid solution can be determined from the initial acid concentration and the pK_a.
    - $K_a = \dfrac{[H_3O^+][A^-]}{[HA]}$
    - $pK_a = -\log K_a$
- Weak bases react with water to produce hydroxide ions in solution. However, ordinarily just a small percentage of the molecules of a weak base in solution will ionize in this way. Thus, the concentration of OH^- in the solution does not equal the initial concentration of the base, and the vast majority of the base molecules remain un-ionized.
- A solution of a weak base involves equilibrium between an un-ionized base and its conjugate acid. The equilibrium constant for this reaction is K_b, often reported as pK_b. The pH of a weak base solution can be determined from the initial base concentration and the pK_b.
   - $K_a = \dfrac{[OH^-][HB^+]}{[B]}$
    - $pK_b = -\log K_b$
- The percent ionization of a weak acid (or base) can be calculated from its pK_a (pK_b) and the initial concentration of the acid (base).

Strong Acids 100% ionized in aqueous solution	Weak Acids Partially ionized** in aqueous solution
hydrochloric acid, HCl hydrobromic acid, HBr hydroiodic acid, HI perchloric acid, $HClO_4$ nitric acid, HNO_3 sulfuric acid, H_2SO_4	hydrofluoric acid, HF acetic acid, $HC_2H_3O_2$ nitrous acid, HNO_2 hypochlorous acid, HOCl formic acid, HCO_2H hydrocyanic acid, HCN sulfurous acid, H_2SO_3 carbonic acid, H_2CO_3 etc.

**"Partially ionized" is not the same as "partially soluble." A weak acid can be very soluble in water, but only a small fraction of the acid molecules are ionized in aqueous solution.

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14. Hydrofluoric acid, HF, is a weak acid that ionizes in aqueous solution according to the equation shown below.
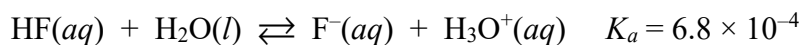
$$HF(aq) \; + \; H_2O(l) \; \rightleftarrows \; F^-(aq) \; + \; H_3O^+(aq) \quad K_a = 6.8 \times 10^{-4}$$

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

Compare your answer to part (a) with the K_a expression on the AP Chemistry Equations and constants sheet.

(b) The pH of 0.50 M HF(aq) is 1.74.

  (i) Calculate the value of $[H^+]$ in 0.50 M HF(aq).

  (ii) Calculate the percent ionization in 0.50 M HF(aq).

15. Acetic acid, $HC_2H_3O_2$, is a weak acid found in household vinegar. Acetic acid ionizes in aqueous solution according to the equation shown below.

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons C_2H_3O_2^-(aq) + H_3O^+(aq) \quad K_a = 1.8 \times 10^{-5}$$
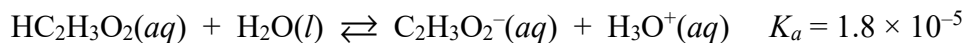
(a) Do you predict that the pH of 0.50 M $HC_2H_3O_2(aq)$ should be less than, greater than, or equal to the pH of 0.50 M $HF(aq)$? Your justification should include a comparison of the K_a values for HF and $HC_2H_3O_2$ and a comparison of the relative strength of these two acids.

(b) Fill in the missing information in the R-I-C-E table below. Let the variable "x" represent the concentration of the hydronium ion at equilibrium.

R eaction	$HC_2H_3O_2(aq)$	+	$H_2O(l)$	$\rightleftharpoons$	$C_2H_3O_2^-(aq)$	+	$H_3O^+(aq)$
I nitial	0.50 M		N/A		0 M		$\approx 0\ M$
C hange			N/A				
E quilibrium			N/A				

(c) Plug in all three values (in terms of x) from the bottom row of your R-I-C-E table into the K_a expression shown below. Then solve for x.

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 1.8 \times 10^{-5}$$

On the AP Chemistry Exam, you will not need to use the quadratic equation to solve any mathematical calculations. Acetic acid is a weak acid. It is only partially ionized in aqueous solution. You can make the following assumption, in order to simplify the math in part (c).

$[H_3O^+]_{equilibrium}$ is much smaller than $[HC_2H_3O_2]_{initial}$

$$(0.50\ M - x) \approx 0.50\ M$$

(d) In part (c), you solved for x, which represents $[H_3O^+]_{equilibrium}$ in 0.50 M $HC_2H_3O_2(aq)$. Calculate the pH of 0.50 M $HC_2H_3O_2(aq)$.
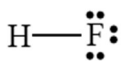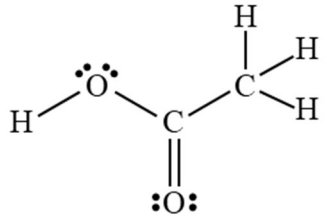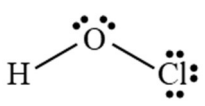
16. Hypochlorous acid, HOCl, is a weak acid that forms when chlorine dissolves in water. HOCl molecules react with water molecules to form hypochlorite ions, OCl^-, and hydronium ions.

(a) Write the balanced chemical equation for the ionization of HOCl in aqueous solution.

(b) Write the K_a expression for hypochlorous acid.

(c) The pH of 0.50 M HOCl(aq) is 3.92. Make a R-I-C-E table to help you calculate the concentrations of HOCl, H_3O^+, and OCl^- at equilibrium in 0.50 M HOCl(aq).

(d) Use the information from part (c) to calculate the value of K_a for HOCl.

Acid	HF	$HC_2H_3O_2$	HOCl
Structural Formula	H—F:	(structural formula)	H—O—Cl:
K_a	6.8×10^{-4}	1.8×10^{-5}	2.9×10^{-8}
$[HA]_{initial}$	0.50 M	0.50 M	0.50 M
pH	1.74	2.52	3.92
$[H_3O^+]_{equilibrium}$	0.018 M	0.0030 M	0.00012 M
% ionization	3.6%	0.6%	0.024%

17. Use the information in the table above to answer the following questions.

   (a) When comparing two different acids, the stronger acid is

   the acid that has the (lower   higher) K_a value.

   (b) When comparing two different acids that have the same $[HA]_{initial}$,

   the stronger acid will have the (lower   higher) pH.

There is an interesting relationship between $[HA]_{initial}$ and the percent ionization of the acid.

	$[HA]_{initial}$	pH	$[H_3O^+]_{equilibrium}$	Percent Ionization
hydrofluoric acid	1.0 M	1.59	0.026 M	2.6%
	0.50 M	1.74	0.018 M	3.6%
	0.10 M	2.10	0.0079 M	7.9%
HF	0.050 M	2.26	0.0055 M	11%
	0.010 M	2.64	0.0023 M	23%

|	$[HA]_{initial}$	pH	$[H_3O^+]_{equilibrium}$	Percent Ionization
acetic acid	1.0 M	2.38	0.0042 M	0.42%
	0.50 M	2.52	0.0030 M	0.60%
	0.10 M	2.89	0.0013 M	1.3%
$HC_2H_3O_2$	0.050 M	3.03	0.00094 M	1.9%
	0.010 M	3.38	0.00042 M	4.2%

18. As the value of $[HA]_{initial}$ decreases…

…the value of $[H_3O^+]_{equilibrium}$ (decreases   increases).

…the pH of the solution (decreases   increases).

…the percent ionization of the acid (decreases   increases).

The inverse relationship between $[HA]_{initial}$ and the percent ionization of the acid can be explained in terms of Le Châtelier's principle.

$$HA(aq) \rightleftarrows H^+(aq) + A^-(aq) \quad K_a = \frac{[H^+][A^-]}{[HA]}$$
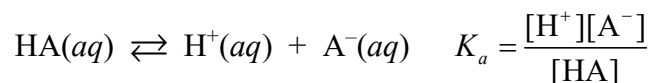
During Unit 7, you learned the following.

If an aqueous equilibrium system is diluted with water, the concentrations of all aqueous species are decreased. As a result of this change, the aqueous equilibrium system shifts toward the side of the equation that has the higher number of moles of aqueous particles.

Suppose that 500 mL of H_2O is added to 500 mL of a weak acid, HA.

$$Q_a = \frac{\frac{1}{2}[H^+]\frac{1}{2}[A^-]}{\frac{1}{2}[HA]} = \frac{1}{2}K_a$$

Dilution with water causes the value of Q_a to become less than K_a. This causes a shift in the equilibrium with a net conversion from reactants to products. As a result, the percent ionization of the acid increases.

You have seen several examples of weak acids. Now let's examine the properties of **weak bases**.

Strong Bases (Group I and II hydroxides) 100% dissociated into ions in aqueous solution	Weak Bases Partially ionized in aqueous solution
sodium hydroxide, NaOH potassium hydroxide, KOH calcium hydroxide, Ca(OH)$_2$ strontium hydroxide, Sr(OH)$_2$	ammonia, NH$_3$ methylamine, CH$_3$NH$_2$ pyridine, C$_5$H$_5$N fluoride ion, F$^-$ acetate ion, C$_2$H$_3$O$_2$$^-$ etc.

19. Ammonia, NH_3, is a weak base that ionizes in aqueous solution according to the equation shown below.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \quad K_b = 1.8 \times 10^{-5}$$

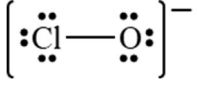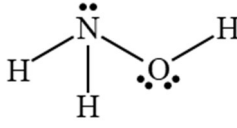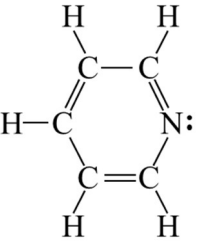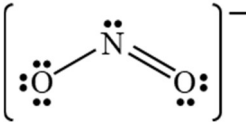(a) Write the expression for the equilibrium constant, K_b, for this reaction.
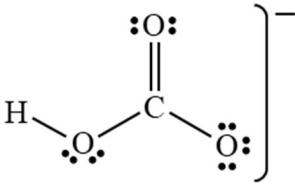
Compare your answer to part (a) with the K_b expression on the AP Chemistry Equations and constants sheet.

(b) The pH of 0.50 M $NH_3(aq)$ is 11.48. Calculate the value of $[OH^-]$ in 0.50 M $NH_3(aq)$.

(c) Use the formula below and your answer to part (b) to calculate the percent ionization of the base in a solution of 0.50 M $NH_3(aq)$. _____

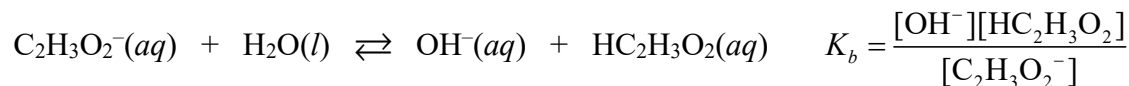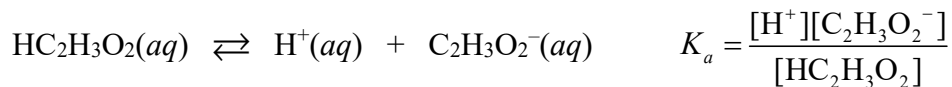$$\text{percent ionization} = \frac{[OH^-]_{equilibrium}}{[B]_{initial}} \times 100\%$$

Two important categories of weak bases are listed in the table below.

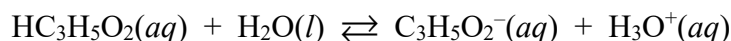| neutral molecules containing an atom with a lone pair of electrons that can accept H^+			anions that are the conjugate base of a weak acid	
ammonia NH_3	methylamine CH_3NH_2	acetate $C_2H_3O_2^-$	hypochlorite ClO^-	
hydroxylamine NH_2OH	pyridine C_5H_5N	nitrite NO_2^-	hydrogen carbonate HCO_3^-	

Acetic acid, $HC_2H_3O_2$, and acetate, $C_2H_3O_2^-$, represent a conjugate acid-base pair. The relationship between the values of K_a and K_b for a conjugate acid-base pair is represented by the equations below.

$$K_a \times K_b = 1.0 \times 10^{-14} \text{ at } 25°C \qquad pK_a + pK_b = 14.00 \text{ at } 25°C$$

The example shown below is for the weak acid acetic acid ($HC_2H_3O_2$) and its conjugate base acetate ($C_2H_3O_2^-$). However, the relationship $K_a \times K_b = 1.0 \times 10^{-14}$ works with any conjugate acid-base pair.

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq) \qquad K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$C_2H_3O_2^-(aq) + H_2O(l) \rightleftharpoons OH^-(aq) + HC_2H_3O_2(aq) \qquad K_b = \frac{[OH^-][HC_2H_3O_2]}{[C_2H_3O_2^-]}$$

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \qquad K_w = [H^+][OH^-] = K_a \times K_b$$

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20. Propanoic acid, $HC_3H_5O_2$, ionizes in aqueous solution according to the equation shown below.

$$HC_3H_5O_2(aq) + H_2O(l) \rightleftharpoons C_3H_5O_2^-(aq) + H_3O^+(aq)$$
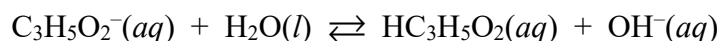
(a) Write the expression for the equilibrium constant, K_a, for this reaction.

(b) The pK_a for $HC_3H_5O_2$ is 4.88. Calculate the value of K_a for $HC_3H_5O_2$. _____

(c) Calculate the pH of 0.50 M $HC_3H_5O_2(aq)$.

The propanoate ion, $C_3H_5O_2^-$, ionizes in aqueous solution according to the equation shown below.

$$C_3H_5O_2^-(aq) + H_2O(l) \rightleftharpoons HC_3H_5O_2(aq) + OH^-(aq)$$

(d) Calculate the value of K_b for the propanoate ion, $C_3H_5O_2^-$. _____

(e) Calculate the pH of 0.50 M $NaC_3H_5O_2(aq)$.