

AP Chemistry
Unit XI Learning Targets
Acid-Base Equilibria (Ch. 16)

At the end of this unit I will be able to . . .

#	Target	Book section
1	list the general properties that characterize acidic and basic solutions and identify the ions responsible for these properties.	16.1, notes
2	define the terms acid and base using the Arrhenius Theory, the Bronsted-Lowry Theory and the Lewis Theory of acids and bases.	16.2 & 16.11
3	identify the conjugate base associated with a given Bronsted-Lowry acid and the conjugate acid associated with a given Bronsted-Lowry base.	16.2
4	explain what is meant by the autoionization of water and write the ion-product constant expression for the process.	16.3
5	calculate pH, pOH, [H ⁺] and [OH ⁻] for a strong acid or a strong base.	16.4 & 16.5
6	make calculations associated with weak acids. This includes solving for the K _a , % ionization and pH of the weak acid.	16.6 & 16.8
7	make calculations associated with weak bases. This includes solving for the K _b , % ionization and pH of the weak base.	16.7 & 16.8
8	predict whether a particular salt solution will be acidic, basic, or neutral and write the associated hydrolysis equation. Using the equation, I can calculate the pH of an acidic or basic salt solution.	16.9
9	explain how acid strength relates to polarity and the strength of the H-X bond.	16.10
10	predict the relative strengths of oxyacids.	16.10
11	define an acid and a base in terms of the Lewis concept.	16.11

Target 1: list the general properties that characterize acidic and basic solutions and identify the ions responsible for these properties.

General Properties of Acids and Bases

Acids	Bases
sour tasting	bitter tasting
turn litmus red	turn litmus blue
phenolphthalein colorless	phenolphthalein pink
reacts w/ metals \rightarrow H_2 gas	feels slippery to touch
reacts w/ carbonates \rightarrow CO_2 gas	reacts with acids \rightarrow salt & H_2O
reacts with bases \rightarrow salt & H_2O	conducts electricity
conduct electricity	$pH > 7$; $[H^+] < [OH^-]$
$pH < 7$; $[H^+] > [OH^-]$	☺

Target 2: define the terms acid and base using the Arrhenius Theory, the Bronsted-Lowry Theory and the Lewis Theory of acids and bases.

There are three common ways to define an acid or a base . . .

1. Arrhenius acid = substance that ionizes in water to produce H^+ (HCl, HNO_3)
Arrhenius base = substance that ionizes in water to produce OH^- (NaOH, $Ca(OH)_2$)
2. Bronsted-Lowry acid = substance capable of donating a proton/ H^+ ion
Bronsted-Lowry base = substance capable of accepting a proton/ H^+ ion
3. Lewis acid = substance capable of accepting an electron pair
Lewis base = substance capable of donating an electron pair

* The Bronsted-Lowry definition of acids and bases is a much more comprehensive definition of acids and bases than the Arrhenius one. The Arrhenius definition, although still commonly used, has 2 serious flaws in defining acids and bases. First, it only applies to acid-base reactions in aqueous solutions. Secondly, it cannot explain why NH_3 is base even though it doesn't contain a hydroxide ion. The Lewis A/B definition is even more general. We'll look at it later in section 16.11.

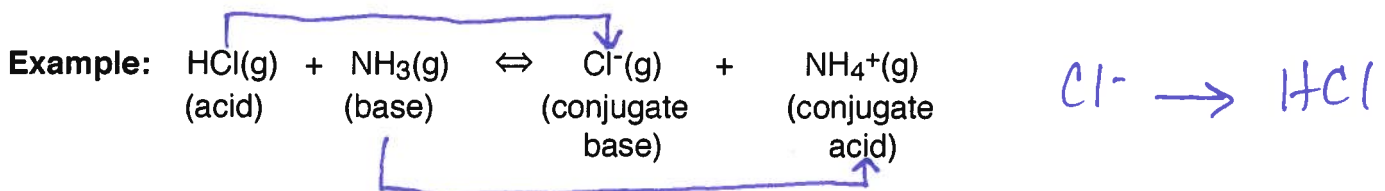
Target 3: Identify the conjugate base associated with a given Bronsted-Lowry acid and the conjugate acid associated with a given Bronsted-Lowry base.

The **conjugate acid** is the substance formed by the addition of a proton to a Bronsted-Lowry base.

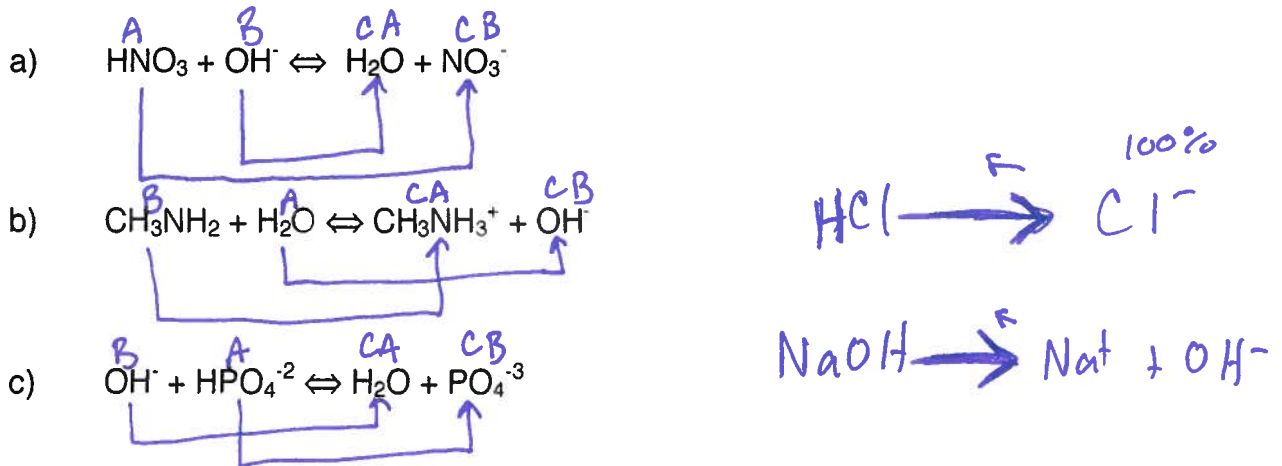
The **conjugate base** is the substance formed by the loss of a proton from a Bronsted-Lowry acid.

A **conjugate acid-base pair** is an acid and a base, such as water and OH^- , that differ only by the presence of a proton, H^+ .

✶ Water is an example of a substance which is said to be **amphoteric**... that is it can act as an acid in some reactions and as a base in others.



Practice: Determine the acid, base, conjugate acid and conjugate base.



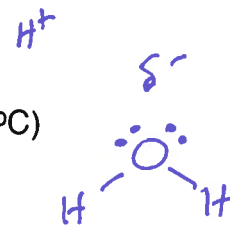
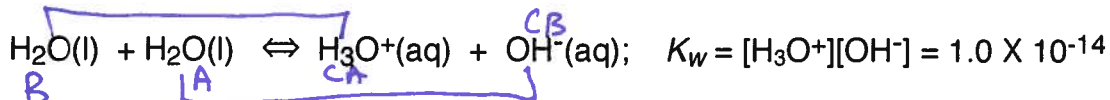
Relative Strengths of Acids and Bases: The stronger an acid, the weaker its conjugate base. The stronger a base, the weaker its conjugate acid.

Target 4: Explain what is meant by the autoionization of water and write the ion-product constant expression for the process.

The autoionization of water is the process whereby water spontaneously forms low concentrations of $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ ions by proton transfer from one water molecule to another. (B/L definition) The equilibrium constant for water, K_w (called the ion-product constant), can be written as follows:



OR



Approximately 1 out of every billion water molecules is actually ionized at any given moment in time. It should be no surprise that pure water is a poor conductor of electricity!

Target 5: Calculate pH, pOH, $[\text{H}^+]$ and $[\text{OH}^-]$ for a strong acid or a strong base.

The concentration of $\text{H}^+(\text{aq})$ can be expressed in terms of pH:

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

Most pH values range from 0 to 14, but can be higher or lower. The pH of a neutral solution at 25°C is 7.00, whereas an acidic solution has a pH less than 7.00 and a basic solution has a pH of greater than 7.00. The pX notation is simply a mathematical notation used to express smaller numbers as bigger numbers. We will encounter other pX values such as pK_a , pK_b , pOH, and pK_w . The pH scale is a logarithmic scale. Figure 16.5 (p 663) illustrates the fact that a change in pH of 1 corresponds to a change in $[\text{H}^+]$ of 10!

pH meters and indicators are commonly used to measure pH. (see figure 16.7 on p 664)

Notes:

- $\text{pH} + \text{pOH} = 14$
 - $\text{pOH} = -\log[\text{OH}^-]$
 - only count significant digits past the decimal point in a log scale (for example, a pH of 2.54 has 2 significant digits!
- Handwritten note: -- inverse log -- $\rightarrow 10^{-2.54} = 2.9 \times 10^{-3} \text{ M (2 sig figs!)$*

Strong acids and strong bases exist entirely as ions in solution (strong electrolytes). There are 7 common strong acids . . . please memorize! **SNICl BrCl !!!**

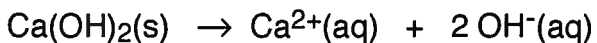
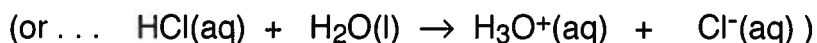
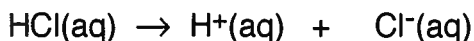
HNO_3 = nitric acid	HCl = hydrochloric acid
H_2SO_4 = sulfuric acid	HBr = hydrobromic acid
HClO_3 = chloric acid	HI = hydroiodic acid
HClO_4 = perchloric acid	

There are few common strong bases . . . please memorize!

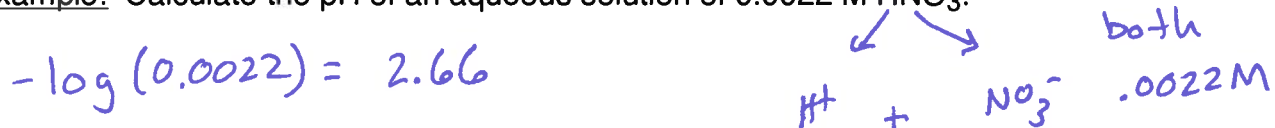
The ionic hydroxides of 1A metals (LiOH , NaOH , KOH , RbOH , and CsOH)
Plus, $\text{Ca}(\text{OH})_2$, $\text{Ba}(\text{OH})_2$, and $\text{Sr}(\text{OH})_2$.

Handwritten note: Sr. Ba(a)

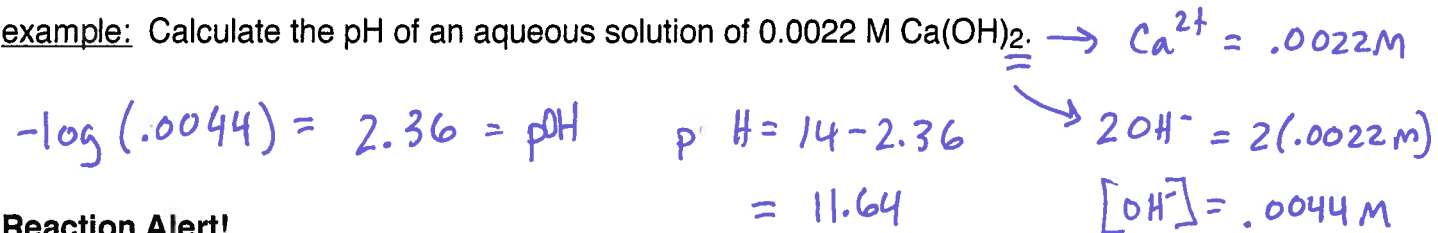
The equations for strong acids and bases are written using single arrows as the dissolution process is complete . . .



example: Calculate the pH of an aqueous solution of 0.0022 M HNO₃.

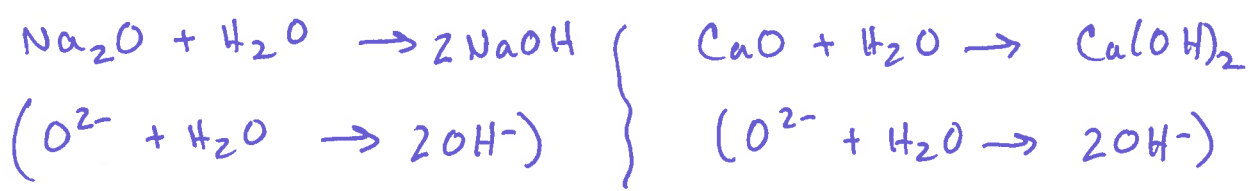


example: Calculate the pH of an aqueous solution of 0.0022 M Ca(OH)₂.



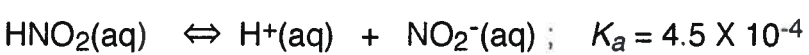
*** Reaction Alert!**

Metal oxides (such as Na₂O and CaO) are often used to make bases:

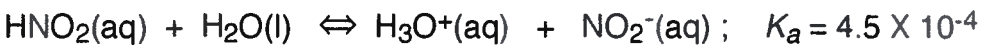


Target 6: Make calculations associated with weak acids. This includes solving for the K_a , % ionization and pH of the weak acid.

The ionization of a weak acid is written as an equilibrium. The extent of ionization can be measured by an equilibrium constant called an **acid-dissociation constant**, K_a :

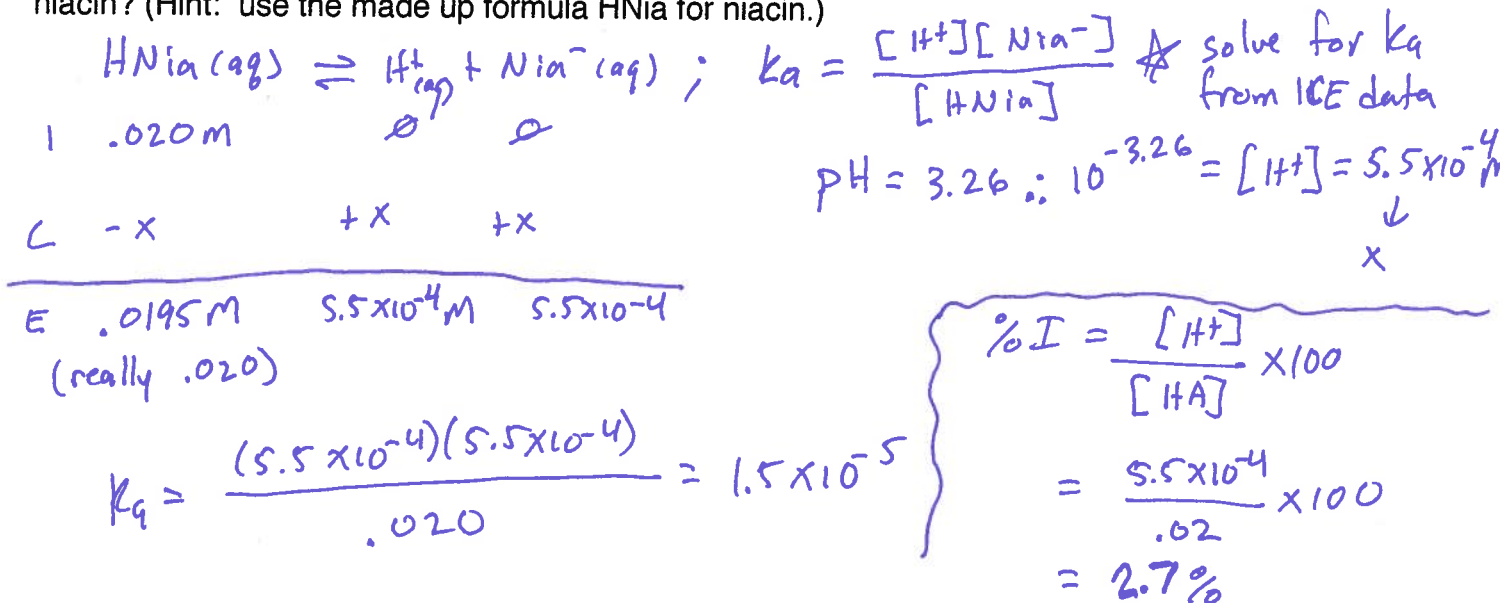


The equilibrium can also be written showing the water in the equation . . .

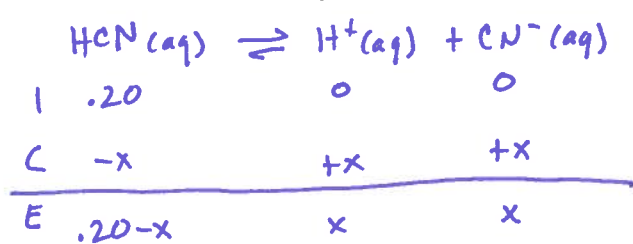


The larger the K_a value (the farther shifted to the right), the stronger the acid! See Table 16.2 (p 667) or Appendix D for a list of some common weak acids and their K_a 's.

Example: Niacin is a vitamin which is a weak acid. A 0.020 M solution of niacin has a pH of 3.26. What is the percentage of the acid which is dissociated? What is the K_a for niacin? (Hint: use the made up formula HNia for niacin.)



H.W. Practice: Calculate the pH of a 0.20 M solution of HCN at 25°C. The K_a for HCN is 4.9×10^{-10} at 25°C.



$$K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]} = 4.9 \times 10^{-10} = \frac{x^2}{.2-x}$$

Since K_a is so small we will ignore

$$4.9 \times 10^{-10} = \frac{x^2}{.2}$$

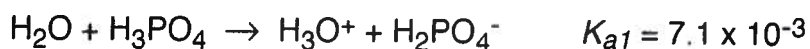
$$x = 9.9 \times 10^{-6} \text{ M} = [\text{H}^+]$$

$$\text{pH} = -\log(9.9 \times 10^{-6}) = \boxed{5.00}$$

$$\% I = \frac{[\text{H}^+]}{[\text{HA}]} \times 100$$

$$= \frac{9.9 \times 10^{-6}}{.2} \times 100 = \boxed{.0050\%}$$

Polyprotic acids are acids which have more than one ionizable proton such as H_2SO_3 and H_3PO_4 . We picture the hydrogens ionizing in successive steps as illustrated by:

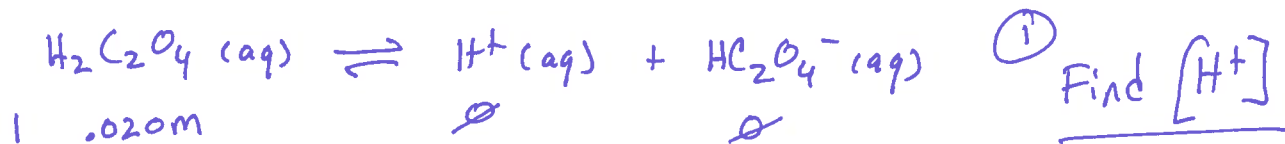


The first proton is always easier to remove than successive protons as indicated by the magnitudes of K_{a1} , K_{a2} and K_{a3} . K_{a1} will always be larger than K_{a2} . K_{a2} will always be

larger than K_{a3} . When solving a pH problem involving polyprotic acids, the pH can be estimated by considering only K_{a1} as long as K_{a1} is 1000X larger than K_{a2} . What's in Appendix D?

Example: Calculate the pH and the oxalate ion concentration in a 0.020 M solution of oxalic acid, $H_2C_2O_4$. The K_{a1} is 5.9×10^{-2} and the K_{a2} is 6.4×10^{-5} .

Since K_{a1} is 1000x larger than K_{a2} , pH can be estimated with the use of only K_{a1}



C	-x	+x	+x
E	.020-x	x	x

$$K_{a1} = \frac{x^2}{.020-x} = 5.9 \times 10^{-2} \quad x = \underline{.01578 M} = [H^+], [HC_2O_4^-]$$

need solver

② Find pH

$$-\log [.01578] = 1.80$$

③ Find oxalate ion concentration

$$HC_2O_4^- \rightleftharpoons H^+ + C_2O_4^{2-}$$

I	.01578	.01578	0
C	-x	+x	+x
E	.01578-x	.01578+x	x

$$K_{a2} = \frac{[H^+][C_2O_4^{2-}]}{[HC_2O_4^-]}$$

$$6.4 \times 10^{-5} = \frac{(.01578+x)(x)}{.01578-x}$$

$$x = 6.34 \times 10^{-5} M = [C_2O_4^{2-}]$$

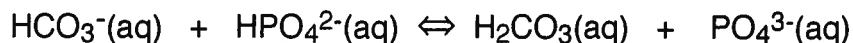
AP Chemistry Review Targets 1-6

Directions: Answer each of the following questions. Place your answers on the provided blanks. This is an open note quiz.

1. List the conjugate base of CH_3COOH . CH_3COO^-

2. List the conjugate acid of H_2AsO_4^- . H_3AsO_4

Consider the following reaction for Questions 3 and 4:



3. Identify the acid in the above reaction. HPO_4^{2-} CB

4. Identify the conjugate base in the above reaction. PO_4^{3-}

5. Select all correct answers. According to Arrhenius, an acid is a substance that _____.

- a) is capable of donating one or more protons ✓
- b) causes an increase in the concentration of the H^+ ions in solution ✓
- c) can accept a pair of e^- to form a coordinate covalent bond not Arrhenius
- d) reacts with the solvent to form a cation formed by the autoionization of water of that solvent not Arrhenius
- e) tastes bitter

6. A substance is said to be amphoteric if it is capable of acting as both an acid and a base.

7. In an acidic solution, the $[\text{H}_3\text{O}^+]$ > $[\text{OH}^-]$. Fill in the blank with <, >, or =.

8. $\text{HSO}_3^-(\text{aq})$ is an ion which is amphoteric. In the space below, write an equation for the reaction in which $\text{HSO}_3^-(\text{aq})$ acts as a base in $\text{H}_2\text{O}(\text{l})$.



9. In the space below, write an equation for the reaction in which $\text{HSO}_3^-(\text{aq})$ acts as an acid in $\text{H}_2\text{O}(\text{l})$.



10. Predict the products of the following reaction:



11. A solution has $[H^+]$ of 0.00032 M. The $[OH^-]$ is 3.13×10^{-11} M and the solution is acidic. (acidic, basic, or neutral)
12. What is the value for K_w at 25°C? 1×10^{-14}

For #13-23, show your work and then circle your final answers.

13. How many times more acidic is a solution with a pH of 3.00 than a solution with a pH of 5.50?

$$3 \xrightarrow{10} 4 \xrightarrow{10} 5 \xrightarrow{10} 6 \quad 10 \times 10 \times 10 = 1000 \times$$

14. If a solution has a pH of 4.52, then its $[H^+]$ is 3.02×10^{-5} M.

$$10^{-4.52} \rightarrow$$

15. If a solution has a pOH of 10.80, its $[H^+]$ is 6.31×10^{-4} M.

16. If a magnesium hydroxide solution has a $[OH^-]$ of 1.00×10^{-11} M, then its pH is 3.00



17. Calculate the pH of a 0.0034 M solution of nitric acid.

HNO_3 is a strong acid!

$$[H^+] = 0.0034 \text{ M}$$

$$pH = -\log(0.0034) = \boxed{2.47}$$

18. Calculate the pOH of a solution made by dissolving 2.50 grams of $HClO_3$ in enough water to make 2.50 L of solution.

$$\frac{2.50 \text{ g}}{2.50 \text{ L}} \times \frac{1 \text{ mole } HClO_3}{84.54 \text{ g } HClO_3} = 0.0118 \text{ M} = [H^+]$$

$$pH = -\log(0.0118) = 1.928$$

$$pOH = 14 - 1.928 = \boxed{12.072}$$

19. Calculate the pH of a solution made by dissolving 14.0 g of strontium hydroxide in enough water to make 3.00 L of solution.

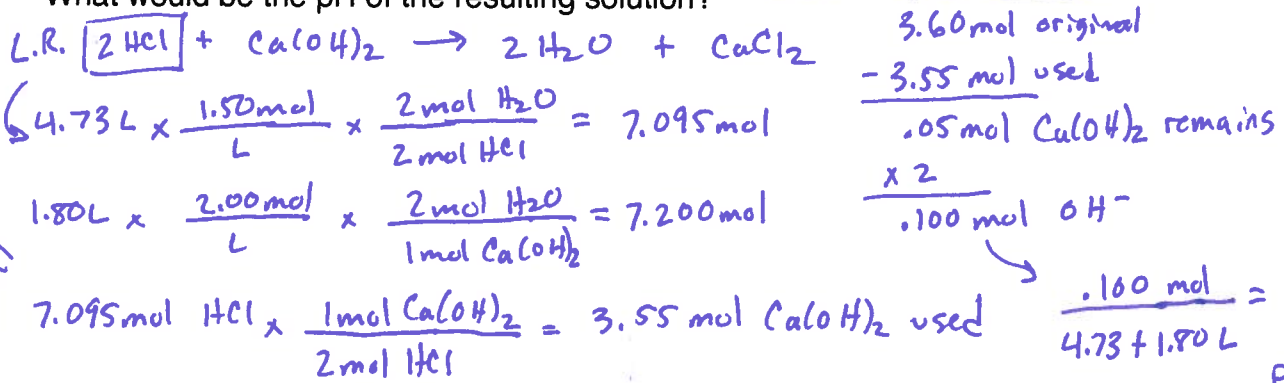
$$\frac{14.0 \text{ g}}{3.00 \text{ L}} \times \frac{1 \text{ mol } Sr(OH)_2}{121.62 \text{ g}} = .0384 \text{ M} \times 2 = .0768 \text{ M} = [OH^-]$$

$$pOH = -\log(0.0768) = 1.115$$

$$pH = 14 - 1.115 = \boxed{12.885}$$

Limiting reactant problem!

20. Assume that you added 4.73 liters of 1.50 M HCl to 1.80 liters of 2.00 M Ca(OH)₂. What would be the pH of the resulting solution?



Ch. 17
molecular

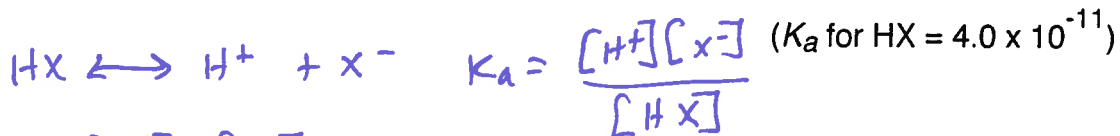
21. Assume you wanted to make 50.0 L of a sodium hydroxide solution that has a pH of 10.00. How many grams of sodium hydroxide would this require?

pH = 10.00 so $[\text{H}^+] = 1.00 \times 10^{-10}\text{ M}$ and $[\text{OH}^-] = 1.00 \times 10^{-4}\text{ M}$

$50.0\text{ L} \times \frac{1.00 \times 10^{-4}\text{ mol}}{\text{L}} \times \frac{40.0\text{ g NaOH}}{1\text{ mol NaOH}} = \boxed{2.00\text{ g NaOH}}$

pH = 12.19

22. What is $[\text{H}^+]$ in a 0.10 M solution of HX, a weak acid, at 25°C?



$[\text{H}^+] = [\text{X}^-]$

so, $4.0 \times 10^{-11} = \frac{x^2}{.10}$

$x = 20 \times 10^{-6}\text{ M}$

23. What is the % ionization of a 3.00 M acetic acid solution at 25°C? For acetic acid the is K_a = 1.8 × 10⁻⁵.

$1.8 \times 10^{-5} = \frac{x^2}{3.00}$

$x = .00735\text{ M}$

$\%I = \frac{.00735}{3.00} \times 100$

$= .245\%$

As the strength of an acid increases (larger K_a) the strength of the base decreases (smaller K_b), so the product is always 1.0×10^{-14} . Ionization constants are usually only listed for one member of a conjugate acid-base pair. This is because the K_b can readily be calculated from K_a and visa-versa.

$$K_w = (K_a)(K_b) = 1.00 \times 10^{-14}$$

Example: Use Appendix D to calculate the K_b for the fluoride ion.

$$K_a = 6.8 \times 10^{-4} \quad 1.00 \times 10^{-14} = (6.8 \times 10^{-4})(K_b) \quad \boxed{K_b = 1.5 \times 10^{-11}}$$

Also note the following equation to be true: $pK_a + pK_b = 14.00$

AP Chemistry Review for Quiz: Targets 1-7

Review the following materials:

- problem sets #1-3
- *definitions*: Bronsted-Lowry, Arrhenius and Lewis acids and bases, strong acid, weak acid, strong base, weak base, amphoteric, K_a , K_b , K_w , autoionization of water
- Finding pH and pOH of acids and bases (both strong and weak) when given concentrations and K values.
- Finding K_a or K_b when given concentrations and pH.
- Acidic and basic anhydrides (sponge method).
- Calculate the percent dissociation of an acid when given concentration and K_a
- Calculate pH when given % dissociation and concentration.
- Calculate pH/pOH of a polyprotic acid.

Directions: Answer each of the following questions.

1. How is an Arrhenius acid different than a Bronsted-Lowry acid?

2. Which of the following is/are Bronsted-Lowry acids?

- (a) $(\text{CH}_3)_3\text{NH}^+$ (b) CH_3COOH (c) HF (d) HNO_2

3. What is the molar concentration of the hydronium ion in pure water at 25°C?

$$1.00 \times 10^{-7} \text{ M}$$

4. What is the pH of a solution at 25°C in which the $[\text{OH}^-]$ is 2.8×10^{-13} ?

$$\text{pOH} = -\log(2.8 \times 10^{-13}) = 12.55 \quad \text{pH} = 14 - 12.55 = \boxed{1.45}$$

5. What is the concentration (in M) of hydronium ions in a solution at 25°C with a hydroxide ion concentration of 2.31×10^{-4} M?

$$1.00 \times 10^{-14} = (x)(2.31 \times 10^{-4}) \quad x = 4.33 \times 10^{-11} \text{ M}$$

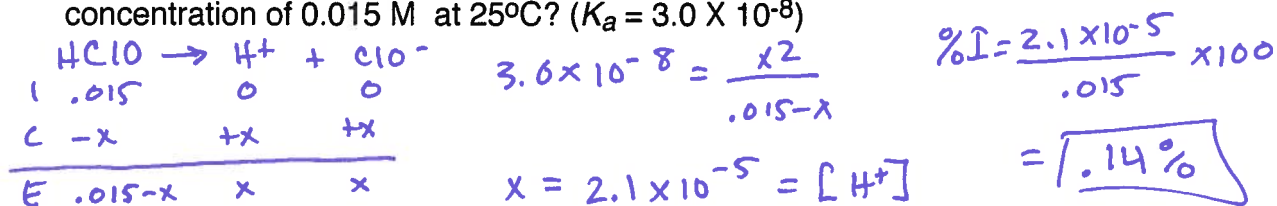
6. Which solution below has the highest hydroxide ion concentration?

solution #	pH
1	3.21
2	7.93
3	9.92

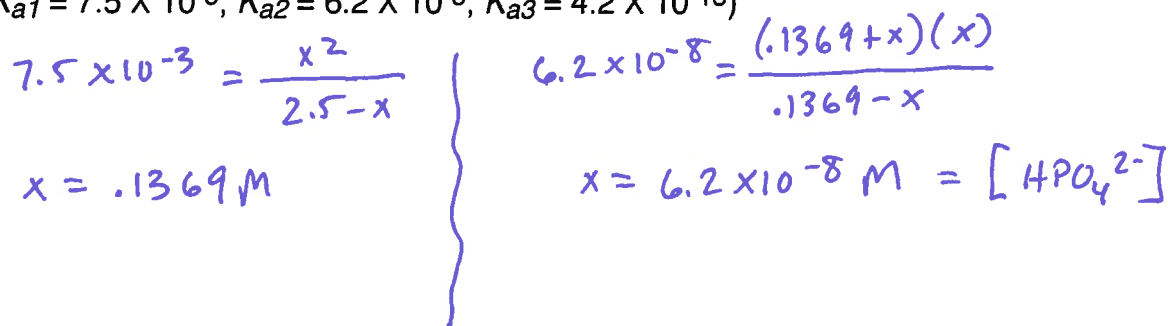
7. What is the pH of a 0.053 M solution of $\text{Ca}(\text{OH})_2$?

$$[\text{OH}^-] = 2(0.053) = .106 \text{ M} \quad -\log(.106) = .975 \quad \text{pH} = 14 - .975 = \boxed{13.03}$$

8. What is the percent ionization of a hypochlorous acid solution (HClO) with an initial concentration of 0.015 M at 25°C? ($K_a = 3.0 \times 10^{-8}$)



9. What is the molar concentration of HPO_4^{2-} in a 2.5 M solution of phosphoric acid? ($K_{a1} = 7.5 \times 10^{-3}$, $K_{a2} = 6.2 \times 10^{-8}$, $K_{a3} = 4.2 \times 10^{-13}$)



10. 25.0 mL of 0.100 M $\text{Ba}(\text{OH})_2$ is added to 25.0 mL of 0.100 M HCl. What is the pH of this solution?

* Ch. 17 material *

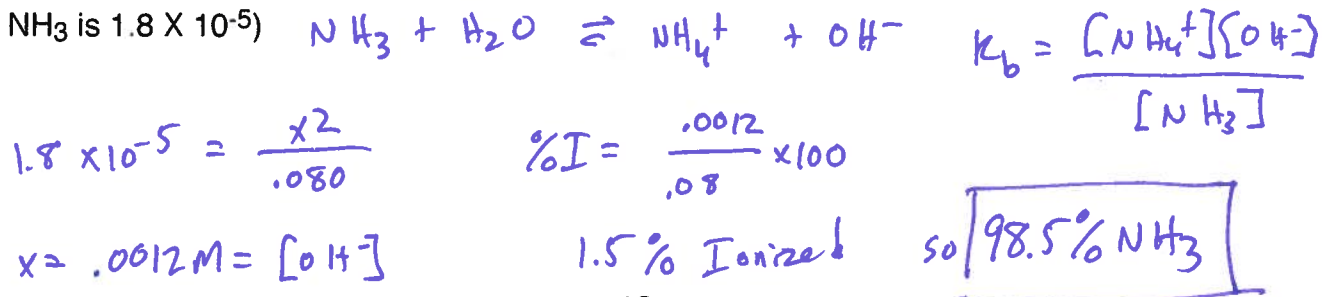
11. Which one of the following is not a weak base in water? Br^- , $\text{C}_2\text{H}_3\text{O}_2^-$, NO_2^-

HBr
S.A. →

12. Assume you have 10.0 liters of a 0.100 M solution of NaOH. How many liters of distilled water should you add to this solution to lower its pH to 11.00?

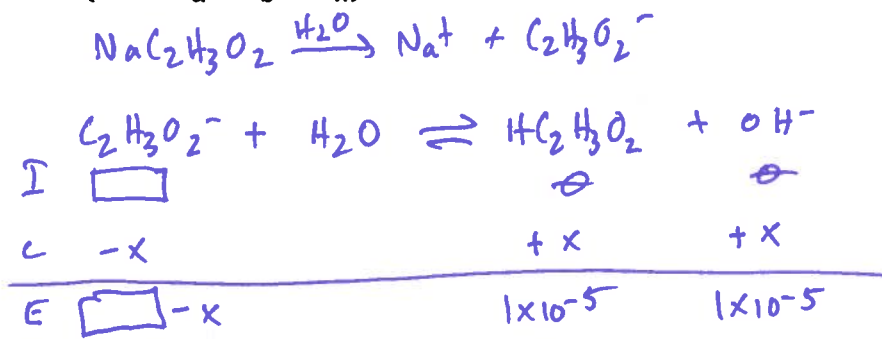
* Ch. 17 material *

13. In a 0.080 M NH_3 solution, what percent of the NH_3 is present as NH_4^+ ? (K_b for NH_3 is 1.8×10^{-5})



14. A solution is made by adding solid sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2$, to enough water to make 3.0 liters of solution. If the solution has a pH of 9.00, how many grams of $\text{NaC}_2\text{H}_3\text{O}_2$ were added to the water? (K_a for acetic acid is 1.8×10^{-5})

{Hint: $K_a \times K_b = K_w$ }



$\text{pH} = 9.00 \therefore \text{pOH} = 5$

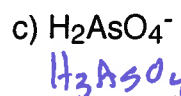
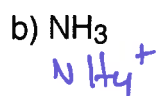
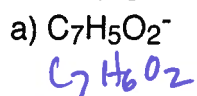
$[\text{OH}^-] = 1 \times 10^{-5}$

$K_b = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$

$5.56 \times 10^{-10} = \frac{(1 \times 10^{-5})^2}{\square}$

15. What is the conjugate acid for the following Bronsted-Lowry bases?

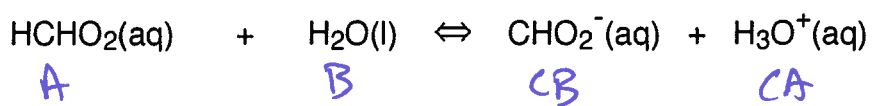
add H^+
ion



$\square = .180 \text{ M} \times 3 \text{ L} =$

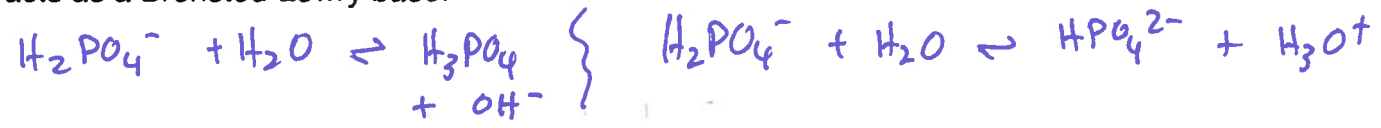
$.540 \text{ mol} \times \frac{82 \text{ g}}{1 \text{ mol}}$

16. Label the Bronsted-Lowry acid, base, conjugate acid, and conjugate base in the following equation:

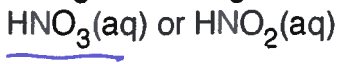


$\square = 44 \text{ g}$

17. Write 2 equations for the reaction of $\text{H}_2\text{PO}_4^-(\text{aq})$ with water. One equation in which $\text{H}_2\text{PO}_4^-(\text{aq})$ acts as a Bronsted-Lowry acid and one equation in which $\text{H}_2\text{PO}_4^-(\text{aq})$ acts as a Bronsted-Lowry base.

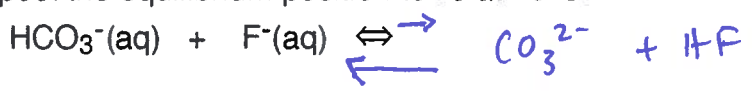


18. Which of the following is a stronger Bronsted-Lowry acid? Why?



look @ K_a values

19. Predict the products of the following acid-base reaction then explain on which side you'd expect the equilibrium position to lie at 25°C.



20. Write the equation which represents the autoionization of water. What is the equilibrium constant for this reaction at 25°C?



21. If the $[\text{OH}^-] = 0.0067 \text{ M}$, calculate the $[\text{H}^+]$ and the pH.

$$1 \times 10^{-14} = (0.0067)(x) \quad x = 1.49 \times 10^{-12}$$

22. Calculate the pH of each of the following strong acid solutions:

a) 0.723 g of HClO_4 in 500.0 mL of solution

$$.723 \text{ g} \times \frac{1 \text{ mol}}{100.5} = \frac{.00719 \text{ mol}}{.5 \text{ L}} = .0144 \text{ M} \quad \text{pH} = -\log(.0144) = 1.84$$

b) 5.00 mL of 1.00 M HCl diluted to 0.200 L

next unit

23. Calculate the pH of a solution formed by mixing 5.00 mL of 0.500 M $\text{Ca}(\text{OH})_2$ with 20.0 mL of 0.100 M NaOH .

next unit

24. Calculate the percent ionization of 1.00 M acetic acid. ($K_a = 1.8 \times 10^{-5}$) How do you think the percent ionization would change for a 0.100 M solution of acetic acid at the same temperature?

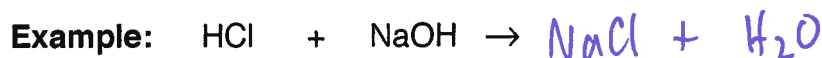
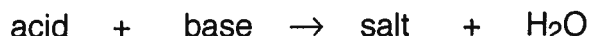
$$\frac{x^2}{1-x} = 1.8 \times 10^{-5} \quad x = .00424 \text{ M} = [\text{H}^+]$$

$$\frac{.00424}{1} \times 100 = .424 \%$$

Target 8: Predict whether a particular salt solution will be acidic, basic, or neutral and write the associated hydrolysis equation. Using the equation, I can calculate the pH of an acidic or basic salt solution.

We assume that when salts dissolve in water, they are completely ionized. Salt solutions can be either acidic, basic, or neutral due to the fact that some ions from the salt can react with water in order to produce an acid or a base. This type of a reaction (a reaction of an anion or cation with water to produce either H⁺ ions or OH⁻ ions) is called salt hydrolysis, or simply **hydrolysis**.

Let's revisit how salts are formed . . .

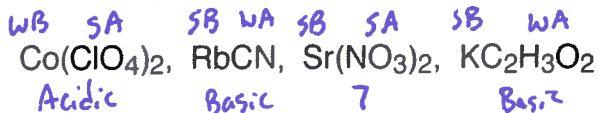


Acid	Base	pH salt solution	Example
Strong	Strong	7	NaCl
Weak	Strong	> 7	NaF
Strong	Weak	< 7	NH ₄ Cl
Weak	Weak	?	NH ₄ F

Practice #1: Predict whether the following salt solutions would be acidic, basic, or neutral.



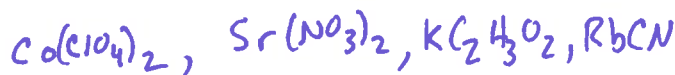
Practice #2: Assume you have 0.10 M solutions of each of the following salts:



List the solutions in order of increasing pH. Hint . . .

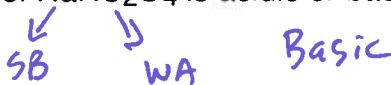
stronger base HCN has a $K_a = 4.9 \times 10^{-10}$

weaker base $\text{HC}_2\text{H}_3\text{O}_2$ has a $K_a = 1.8 \times 10^{-5}$



most acidic $\xrightarrow{\hspace{10em}}$ most basic

Practice #3: Predict whether an aqueous solution of NaHC_2O_4 is acidic or basic. For $\text{H}_2\text{C}_2\text{O}_4$, $K_{a1} = 5.9 \times 10^{-2}$ and K_{a2} is 6.4×10^{-5} .



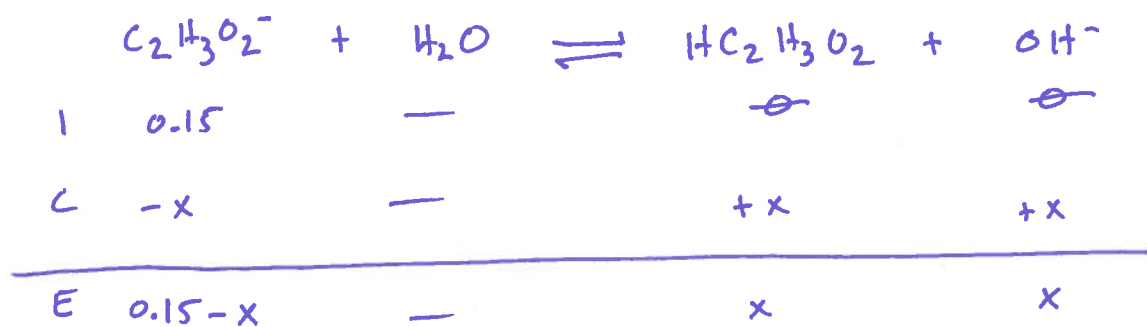
Calculating the pH of salt solutions!

Example #1: Calculate the pH of a 0.15 M solution of sodium acetate. The K_a of acetic acid is 1.8×10^{-5} .

step 1: Write the equation for the initial dissociation of the salt. Determine which of the ions will hydrolyze (if any).



step 2: Write the equilibrium equation for the hydrolysis of the acetate ion. Construct a table summarizing the changes in concentration.



step 3: Choose an appropriate equilibrium constant (K_a or K_b) to use. Solve for either the $[\text{OH}^-]$ or $[\text{H}^+]$. producing OH^- so use K_b !

$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = \frac{x^2}{0.15 - x} = 5.6 \times 10^{-10}$$

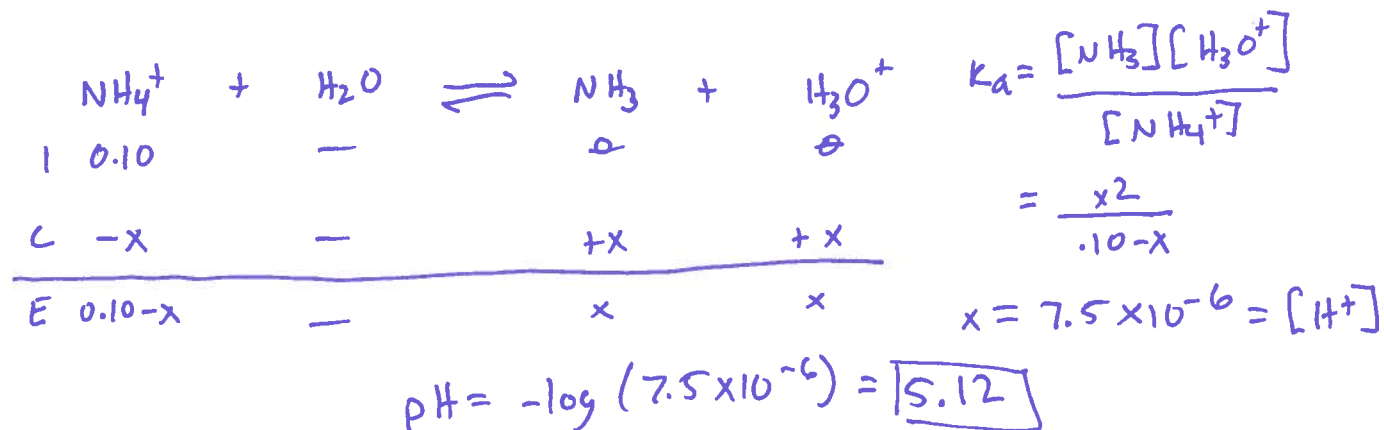
$$K_b = \frac{1 \times 10^{-14}}{K_a} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = x = 9.20 \times 10^{-6} \text{ M} = [\text{OH}^-]$$

step 4: Solve for the pH of the salt solution.

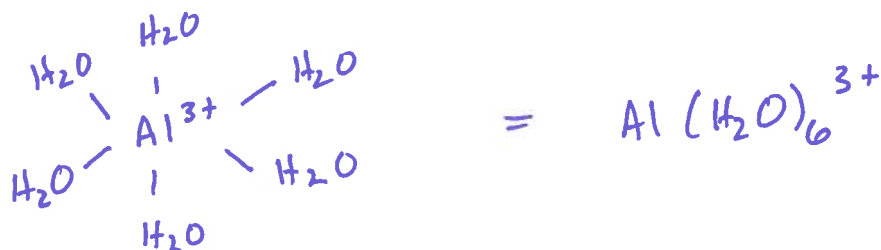
$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ &= -\log (9.2 \times 10^{-6}) \\ &= 5.04 \end{aligned}$$

$$\begin{aligned} \text{pH} &= 14 - 5.04 \\ &= 8.96 \end{aligned}$$

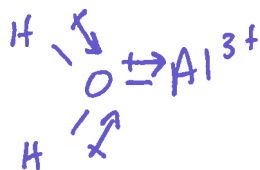
Practice: What is the pH of a 0.10 M NH_4Cl solution? The K_a for NH_4^+ is 5.6×10^{-10} .



Some salts contain small, highly charged metal ions (for example, Al^{3+} , Cr^{3+} , Fe^{3+} , Bi^{3+} , and Be^{2+}) which produce an acidic solution. For example, when aluminum chloride (AlCl_3) dissolves in water, the Al^{3+} ions take the hydrated form $\text{Al}(\text{H}_2\text{O})_6^{3+}$. Because metal ions are positively charged, they attract the unshared pair of electrons from the water molecules. (How did we know to use 6 water molecules as ligands).



Let's consider one bond between the metal ion and the oxygen atom of one of the six water molecules.



The positively charged Al^{3+} ion draws electron density toward itself, making the O-H bond more polar. Consequently, the H atoms have a greater tendency to ionize than those in water molecules not involved in hydration. The resulting ionization process can be written as . . .



Note that the production of H^+ by the hydrated metal ions allows us to classify these metal ions as acids.

The equilibrium expression for the metal ion hydrolysis is given by

$$K_a = \frac{[Al(OH)(H_2O)_5^{2+}][H^+]}{[Al(H_2O)_6^{3+}]}$$

The complex ion could undergo further hydrolysis, although we usually do not consider it. Also realize that the greater the charge and the smaller the radius of the metal ion, the more polar the bond. The more polar the bond, the further it ionizes and K_a increases. For example, Fe^{3+} is more highly charged and smaller than Cu^{2+} and is therefore more acidic!

Target 9: Explain how acid strength relates to polarity and the strength of the H-X bond.

There are several factors which determine the strength of an acid. Factors such as the properties of the solvent, temperature, and molecular structure. We will focus our attention on the relationship between the strength of an acid and its structure. We will discuss the factors which affect both **binary acids** (contain only 2 elements) and **ternary acids** (contain 3 different elements). Two of the major factors are bond polarity and bond dissociation energy of the H-X bond.

Binary Acids- the major factor in determining acid strength for a series of acids which all contain elements from the same family is the bond dissociation energy.

Bond	bond diss. energy (kJ/mol)
H-F	568
H-Cl	432
H-Br	366
H-I	298

The strength of these acids are . . . $HF \ll HCl < HBr < HI$

The more tightly held the hydrogen, the weaker the acid.

Now let's consider a series of hydrogen compounds in a particular period on the periodic table.



This trend can be explained on the basis of bond polarity. As the H-X bond becomes more polar, the compound has a greater tendency to ionize. Note that relative importance of bond polarity and bond energy is reversed when compared to the trend of binary hydrogen compounds in a family.

Target 10: Predict the relative strengths of oxyacids.

Ternary Acids- Ternary acids will always contain hydrogen, oxygen, and a nonmetal which will occupy a central position in the structure of the acid. The strength of a ternary acid, sometimes called an oxyacid, depends upon the tendency for an O-H bond to ionize.



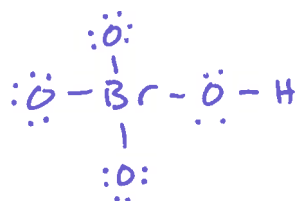
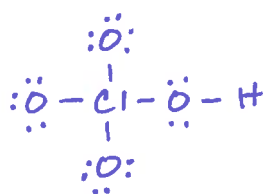
Any factor that draws electrons toward the Z (central) atom strengthens the Z-O bond and weakens the O-H bond. The easier it is to break the O-H bond, the stronger the acid. There are 2 factors which can strengthen the Z-O bond:

- 1) Z having a high electronegativity, and
- 2) Z having a high oxidation state.

Both of these factors tend to pull electrons toward Z. *and thus weaken the O-H bond*

Let's divide the ternary acids into 2 groups:

group 1 - Oxyacids which have different central atoms from the same group of the periodic table and have the same oxidation number.

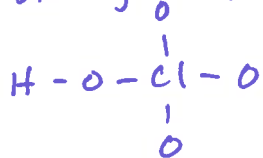


The strength of the acid increases with an increase in electronegativity of the central atom

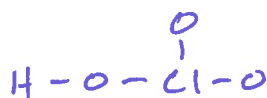
Cl-O bond is stronger than the Br-O bond, therefore the perchloric acid has a weaker O-H bond, and is therefore a stronger acid!

group 2 - Oxyacids having the same central atom, but different numbers of oxygens.

★ strongest ★



Cl = +7



Cl = +5



Cl = +3

weakest



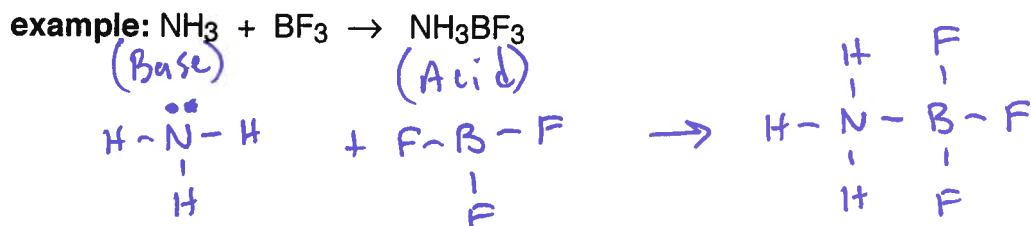
Cl = +1

As you can see, as the number of oxygens increase, the oxidation number also increases. The higher the oxidation number, the the stronger the Cl-O bond in the Cl-O-H group, and therefore the weaker the O-H bond.

Target 11: Define an acid and a base in terms of the Lewis concept.

Lewis acid = electron-pair acceptor

Lewis base = electron-pair donor



The Lewis definition of acids and bases is *broader* than the Bronsted-Lowry definition as the Lewis definition includes reactions where there is no proton transfer. Everything that is a base in the Bronsted-Lowry definition (a proton acceptor) is also a base using the Lewis definition (electron-pair donor). The difference lies in the fact that a Lewis base can donate electron pairs to other things besides H^+ .

Practice: Identify the Lewis acid and base in each of the following reactions.

- $\text{SnCl}_4(\text{s}) + 2\text{Cl}^-(\text{aq}) \rightleftharpoons \text{SnCl}_6^{2-}(\text{aq})$
A B
- $\text{Hg}^{2+}(\text{aq}) + 4\text{CN}^-(\text{aq}) \rightleftharpoons \text{Hg}(\text{CN})_4^{2-}(\text{aq})$
A B
- $6\text{NH}_3(\text{aq}) + \text{Co}^{3+}(\text{aq}) \rightleftharpoons \text{Co}(\text{NH}_3)_6^{3+}(\text{aq})$
B A

Keys to identifying Lewis acids and bases

Lewis acids = often times + charged or are metals

Lewis bases = often times - charged or species that can become + ions