

Lecture Notes for Chapter 15: Acids and Bases

I. Acids and Bases

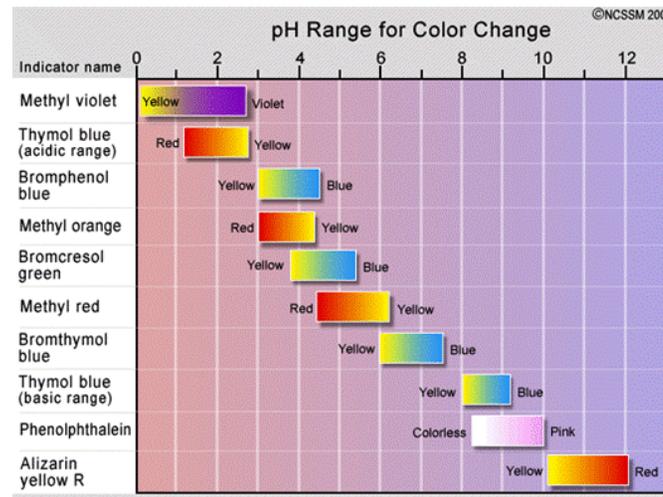
- There are several ways to define acids and bases. Perhaps the easiest way to start is to list some of the properties of acids and bases.
- The table below summarizes some properties that will be helpful as we learn more about acids and bases. In general, though, it is useful to say that:
 - Acids** are substances which donate hydrogen and form lots of H_3O^+ (hydronium) ions when dissolved in water
 - Bases** are substances which accept hydrogen and form lots of OH^- (hydroxide) ions in solution.

Property	Acids	Neutral Solutions	Bases
Taste	Sour taste	-	Bitter taste
Feel	-	-	Slippery
Litmus Test	Turns blue litmus red	Will not change the color of litmus	Turns red litmus blue
Phenolphthalein Test	Colorless in phenolphthalein	Colorless in phenolphthalein	Turns phenolphthalein pink
Arrhenius Definition	Produces H^+ ions in aqueous solution	-	Produces OH^- ions in aqueous solution
Brønsted-Lowry Definition	Donates protons	-	Accepts protons
Lewis Definition	Electron pair acceptor	-	Electron pair donor
Hydroxide ion concentration	$[\text{OH}^-] < 10^{-7} \text{ M}$ SMALL	$[\text{OH}^-] = 10^{-7} \text{ M}$ MEDIUM	$[\text{OH}^-] > 10^{-7} \text{ M}$ LARGE
Hydronium ion concentration	$[\text{H}_3\text{O}^+] > 10^{-7} \text{ M}$ LARGE	$[\text{H}_3\text{O}^+] = 10^{-7} \text{ M}$ MEDIUM	$[\text{H}_3\text{O}^+] < 10^{-7} \text{ M}$ SMALL
pH	pH < 7	pH = 7	pH > 7
pOH	pOH > 7	pOH = 7	pOH < 7

II. What are Acids and Bases?

- Taste & feel
- Are acids necessarily dangerous? And are bases therefore necessarily safe?
- Litmus** as an indicator
- Phenolphthalein** as an indicator
- In chemistry, an **indicator** is a substance that changes color when the pH of its environment is changed. If you dissolve an indicator in a solution, it will change colors when the solution's pH changes. At what pH value will the indicator change its color? Answer: it depends on the

indicator. This is called the **pH range** of the indicator.



f. Arrhenius definition of acids and bases

- Svante Arrhenius: Swedish guy
- Suggested that acids produced H^+ ions in solution, and that bases produced OH^- ions in solution
- Vocab moment: H^+ ion is called a **hydrogen ion**. OH^- is called a **hydroxide ion**.
- GOOD: Explained for the first time many behaviors of acids and bases, especially their ability to neutralize one another in aqueous solution:



v. SHORTCOMINGS:

- Could not account for why some acids don't contain H: why does CO_2 produce acidic solutions when dissolved in water? Why isn't methane (CH_4) a good acid, even though it contains H?
- Why do some bases such as NH_3 (ammonia) not contain the hydroxide ion? Why can NH_3 act as a base or an acid, depending on the situation?
- ACIDS form H^+ (hydrogen) ions in solution.
- BASES form OH^- (hydroxide) ions in solution.

g. Brønsted-Lowry definition of acids and bases

- Danish guy and English guy – developed theories independently
- For our chemistry course, this is the most useful and important theory of acids and bases
- First, let's define what we mean by "proton" when we speak of donating and accepting protons.

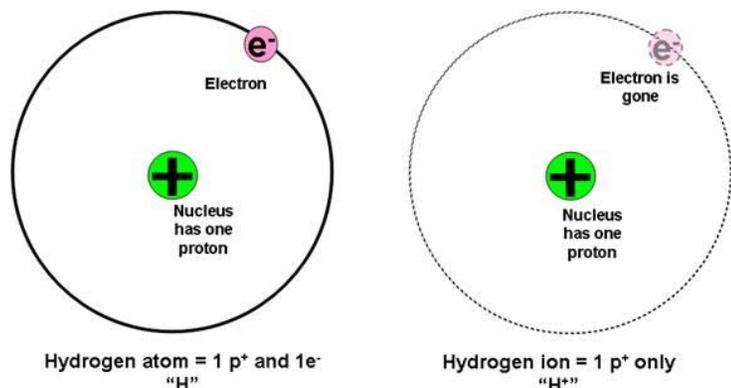


Figure 1: The difference between a hydrogen atom (H) and a hydrogen ion (H⁺). A hydrogen ion is the same thing as a proton.

iv. The Brønsted-Lowry definition of acids: ACIDS are proton (H⁺) donors.

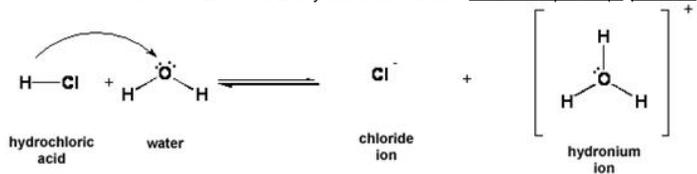
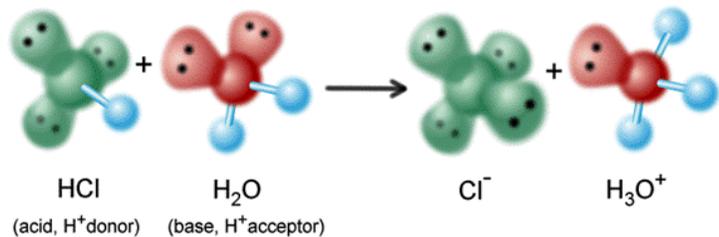


Figure 2: HCl donates H⁺ (a proton) to water. HCl is an example of a Brønsted-Lowry acid.



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v. The Brønsted-Lowry definition of bases: BASES are proton (H⁺) acceptors.

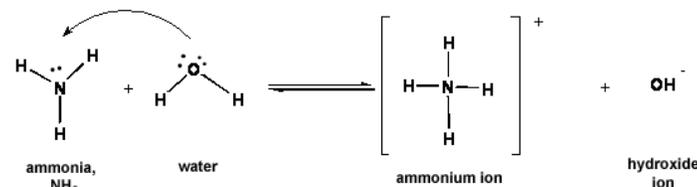
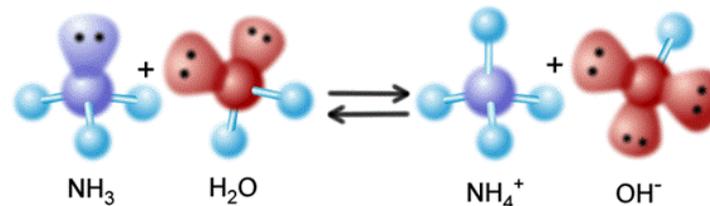


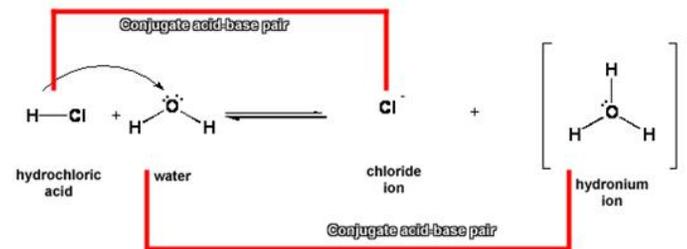
Figure 3: Ammonia (NH₃) accepts a proton from water to make the ammonium ion. NH₃ is an example of a Brønsted-Lowry base.



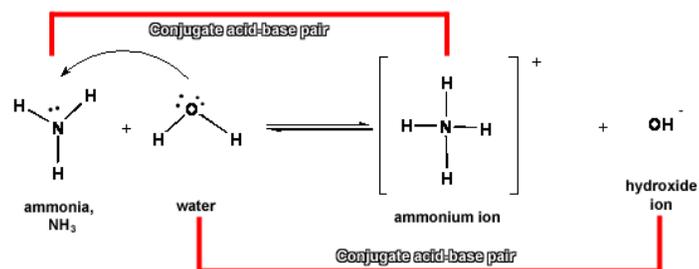
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vi. Conjugate acids and bases

- Substances on opposite sides of the equation that are one proton (H⁺) different from each other are called **conjugate acid-base pairs**.
- In the reaction between HCl and water, HCl and Cl⁻ are a conjugate acid-base pair. The other conjugate acid-base pair is H₂O/H₃O⁺.
 - HCl is the conjugate acid of Cl⁻, and Cl⁻ is the conjugate base of HCl.
 - H₃O⁺ is the conjugate acid of H₂O, and H₂O is the conjugate base of H₃O⁺.



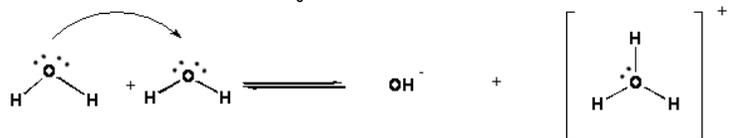
- In the reaction between NH₃ and water, ammonia/ammonium ion are a conjugate acid-base pair, and H₂O/OH⁻ are the other conjugate acid-base pair.
 - NH₃ is the conjugate base of NH₄⁺, and NH₄⁺ is the conjugate acid of NH₃.
 - H₂O is the conjugate acid of OH⁻, and OH⁻ is the conjugate base of H₃O⁺.



- h. Lewis definition of acids and bases
- This is the best and most current theory. However, for our chemistry course, this is the least important theory of acids and bases. The Lewis theory will be very useful in Biology class and most college chemistry classes, though.
 - ACIDS are electron pair acceptors
 - BASES are electron pair donors

III. The Self-Ionization of Water and the Ion Product Constant for water, K_w

- Water can act as an acid or as a base.**
- As you can see from the reactions above, water can act as either an acid or a base, depending on the situation. Such a substance is said to be "amphoteric" or "amphiprotic", since it can either accept or donate a proton under various conditions.
- Pure water is not pure H_2O ! Can you believe it?!? The water reacts with itself to make the OH^- ion (hydroxide ion) and the H_3O^+ ion (hydronium ion).

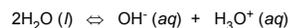


- d. The concentrations of these ions in pure water at 25°C is always

$$[\text{OH}^-] = [\text{H}_3\text{O}^+] = 1.00 \times 10^{-7} \text{ M}$$

***Neutral solutions and pure water ONLY!**

- e. Also, we can write this chemical reaction as



Therefore, we can write the equilibrium constant (K_{eq}) expression for this reaction as

$$K_{\text{eq}} = [\text{OH}^-][\text{H}_3\text{O}^+] = [10^{-7} \text{ M}][10^{-7} \text{ M}] = 10^{-14}$$

Notice that we have left H_2O out of the K_{eq} expression because it is a liquid.

- f. This equilibrium constant has a special name; it is called the **ion product constant of water**. It is unitless. It is given a special symbol, too: K_w . This product will be true for any aqueous solution – not

just pure water – at 25°C , which is the only temperature which we will consider. **Acidic, basic, and neutral aqueous solutions always have an ion product constant that is equal to 1.00×10^{-14} .**

- g. The following is an important formula:

$$K_w = 1.00 \times 10^{-14}$$

IV. The pH Scale

- What is pH?
- Introduction to the logarithm, or \log_{10}

Complete the table below. The pH and pOH columns will be explained after the notes that follow:

[OH ⁻]	[H ₃ O ⁺]	Acidic, Basic, or Neutral Solution	pH	pOH
$1.0 \times 10^{-7} \text{ M}$				
	$1.0 \times 10^{-9} \text{ M}$			
$1.0 \times 10^{-5} \text{ M}$				
	$6.2 \times 10^{-13} \text{ M}$			
$9.3 \times 10^{-9} \text{ M}$				

Solve for the missing value:

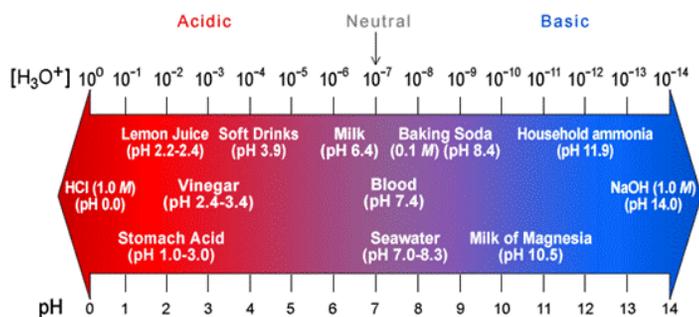
Problem	Answer	Problem	Answer
$10^3 = ?$		$10^{-7} = ?$	
$10^2 = 1000$		$10^2 = 0.0001$	
$10^{-4} = ?$		$10^{-3} = ?$	
$10^2 = 0.1$		$10^2 = 0.0368$	

- c. The pH scale is a logarithmic scale

- This means that the pH of a solution for which $[\text{H}_3\text{O}^+] = 0.01 \text{ M}$ is $\text{pH}=2$, but the pH for a solution for which $[\text{H}_3\text{O}^+] = 0.000001 \text{ M}$ is $\text{pH} = 6$.
- Which solution is more acidic? Clearly the solution for which $[\text{H}_3\text{O}^+] = 0.01 \text{ M}$.
- How many times more acidic is it? Simply divide the two $[\text{H}_3\text{O}^+]$ values. The answer is $0.01 \div 0.000001 = 10^{-2} \div 10^{-6} = 10,000$ times as acidic.
- Common mistake:** many students will incorrectly say that the $\text{pH}=6$ solution is more acidic. In fact they will incorrectly say that it is three times as acidic.
- Common mistake:** many students will correctly say that the $\text{pH}=2$ solution is more acidic, but they will incorrectly say that it is three times as acidic.
- The pH scale is logarithmic**, which is to say that pH units are each 10 times further apart from each other as you go up or down the scale. That is, DO NOT divide $(\text{pH}=6) \div (\text{pH}=2)$ to determine how many times more acidic or basic one solution is than another. You should instead realize that those pH units represent **powers of 10**. Thus, the correct calculation is, as shown above, $10^{-2} \div 10^{-6}$.
- You should instead realize that a change in pH of 1 unit represents a **ten-fold increase in the concentration of $[\text{H}_3\text{O}^+]$** , a change in pH of 2 units represents a **one hundred-fold increase in the concentration of $[\text{H}_3\text{O}^+]$** , a change in pH of 3 units represents a **one thousand-fold increase in the concentration of $[\text{H}_3\text{O}^+]$** , etc.

- viii. The same logic is true for pOH calculations. You should instead realize that a change in pOH of 1 unit represents a **ten-fold increase in the concentration of [OH⁻]**, a change in pOH of 2 units represents a **one hundred-fold increase in the concentration of [OH⁻]**, a change in pOH of 3 units represents a **one thousand-fold increase in the concentration of [OH⁻]**, etc.
- d. The pH scale ranges from low numbers (acidic) to high numbers (basic).
- e. Negative pH is possible, so is a pH = 0. (Calculate the pH of a solution of [H₃O⁺] = 1.00 M. Calculate the pH of a solution of [H₃O⁺] = 10.00 M.) pH and pOH values do not have to be whole numbers.
- f. The further a pH value is from the neutral value of pH=7, the more acidic or basic that solution is.
- For instance, pH=10.2 is more basic than pH=8.11.
 - For instance, pH = 3.09 is more acidic than pH = 6.92.

pH Scale



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- g. Sometimes it is useful to keep track of pOH
- h. The four important equations:
- [H₃O⁺][OH⁻] = 1 × 10⁻¹⁴ = K_w
 - pH = -log [H₃O⁺], also written as pH = -log [H⁺]
 - pOH = -log [OH⁻]
 - pH + pOH = 14
- i. Example problem: Determining [OH⁻] from [H₃O⁺].

⑥ If [H₃O⁺] = 1 × 10⁻⁴ M, find [OH⁻].

$$[H_3O^+][OH^-] = 1.00 \times 10^{-14}$$

$$[1 \times 10^{-4}][OH^-] = 1.00 \times 10^{-14}$$

$$[OH^-] = \frac{1.00 \times 10^{-14}}{1 \times 10^{-4}} = 1 \times 10^{-10} M$$

⑦ Is this an acidic or basic solution?

It is acidic, because [H₃O⁺] > 10⁻⁷ M (and because [OH⁻] < 10⁻⁷ M).

- j. Example problem: Determining [OH⁻] from [H₃O⁺]. This time a calculator is probably required.

⑥ If [H₃O⁺] = 7.5 × 10⁻⁸ M, find [OH⁻].

$$[H_3O^+][OH^-] = 1.00 \times 10^{-14}$$

$$[OH^-] = \frac{1.00 \times 10^{-14}}{7.5 \times 10^{-8}} = 1.3 \times 10^{-7} M$$

⑦ Is this solution acidic?

No, it is basic, because [H₃O⁺] < 10⁻⁷ M. (or, you could say that [OH⁻] > 10⁻⁷ M.)

- k. Example problem: Determining pH from [H₃O⁺].

Find the pH for a solution for which [H₃O⁺] = 1 × 10⁻⁴ M.

$$pH = -\log [H_3O^+] = -\log [1 \times 10^{-4}]$$

$$= -[-4]$$

$$= 4$$

- l. Example problem: Determining pH from [H₃O⁺]. This time the concentration is not simply a power of ten, so you will need a calculator.

Find the pH for a solution for which [H₃O⁺] = 7.5 × 10⁻⁸ M.

$$pH = -\log [H_3O^+] = -\log [7.5 \times 10^{-8} M]$$

$$= -[-7.12]$$

$$= 7.12$$

- m. Example problem: Determining pOH from [OH⁻].

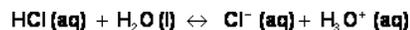
Find the pOH for a solution for which $[OH^-] = 1 \times 10^{-10} M$
 $pOH = -\log [OH^-] = -[-10] = 10$

- n. Example problem: Determining pOH from $[H_3O^+]$.

Find the pOH for a solution for which the $[H_3O^+] = 7.5 \times 10^{-8} M$. We earlier found that $pH = 7.12$ (see earlier example!).
 Thus, $pOH + pH = 14$
 $pOH = 14 - pH$
 $pOH = 14 - 7.12$
 $pOH = 6.88$

V. Acid Strength and Base Strength

- a. First, it is important to emphasize the difference between the terms **strong** and **concentrated**. When a person tastes something that is very flavorful, such as a glass of concentrated Kool-Aid, he or she might remark "Wow, that's pretty strong." By this he or she means that there appears to be a lot of drink mix (solute) dissolved in a relatively small amount of water (solvent). As chemists, though, we would say that the Kool-Aid is **concentrated**, not strong.
- CONCENTRATED = high molarity
 - STRONG = high degree of ionization
 - DILUTE = low molarity
 - WEAK = low degree of ionization
- b. Some acids are strong, other acids are weak.
- Some acids, such as hydrochloric acid, will **completely react** with water. HCl is therefore a STRONG acid. HCl will dissociate into H^+ and Cl^- ions, with the hydrogen ion quickly jumping onto a water molecule to make H_3O^+ :



Before Reaction :

100% 0%

After Reaction :

0% 100%

- No matter how much you dilute HCl with water, it will always be a strong acid. If you dilute HCl, you will have created a dilute, strong acid solution.
- On the other hand, some acids, such as acetic acid, will only partially react with water. Acetic acid is therefore a WEAK acid. CH_3COOH will dissociate into H^+ and CH_3COO^- (acetate) ions, with the hydrogen ion quickly jumping onto a water molecule to make H_3O^+ :



Before Reaction :

100% 0%

After Reaction :

99% 1%

- No matter how concentrated you manage to make a solution of acetic acid, it will always be a weak acid. If you evaporate or otherwise remove most of the water from an acetic acid solution, you will have created a concentrated, weak acid solution.
- c. Base Strength
- Bases can be either strong or weak, too.
 - Sodium hydroxide ionizes 100% to make Na^+ and OH^- ions, so it is an example of a **strong base**.
 - No matter how much you dilute NaOH with water, it will always be a strong base. If you dilute NaOH, you will have created a dilute, strong base solution.
 - Ammonia (NH_3) reacts with water (see equation illustrated earlier in these notes) only partially. Ammonia is a **weak base**, just like acetic acid is a weak acid. Out of every 100 molecules of ammonia that dissolves into water, only about one has reacted to make an ammonium ion (NH_4^+) at any one time.
 - No matter how concentrated you manage to make a solution of ammonia, it will always be a weak base. If you evaporate or otherwise remove most of the water from an ammonia solution, you will have created a concentrated, weak base solution.
- d. So, to summarize this section, we can say that **STRONG \neq CONCENTRATED**, necessarily, and **WEAK \neq DILUTE**, necessarily. It is possible to be both strong and concentrated, but it is also possible to be strong and dilute at the same time. It is possible to be both weak and dilute, but it is also possible to be strong and dilute at the same time. Here is a table showing examples of these combinations:

Example: Acids		
ACID STRENGTH →	STRONG	WEAK
ACID CONCENTRATION ↓		
CONCENTRATED	12.0 M HCl (hydrochloric acid) "12.0 M HCl is an example of a concentrated, strong acid solution."	12.0 M HC ₂ H ₃ O ₂ (acetic acid) "12.0 M HC ₂ H ₃ O ₂ is an example of a concentrated, weak acid solution."
DILUTE	0.1 M HCl (hydrochloric acid) "0.1 M HCl is an example of a dilute, strong acid solution."	0.1 M HC ₂ H ₃ O ₂ (acetic acid) "0.1 M HCl is an example of a dilute, weak acid solution."

Example: Bases		
BASE STRENGTH →	STRONG	WEAK
BASE CONCENTRATION ↓		
CONCENTRATED	12.0 M NaOH (sodium hydroxide) "12.0 M NaOH is an example of a concentrated, strong base solution."	12.0 M NH ₃ (ammonia) "12.0 M NH ₃ is an example of a concentrated, weak base solution."
DILUTE	0.1 M NaOH (sodium hydroxide) "0.1 M NaOH is an example of a dilute, strong base solution."	0.1 M NH ₃ (ammonia) "0.1 M NH ₃ is an example of a dilute, weak base solution."

VI. Calculating the pH of Strong Acid and Strong Base Solutions

- For strong acid solutions, [H₃O⁺] = the initial concentration of the acid. Why? The answer is that because every molecule of a strong acid falls apart (dissociates, ionizes) into a H₃O⁺ ion, the concentration of an acid such as HCl is the same as the concentration of H₃O⁺ ions that it produces.
- For this reason, the pH of a strong acid is found according to this formula:

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

Also, the pOH of strong bases can be determined in a similar way:

$$\text{pOH} = -\log [\text{OH}^-] = -\log [\text{STRONG base}]$$

- Example: Find the pH of a 0.0100 M solution of HCl.
Answer: HCl is a strong, monoprotic acid. Therefore, at equilibrium, all of the HCl will disappear and produce H₃O⁺ ions in a 1:1 ratio. This means that the concentration of H₃O⁺ at equilibrium = the

concentration of the HCl. Therefore,
 $\text{pH} = -\log [\text{H}_3\text{O}^+]$
 $= -\log [0.0100]$
 $= -[-2]$
 $= 2$

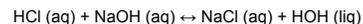
- Example: Find the pOH of a 0.0100 M solution of the strong base NaOH.
Answer: NaOH a strong base. Therefore, at equilibrium, all of the NaOH will disappear and produce OH⁻ ions in a 1:1 ratio. This means that the concentration of OH⁻ at equilibrium = the concentration of the NaOH. Therefore,
 $\text{pOH} = -\log [\text{OH}^-]$
 $= -\log [0.0100]$
 $= -[-2]$
 $= 2$

Now, what is the pH of this 0.0100 M solution of NaOH?

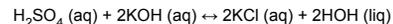
Answer: $\text{pH} + \text{pOH} = 14$
 $\text{pH} + 2 = 14$
 $\text{pH} = 14 - 2$
 $\text{pH} = 12$

VII. Neutralization

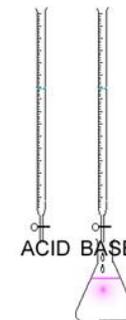
- The reaction that takes place between acids and bases is called **neutralization**.
- An acid + a base yields a **salt plus water**.
- Example of a neutralization reaction: hydrochloric acid plus sodium hydroxide.



- Another example of a neutralization reaction: sulfuric acid plus potassium hydroxide.



- A lab procedure that uses a known volume of a solution of known concentration to determine the concentration of another solution is called a **titration**. The titration is carried out using two burets and a flask.
- The titration has reached its **endpoint** when the number of moles of H⁺ from the acid and the moles of OH⁻ from the base have exactly neutralized one another. At the endpoint, moles of H⁺ = moles of OH⁻ in the flask.
- An **indicator** such as phenolphthalein is used to signal when the titration has reached its endpoint. In this case, the phenolphthalein will turn pink when the solution reaches the endpoint.



- Example problem: determining the number of moles of H⁺ in a sample of an acid.

How many moles of H^+ are in 2.00 moles of HCl ?

$$2.00 \text{ moles } HCl \times \frac{1 \text{ mol } H^+}{1 \text{ mol } HCl} = 2.00 \text{ mol } H^+$$

(moles of H^+ = moles of acid for a monoprotic acid.)

- i. Example problem: determining the number of moles of H^+ in a sample of an acid. This time the acid is a triprotic acid, so the number of moles of H^+ is not the same as the number of moles of acid.

How many moles of H^+ are in 2.00 moles of H_3PO_4 ?

$$2.00 \text{ mol } H_3PO_4 \times \frac{3 \text{ mol } H^+}{1 \text{ mol } H_3PO_4} = 6.00 \text{ mol } H^+$$

- j. Example problem: determining the number of moles of H^+ in an acid solution. This time the acid is dissolved in water, so the number of moles of H^+ must be determined by using the molarity formula, $M=n/V$.

How many moles of H^+ are in 2.00 liters of 0.100 M $HCl(aq)$?

$$M = \frac{n}{V}$$

$$0.100 \frac{\text{mol}}{\text{liter}} = \frac{n}{2.00 \text{ liters}}$$

$$0.200 \text{ mol} = n$$

- k. Example problem: titration of an acid with a base. Solve this the same way that you would solve any solution stoichiometry problem. You can see that I used the "V-M-n" table that you see below to keep track of volume, molarity (concentration) and number of moles. I also used the formula $M=n/V$.

ml	V	ml
M	M	M
mol	n	mol

How much 0.100 M $NaOH$ will be required to neutralize 50.9 ml of 0.150 M HCl ?

$$NaOH + HCl \rightarrow H_2O + NaCl$$

(E) $\begin{array}{|c|c|c|} \hline 0.07635 & V & .0509 \text{ L} \\ \hline 0.100 \text{ M} & M & .150 \frac{\text{mol}}{\text{L}} \\ \hline \end{array}$
 (A) $\begin{array}{|c|c|c|} \hline 0.007635 & n & 0.007635 \text{ mol} \\ \hline \end{array}$
 (B) $M = \frac{n}{V} \Rightarrow 0.150 \frac{\text{mol}}{\text{L}} = \frac{n}{0.0509 \text{ L}} \Rightarrow n = 0.007635$
 (C) $0.100 \text{ M} = \frac{0.007635 \text{ mol}}{V} \Rightarrow V = \frac{0.007635}{0.100} = 0.07635 \text{ L} = 76.4 \text{ mL}$

VIII. Buffered Solutions