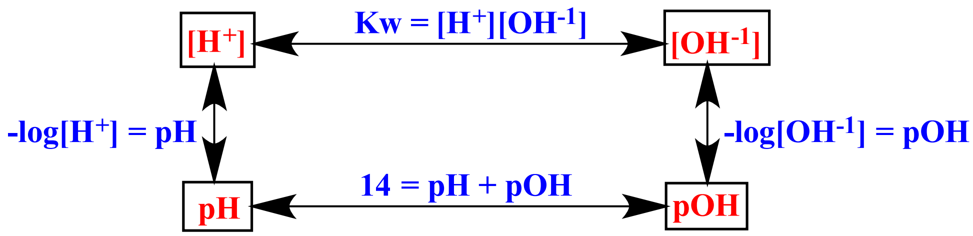
<u>Acid/Base Video 1 Review Sheet Key:</u>



1) Given the $[OH^{-1}] = 1.21 \times 10^{-9}$ M, please calculate the pH.

Decided to convert given hydroxide concentration to proton concentration then to pH:

Given: $[OH^{-1}] = 1.21 \times 10^{-9}$ M

$\boxed{[H^+]}$ $Kw = 1 \times 10^{-14} = [H^+][OH^{-1}]$

rearrange

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^{-1}]} = \frac{1 \times 10^{-14}}{(1.21 \times 10^{-9})} = \underline{8.26 \times 10^{-6}}$$

$\boxed{pH}$ $pH = -\log[H^+] = -\log(8.26 \times 10^{-6}) = \underline{5.083}$

Note: <u>3</u> significant figures after the decimal point are required after taking the log of <u>3</u> significant figures.

2) Given a 0.10M HNO_3 solution (strong acid), please calculate pH, pOH, $[H^+]$, and $[OH^-]$:

A strong acid; thus, HNO_3 is 100% dissociated as shown below, which gives proton concentration, followed by pH calculation, followed by pOH calculation, and hydroxide concentration:

	$HNO_{3(aq)}$ $\longrightarrow$	$H^+_{(aq)}$ +	$NO^-_{3(aq)}$
t=0	0.10M	0	0
$t_{eq.}$	0	**0.10M**	0.10M

$\boxed{[H^+]}$ **Thus, $[H^+] = \underline{0.10M}$**

$\boxed{pH}$ $pH = -\log[H^+] = -\log(0.10) = \underline{1.00}$

$\boxed{pOH}$ $pOH = 14 - 1.00 = \underline{13.00}$

$\boxed{[OH^-]}$ $Kw = 1 \times 10^{-14} = [H^+][OH^{-1}]$

rearrange

$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{0.10} = \underline{1.0 \times 10^{-13}}$$

3) Given a 0.18M NaOH solution (strong Base), please calculate pH, pOH, $[H^+]$, and $[OH^-]$:

NaOH is a strong base; thus, NaOH is 100% dissociated as shown below, which gives hydroxide concentration, followed by proton calculation, followed by pH calculation, and pOH:

	$NaOH_{(s)}$	$\rightarrow Na^+_{(aq)}$	$+\ OH^-_{(aq)}$
t=0	0.18M	0	0
$t_{eq.}$	0	0.18M	**0.18M**

$\boxed{[OH^-]}$ **Thus, $[OH^{-1}]$ = <u>0.18M</u>**

$\boxed{[H^+]}$

$Kw = 1 \times 10^{-14} = [H^+][OH^{-1}]$

rearrange

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^{-1}]} = \frac{1 \times 10^{-14}}{(0.18)} = \underline{5.6 \times 10^{-14}}$$

$\boxed{pH}$ **$pH = -\log[H^+] = -\log(5.6 \times 10^{-14}) = \underline{13.25}$**

$\boxed{pOH}$ **$pOH = -\log[OH^-] = -\log(0.18) = \underline{.74}$**

4) Please identify the acid, base, conjugate acid, and conjugate base as well as the acid-conjugate base and base-conjugate acid pairs for the following equilibrium:

Acid-Conjugate Base Pair

$$\underset{\text{Acid}}{HF_{(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Conjugate Acid}}{H_3O^+_{(aq)}} + \underset{\text{Conjugate Base}}{F^-_{(aq)}}$$

Base-Conjugate Acid Pair

<u>Circle the correct underlined word(s) that makes the statement correct:</u>

5) A **<u>strong / weak</u>** conjugate base is produced when a stable conjugate base is formed.

If a stable conjugate base is formed it is low in energy, stable, and unreactive; thus, it is called a "WEAK" conjugate base.

6) A(n) **<u>stable / unstable</u>** conjugate base is formed when a weak acid is dissolved within water.

If a weak acid dissociates its conjugate base is high in energy, reactive, and "UNSTABLE".

7) A conjugate base **<u>low / high</u>** in energy is produced when the strong acid HCl is dissociated.

When a strong acid dissociates it yields a conjugate base that is unreactive, stable, and "LOW" in energy.

8) A **<u>conjugate acid low in energy / conjugate base high in energy</u>** is produced when acetic acid (a weak acid) dissociates.

If a weak acid dissociates its conjugate base is reactive, unstable ands it is said to produce a "CONJUGATE BASE HIGH IN ENERGY".

<u>TRUE or FALSE: Circle T or F</u>

9) If the pH is equal to 4.5 then the proton concentration is less than the hydroxide concentration. **T** or **F**

If pH is less than 7 (given 4.5), then the $[H^+] > [OH^-]$; thus, FALSE.

10) If the pOH is equal to 2 then the solution is acidic. **T** or **F**

If pOH = 2, then pH = 12, then basic and not acidic; thus, FALSE.

11) The pH of orange juice is 4. If a solution of vinegar is added with a pH of 4.5 then the pH of the final solution is 8.5. **T** or **F**

pH values are not additive; i.e., adding two acidic solutions together will stay acidic not turn basic with pH of 8.5; thus, FALSE.

12) If the concentration of protons is equal to the concentration of hydroxide ions then the solution should have a pH of approximately 0.0. **T** or **F**

When proton and hydroxide concentrations are equal the solution is neutral or pH of 7; thus, FALSE.

13) If the concentration of protons is less than the concentration of hydroxide ions then the solution should have a pH less than 7.0. **T** or **F**

When $[H^+] < [OH^-]$ solution is basic with pH > 7; thus, FALSE.

14) If the concentration of hydroxide ions is slightly less than the concentration of hydronium ions then the solution should have a pH of approximately 6. **T** or **F**

First, hydronium ion concentration is the same as proton concentration, and if hydroxides are slightly less than proton concentration then the solution is slightly acidic or a little less than 7; thus, TRUE.

15) When a basic solution is neutralized the pH goes down to 7.0. **T** or **F**

If a solution is basic then pH is greater than 7, if the solution is neutralized then the pH drops to 7; thus, TRUE.

16) An Arrhenius acid is considered to be the proton donor and an Arrhenius base is the proton acceptor. **T** or **F**

This is the definition of a Bronsted-Lowry acid and base; thus, FALSE. An Arrhenius acid is a proton donor and an Arrhenius base is a hydroxide donor.

17) By definition a weak acid is only partially dissociated due to the fact that the conjugate base is low in energy and reactive. **T** or **F**

A weak acid is partially dissociated due to the fact that the conjugate base is HIGH in energy and reactive; thus, FALSE.

18) The reason a strong acid is only partially dissociated is because the conjugate base is low in energy and unreactive.

All strong acids are 100% dissociated; thus, FALSE.

19) When a solution that contains Cyanide (CN^{-1}) reacts with liquid water the conjugate acid has the molecular formula HCN. **T** or **F**

When cyanide is dissolved in water it will take a proton from water and reform the weak acid HCN; thus, TRUE.

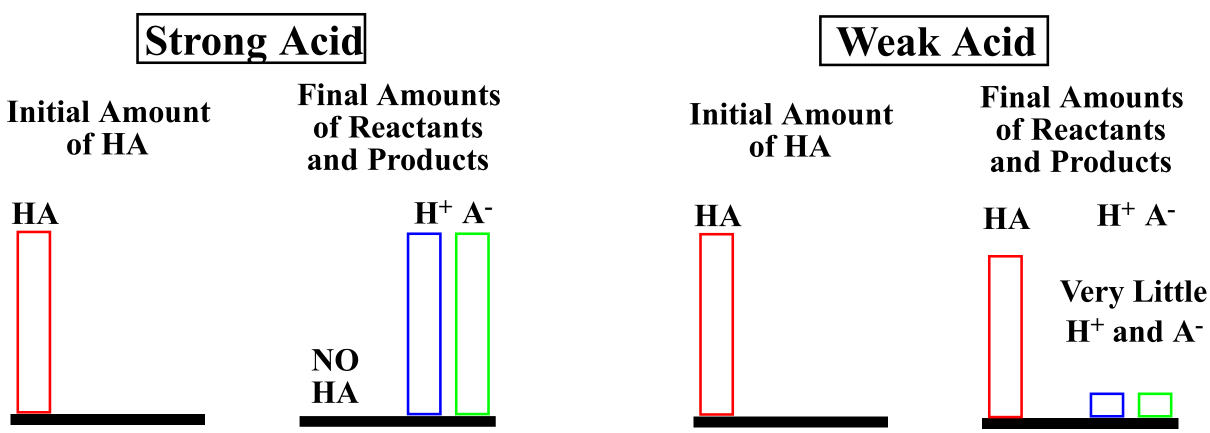
20) Define a strong and weak acid.

"Strong" and "weak" acid refer to the extent of dissociation of the acid as it relates to the stability of the conjugate base as summarized below:

Strong Acid (Completely Ionized) Produces	WeakAcid (Not Completely Ionized) Produces
⇩	⇩
Stable Conjugate Base	**Unstable Conjugate Base**
⇩	⇩
Conjugate Base Low in E	**Conjugate Base High in E**
⇩	⇩
Conjugate Base is Unreactive	**Conjugate Base is Reactive**
⇩	⇩
"Weak" Conjugate Base	**"Strong" Conjugate Base**

Another way of contrasting a "strong" and "weak" acid (HA) is by comparing amounts of H_3O^+ (same as saying H^+) and amount of A^- (conjugate base) formed after dissociation, which is visualized in the following diagram:

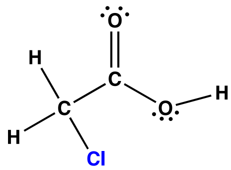
$$\underset{\text{Acid}}{HA_{(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Conjugate Acid}}{H_3O^+_{(aq)}} + \underset{\text{Conjugate Base}}{A^-_{(aq)}}$$



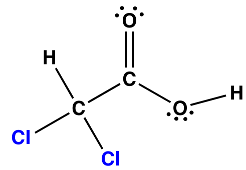
21) Define a strong and weak base.
"Strong" and "weak" base refer to the extent of dissociation of the base or as it relates to the stability of the conjugate acid formed; i.e., if conjugate acid is stable then equilibrium will lie far to the right which also increases the hydroxide concentration (strong base). If conjugate acid is unstable then it will want to take steps toward stability (back to reactants); thus, small increase in hydroxide ion concentration (weak base).

22) Which is the stronger acid? Why?

Chloroacetic Acid **(CH_2ClCO_2H)**



Dichloroacetic Acid **($CHCl_2CO_2H$)**



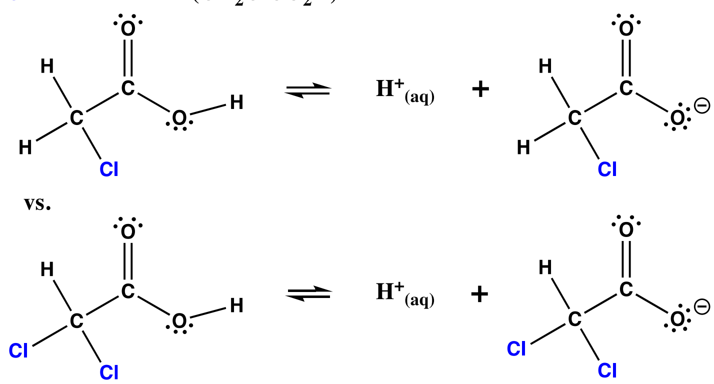
The stronger acid will have the more stable conjugate base.

Both carboxylic acids will have the following equilibriums:

$$CH_2ClCO_2H_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_2ClCO_2^-{}_{(aq)} \qquad CHCl_2CO_2H_{(aq)} \rightleftharpoons H^+_{(aq)} + CHCl_2CO_2^-{}_{(aq)}$$

To better visualize the stability of the conjugate bases let's draw the two-dimensional Lewis structures for the conjugate bases for the equilibrium:

Chloroacetic Acid **(CH_2ClCO_2H)**



Dichloroacetic Acid **($CHCl_2CO_2H$)**

While both conjugate bases have stabilization due to a resonance contributor shown below, the dichloroacetic acid has twice the electron withdrawing inductive effect. Recall, chlorine has an electronegativity value of approximately 3 (although, some have the value as high as 3.16). Thus, the conjugate base of dichloroacetic acid is more stable and therefore dichloroacetic acid will be the more dissociated, stronger acid.