Key points about acids and bases so far

- Pay attention to charge on the acid or base
- There is always an acid and base on the reactant side and their conjugate acid or base on the product side
- An increase in $K_a$ causes a decrease in $K_b$
- A decrease in $K_a$ causes an increase in $K_b$
- $K_a$ = equilibrium constant, water + weak acid
  When you have two different weak acids the larger the $K_a$ the stronger the acid

Two clicker questions to practice what we went over in the previous lecture

A) $\text{HPO}_4^{2-}$ base on reactant side
B) $\text{HCO}_3^-$ acid on reactant side (has an $\text{H}^+$)
C) $\text{CO}_3^{2-}$ base
D) $\text{H}_2\text{PO}_4^-$ acid (has an $\text{H}^+$ and one more $\text{H}^+$ than $\text{HPO}_4^{2-}$)

**Ka Reaction**

$\text{HA} \text{(aq)} + \text{H}_2\text{O} \text{(l)} \rightleftharpoons \text{A}^- \text{(aq)} + \text{H}_3\text{O}^+ \text{(aq)}$

$\text{HA}$-Weak Acid

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

**Weak Acid**

$\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{HPO}_4^{2-} (aq) + \text{H}_3\text{O}^+ (aq)$

$\text{H}_2\text{PO}_4^-$ (aq)- acid loses $\text{H}^+$
$\text{H}_3\text{O}^+$ (aq)- base accepts $\text{H}^+$

**Kb Reaction**

$\text{B} (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{HB}^+ (aq) + \text{OH}^- (aq)$

$\text{H}_2\text{O} (l)$- acid loses and $\text{H}^+$ to $\text{OH}^-$ (aq)
$\text{B} (aq)$- base , $\text{HB}^+$ gained the the $\text{H}^+$

$K_b = [\text{HB}^+] [\text{OH}^-] / [\text{B}]$

Could $\text{H}_2\text{PO}_4^-$ be a weak base? Yes

$\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{OH}^- (aq) + \text{H}_3\text{PO}_4$

$\text{H}_2\text{PO}_4^-$ = base , $\text{H}_3\text{PO}_4$= accepts the $\text{H}^+$
$\text{H}_2\text{O} (l)$ = acid in this case because it loses an $\text{H}^+$ and become $\text{OH}^-$ on the product side

Could $\text{CO}_3^{2-}$ be an acid? No

$\text{CO}_3^{2-} + \text{H}_2\text{O} \rightarrow ?$

-could not be an acid because there is no $\text{H}^+$ for the carbonate to give away
**Could there be a Kb reaction for CO$_3^{2-}$? Yes**

CO$_3^{2-}$ + H$_2$O $\rightleftharpoons$ HCO$_3^-$ + OH$^-$

CO$_3^{2-}$ – base
HCO$_3^-$ accepts an H$^+$ from CO$_3^{2-}$
H$_2$O – acid loses an H$^+$ to OH$^-$

**Base:**
HPO$_4^{2-}$ $K_a = 3.6 \times 10^{-13}$ would be for its weak acid reaction, but we don’t need $K_a$ for it for this problem

**Acids:**
HCO$_3^-$ $K_a = 5.61 \times 10^{-11}$
H$_2$PO$_4^-$ $K_a = 6.23 \times 10^{-8}$

The question was, does the equilibrium for the following rxn lie to the right or to the left:
HPO$_4^{2-}$ (aq) + HCO$_3^-$ (l) $\rightleftharpoons$ CO$_3^{2-}$ (aq) + H$_2$PO$_4^-$ (aq)

B) Equilibrium lies to left
- Need to look at the acid values to find which equilibrium constant for the acid is larger, and that’s the stronger acid so it wins
  - $K_a = 6.23 \times 10^{-8}$ is larger than $K_a = 5.61 \times 10^{-11}$, so H$_2$PO$_4^-$ is the stronger acid

**What is the pH scale?**
- an increase in H$^+$ causes a decrease in pH
- a decrease in H$^+$ causes an increase in pH
- the more acidic - the smaller the pH more H$^+$

**What does pH scale mean mathematically?**
Scale A - linear
Scale C - logarithmic
pH doesn’t always have to be an integer

**What you need to be good at?**
Molarity = mols solute/L solution (imp. during titrations)

**Adding strong acid to strong base or 1 L of water**
HCl = strong acid so it completely dissociates in water, into 1mol H$^+$ ions plus 1 mol Cl$^-$ ions
H₂SO₄ = the first H⁺ is strong, but the second one (on HSO₄⁻) is weak acid doesn’t
dissociate very much, so when you have 1 mol of H₂SO₄, you get a little more than 1 mol
H⁺, because you get 1 mol of H⁺ from the first dissociation that is complete, and a tiny
bit more H⁺ from the dissociation of HSO₄⁻ which is a weak acid so only a tiny bit of it
dissociates. So, if you prepare a solution of 1 mol of H₂SO₄ in water, there will be lots of
solutes in the solution:
  • there will be no H₂SO₄
  • there will be a little more than 1 mol of H⁺
  • there will be a little less than 1 mol of HSO₄⁻
  • there will be a tiny bit (much less than 1 mol) of SO₄²⁻

NaOH = strong base, dissociates into 1 mol OH⁻ ions and 1 mol Na⁺ ions

Example of strong base calculations

a) NaOH- strong base (arrhenius base)
   NaOH- when added to water completely dissociates
   1 mol NaOH 1 mol OH⁻
   .0012 M NaOH solution:
   
   \[
   [\text{OH}^{-}] = 0.0012 \text{ M}
   \]
   
   \[
   \text{pOH} = -\log(0.0012) = 2.92
   \]
   
   \[
   \text{pH} = 14 - 2.92 = 11.08
   \]
   
   11.08: when take the log of something you get the significant digits after the
decimal place (2 sig figs)

b) Sr(OH)₂ is a strong base
   pH= 10.46
   Goal: to find the concentration of Sr(OH)₂
   find the concentration of OH⁻, get that from pOH, get pOH from pH
   
   1) \[ \text{pOH} = 14 - \text{pH} \]
   \[
   \text{pOH} = 14 - 10.46
   \]
   \[
   \text{pH} = 3.54
   \]
   
   2) \[ [\text{OH}^{-}] = 10^{-\text{pOH}} \]
   \[
   = 10^{-3.57}
   \]
   \[
   = .000288 \text{ M (keep more sig figs than needed, round back at the end)}
   \]
   
   3) \[ 0.000288 \text{ mol OH}^{-} / \text{ L} \times 1 \text{ mol Sr(OH)₂} / 2 \text{ mol OH}^{-} \]
   \[
   = .00014 \text{ M the Sr(OH)₂ concentration}
   \]

How much does a weak acid or base dissociate?
Weak acids do not completely dissociate (they mostly stay whole and do not fully
break apart into H⁺ and the conjugate base), need to write an equilibrium table