

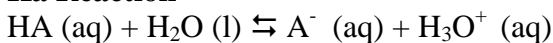
Key points about acids and bases so far

- 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) HPO_4^{2-} base on reactant side
B) HCO_3^- acid on reactant side (has an H^+)
C) CO_3^{2-} base
D) H_2PO_4^- acid (has an H^+ and one more H^+ than HPO_4^{2-})

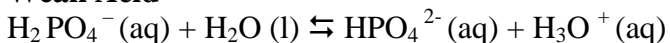
K_a Reaction



HA-Weak Acid

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

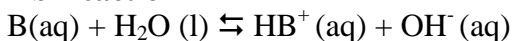
Weak Acid



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

$\text{H}_3\text{O}^+ \text{ (aq)}$ - base accepts H^+

K_b Reaction

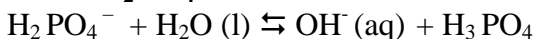


$\text{H}_2\text{O (l)}$ - acid loses and H^+ to $\text{OH}^- \text{ (aq)}$

B (aq)- base , HB^+ gained the the H^+

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

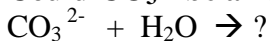
Could H_2PO_4^- be a weak base? Yes



H_2PO_4^- = base , H_3PO_4 = accepts the H^+

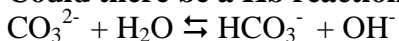
$\text{H}_2\text{O (l)}$ = acid in this case because it loses an H^+ and become OH^- on the product side

Could CO_3^{2-} be an acid? No



-could not be an acid because there is no H^+ for the carbonate to give away

Could there be a Kb reaction for CO_3^{2-} ? Yes



CO_3^{2-} – base

HCO_3^- - accepts an H^+ from CO_3^{2-}

H_2O – acid loses an H^+ to OH^-

Base:

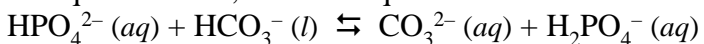
HPO_4^{2-} $K_a = 3.6 \times 10^{-13}$ would be for its weak acid reaction, but we don't need K_a for it for this problem

Acids:

HCO_3^- $K_a = 5.61 \times 10^{-11}$

H_2PO_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:



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_2PO_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 1 mol H^+ ions plus 1 mol Cl^- ions

H_2SO_4 = the first H^+ is strong, but the second one (on HSO_4^-) is weak acid doesn't dissociate very much, so when you have 1 mol of H_2SO_4 , 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 HSO_4^- which is a weak acid so only a tiny bit of it dissociates. So, if you prepare a solution of 1 mol of H_2SO_4 in water, there will be lots of solutes in the solution:

- there will be no H_2SO_4
- there will be a little more than 1 mol of H^+
- there will be a little less than 1 mol of HSO_4^-
- there will be a tiny bit (much less than 1 mol) of SO_4^{2-}

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:

$$[\text{OH}^-] = .0012 \text{ M}$$

$$\text{pOH} = -\log(.0012) = 2.92$$

$$\text{pH} = 14 - 2.92 = 11.08$$

11.08: when take the log of something you get the significant digits after the decimal place (2 sig figs)

b) $\text{Sr}(\text{OH})_2$ is a strong base

$$\text{pH} = 10.46$$

Goal: to find the concentration of $\text{Sr}(\text{OH})_2$

find the concentration of OH^- , get that from pOH , get pOH from pH

$$1) \text{pOH} = 14 - \text{pH}$$

$$\text{pOH} = 14 - 10.46$$

$$\text{pH} = 3.54$$

$$2) [\text{OH}^-] = 10^{-\text{pOH}}$$

$$= 10^{-3.57}$$

$$= .000288 \text{ M (keep more sig figs than needed, round back at the end)}$$

$$3) .000288 \text{ mol OH}^- / 1\text{L} \times 1 \text{ mol Sr}(\text{OH})_2 / 2 \text{ mol OH}^-$$

$$= .00014 \text{ M the Sr}(\text{OH})_2 \text{ concentration}$$

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