

Chemistry 30

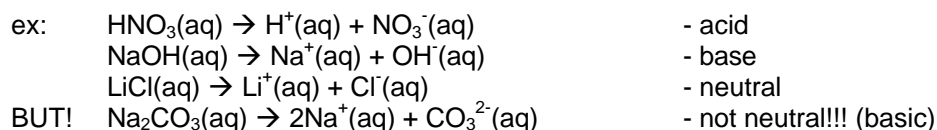
Acids & Bases

Properties of Acids (& Bases)

- dissolve in $\text{H}_2\text{O}(\text{l})$
- blue litmus \rightarrow red (or red litmus \rightarrow blue)
- taste sour (taste bitter)
- react with Mg, Zn to form $\text{H}_2(\text{g})$
- neutralize bases (neutralize acids)
- conduct

Arrhenius Definition of Acids and Bases

- all acids produce hydrogen ions ($\text{H}^+(\text{aq})$) when dissolved in water
- all bases produce hydroxide ions ($\text{OH}^-(\text{aq})$) when dissolved in water
- the properties of acids and bases are due to the $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ produced
- all neutral substances that are neutral are neutral because they don't produce $\text{H}^+(\text{aq})$ when dissolved

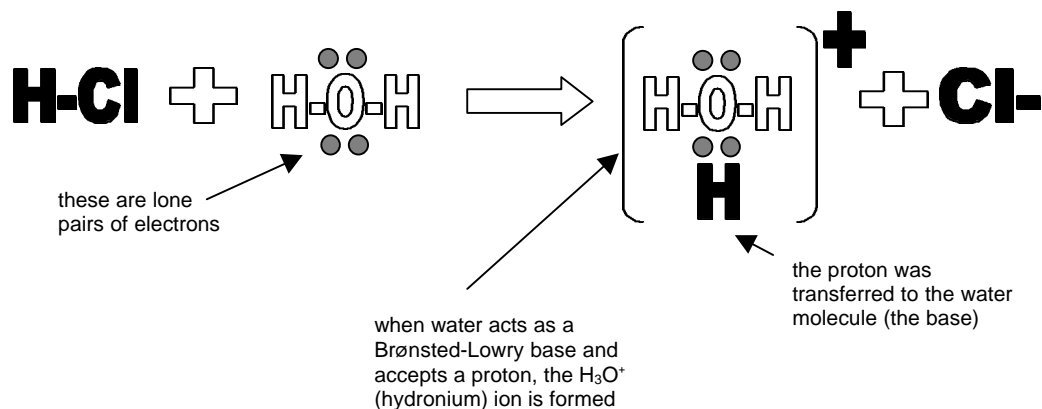


The Arrhenius definition isn't enough to explain all acidic & basic substances

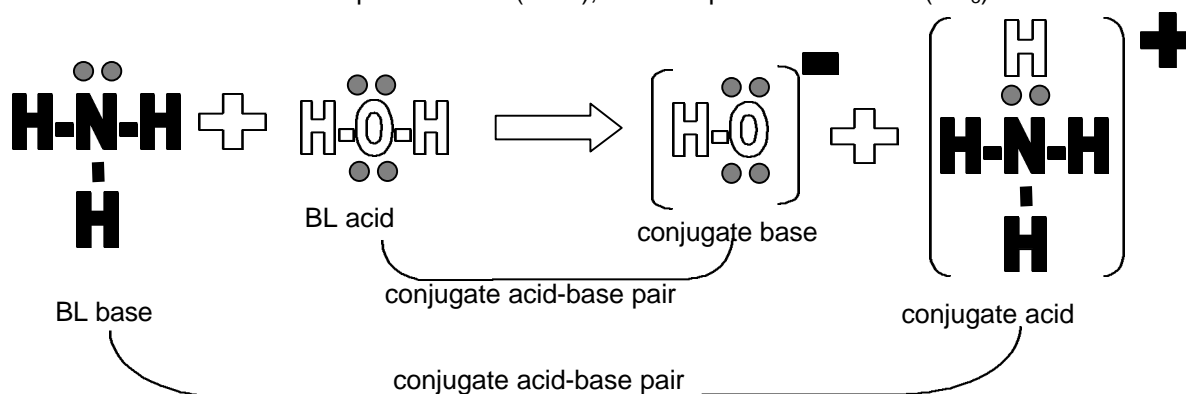
Brønsted-Lowry Acids and Bases

- an acid is a proton donor (i.e. H^+ donor)
- a base is a proton acceptor (i.e. H^+ acceptor) – must have an lone electron pair to accept the proton
- an acid can only behave like an acid if there's a base present to accept the H^+ ion (when HCl is dissolved in water, it behaves like an acid, but in a nonpolar solvent, HCl is NOT an acid)
- water can behave as both a proton donor (acid) and a proton acceptor (base) – the positive hydrogen ion is attracted to the negative lone pair, and “bonds” to the water molecule instead of the chloride ion, forming the $\text{H}_3\text{O}^+(\text{aq})$ (hydronium ion) – water is acting as a proton acceptor (base), and $\text{HCl}(\text{aq})$ is the proton donor (acid)

ex:



- water can also behave as a proton donor (base), for example with ammonia (NH₃)



- Brønsted-Lowry acids and bases always come in pairs – termed conjugate pairs
- when a BL acid (in this case H₂O) donates its proton, it becomes a BL base (OH⁻), the “conjugate base” of water. The pair, H₂O and OH⁻ are called a “conjugate acid-base pair”
- NH₃ behaves as a BL base (accepting a proton), its conjugate acid is NH₄⁺

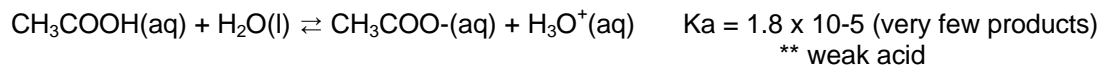
These reactions are really equilibrium reactions:



- when a base accepts a proton → acid (conjugate acid)
- when an acid donates a proton → base (conjugate base)
- conjugate acid-base pair – differ only by ONE H⁺ ion
- conjugate base always has a charge –1 compared to its conjugate acid
- ↑ strength of the acid, ↓ strength of the conjugate base

Weak and Strong Acids

- Just like many other reactions, not all acids and bases will completely dissociate – but reach equilibrium instead → K_a (just a special case of K_{eq} for acids, K_b for bases)



- the stronger the acid, the more likely it is to give up a proton (same with bases, the stronger the base, the more likely it is to accept a proton)
- strength has NOTHING to do with the concentration of the acid (or base)

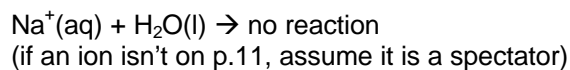
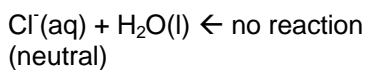
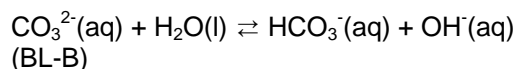
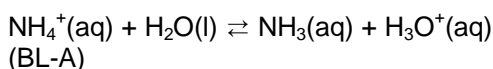
p. 11 in the data booklet -strongest acids are at the top left of the table (↑↑ K_a – 100% ionization)
 -strongest bases are at the bottom left of the table (↓↓ K_a – means favours the reverse reaction)

- There are only 6 “strong” acids; the “strong” bases are any soluble hydroxides (salts with OH⁻ ions that dissolve in water)

- strong acids and bases ionize 100% (no equilibrium) (ex: $\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$)
- weak acids and bases ionize < 100% (equilibrium) (ex: $\text{HF(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{F}^-(\text{aq})$)
- conjugates of the strong acids and bases are SO weak, they will not ionize at all (this is true for all of the 6 strong acids and the soluble hydroxides)
(ex: $\text{Cl}^-(\text{aq}) + \text{H}_2\text{O(l)} \leftarrow \text{HCl(aq)} + \text{H}_2\text{O(l)}$) so no reaction— this reaction will not take place in the forward direction)
- the stronger an acid, the weaker its conjugate base (and vice versa)

Acidity/Basicity of Salts (Salt Hydrolysis)

- a salt is an ionic compound
- whether a salt is acidic, basic or neutral depends on whether the ions that form the salt behave as BL acids, BL bases or neither
- each ion acts independently



- $\text{NH}_4\text{Cl(aq)}$ in water, the $\text{NH}_4^+(\text{aq})$ behaves as a BL acid, the $\text{Cl}^-(\text{aq})$ as a BL neutral, so overall, the salt will be an acid
- NaCl(aq) in water, the $\text{Na}^+(\text{aq})$ is neutral, the $\text{Cl}^-(\text{aq})$ is neutral, so the salt is neutral
- $(\text{NH}_4)_2\text{CO}_3(\text{aq})$ in water, the $\text{NH}_4^+(\text{aq})$ is an acid, the $\text{CO}_3^{2-}(\text{aq})$ is a base, so the salt will be CLOSE (not precisely) to neutral

Writing Brønsted-Lowry reactions between acids and bases

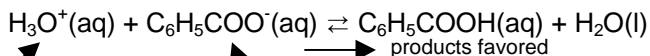
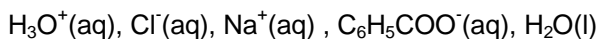
- So far, we've looked only at how acids, bases and salts behave in water
- Brønsted-Lowry theory allows for acid base reactions both with water and in the absence of water
- using Brønsted-Lowry theory, we can write net ionic equations for the reaction taking place and indicate the position of the equilibrium

ex: A solution of KOH(aq) is reacted with a solution of $\text{NaHCO}_3(\text{aq})$

Steps

1. list all of the ions produced (for weak acids or insoluble compounds – write them in their PREDOMINANT form – the undissociated form) → $\text{Na}^+(\text{aq}), \text{HCO}_3^-(\text{aq}), \text{K}^+(\text{aq}), \text{OH}^-(\text{aq}), \text{H}_2\text{O(l)}$
2. identify all ions as either BL-Acids (A), BL-Bases (B) or spectators (**)(if the ion is not on p.11 – it is a spectator) → ** A or B SA ** B SB A or B
 - $\text{H}_2\text{O(l)}$ and $\text{HCO}_3^-(\text{aq})$ appear on both sides of p.11's table (they can behave as both acid or base – called amphiprotic or amphoteric – they must be considered as both)
3. identify the strongest acid (SA) – closest to the top of p.11's table on the left side, and strongest base (SB) – closest to the bottom of the table on the right side – the SA and SB will be the first to react with each other – they are most likely to donate and accept a proton
4. Write the NET equation – just the reaction between the SA and SB (leave out all of the spectator ions) – transfer a proton from the SA to the SB → $\text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{H}_2\text{O(l)} + \text{CO}_3^{2-}(\text{aq})$
5. Predict the equilibrium status (which, forward or reverse predominates – look at the two acids, $\text{HCO}_3^-(\text{aq})$ and $\text{H}_2\text{O(l)}$ (on the product side) – whichever acid is stronger (higher on p.11) pushes the equilibrium away by donating a proton → $\text{HCO}_3^-(\text{aq})$ is a stronger acid than $\text{H}_2\text{O(l)}$, so the products are favoured

ex: Predict the reaction between HCl(aq) and NaC₆H₅COO(aq)

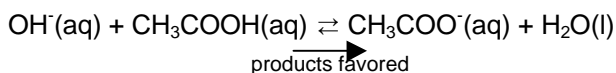
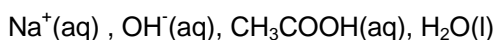


SA – so proton donor

SB – so proton acceptor

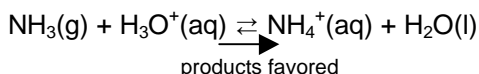
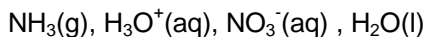
split HCl(aq) into H₃O⁺(aq) and Cl⁻(aq) because HCl(aq) is a strong acid (100% ionization). Split NaC₆H₅COO(aq) because it is soluble (100% dissociation)

ex: NaOH(aq) in CH₃COOH(aq)



CH₃COOH(aq) is NOT split up because it is a weak acid – K_a is very small, so most of the CH₃COOH(aq) is found undissociated, so it is written as is

ex: NH₃(aq) in HNO₃(aq)

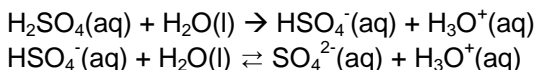


Polyprotic/Polybasic Acids and Bases

- acids (or bases) that can lose (or gain) more than one proton per molecule
- dibasic/diprotic (two protons): H₂CO₃(aq), H₂SO₄(aq), H₂SO₃(aq)
- tribasic/triprotic (three protons): H₃PO₄(aq), H₃BO₃(aq)
- the first proton is the “easiest” to gain (or lose) (K_a or K_b is the largest)
- others are harder because of the charge (a polyprotic acid that has lost a proton will already have a negative charge and will not lose another proton as easily because of the charge)

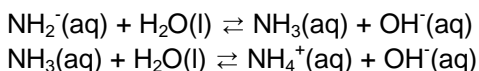
** not all molecules that have more than one hydrogen are polyprotic (ex: organic acids)

ex: Write the ionization steps for H₂SO₄(aq)



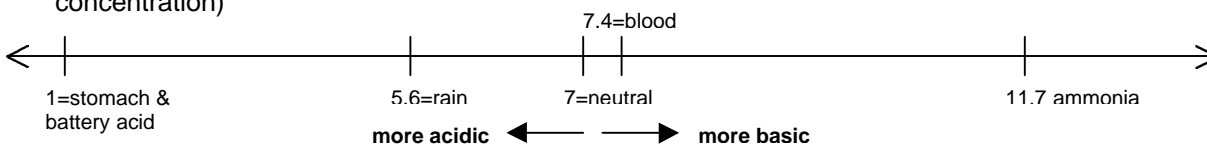
(this step is one way because H₂SO₄ is strong)
(this is equilibrium because HSO₄⁻ is not strong)

ex: NH₂⁻(aq) is dibasic. Write its ionization steps



pH – The scale, what it means, and calculations

- pH = the “power” (exponent) of hydrogen (1909)
- developed to easily compare acidity and basicity
- an difference in pH of one is a difference in concentration of 10 (difference of 3 = 1000x concentration)



$$\text{pH} = -\log [\text{H}_3\text{O}^+(\text{aq})]$$

$$\text{pOH} = -\log [\text{OH}^-(\text{aq})]$$

$$10^{-\text{pH}} = [\text{H}_3\text{O}^+(\text{aq})]$$

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

pH and pOH are related through the ionization constant of water (more on this later), by:

$$\text{pH} + \text{pOH} = 14.00$$

$$[\text{H}_3\text{O}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.00 \times 10^{-14}$$

ex: What is the pH of a solution of HCl(aq) if its concentration is 3.00×10^{-2} mol/L?

Since HCl(aq) is a strong acid: $\text{HCl}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ (this is 100% reaction)

therefore $[\text{H}_3\text{O}^+(\text{aq})] = [\text{HCl}(\text{aq})] = 3.00 \times 10^{-2}$ mol/L

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+(\text{aq})] \\ &= -\log(3.00 \times 10^{-2}) \\ &= 1.523 \end{aligned}$$

SD rules:

in pH only the numbers after the decimal count
(numbers before come from the exponent- if you remember your log math so don't count)

ex: What is the pH of 0.00453 mol/L NaOH(aq)?

Since NaOH(aq) is soluble: $\text{NaOH}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq})$ (this is 100% reaction)

therefore $[\text{OH}^-(\text{aq})] = 0.00453$ mol/L

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-(\text{aq})] \\ &= -\log(0.00453) \\ &= 2.344 \\ \text{pH} &= 14.00 - 2.344 \\ &= 11.656 \text{ (keep 3 SD)} \end{aligned}$$

ex: Calculate the pH if 4.75 g of NaOH(s) is dissolved in water to make 750 mL of solution

Find [NaOH] first, then [OH⁻(aq)], then from this, pOH and finally pH

$$\frac{4.75\text{g}}{40.00\text{g}} \times \frac{1\text{mol}}{40.00\text{g}} \times \frac{1}{0.750\text{L}} = 0.158\text{mol/L NaOH}$$

Since [NaOH(aq)] is soluble, $[\text{NaOH}(\text{aq})] = [\text{OH}^-(\text{aq})] = 0.158$ mol/L

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-(\text{aq})] \\ &= -\log(0.158) \\ &= 0.800 \text{ (if you rounded in the middle, it would be 0.801)} \\ \text{pH} &= 14.00 - 0.800 = 13.200 \end{aligned}$$

ex: Calculate the mass of HCl(g) required to make 500 mL of a solution with a pOH of 12.00

$$\begin{aligned} \text{pOH} &= 12.00 \\ \text{pH} &= 14.00 - 12.00 = 2.00 \end{aligned}$$

This one is exactly the opposite- from pOH, find pH, then find $[\text{H}_3\text{O}^+(\text{aq})]$, then $[\text{HCl}]$, and finally mass

$$\begin{aligned} [\text{H}_3\text{O}^+(\text{aq})] &= 10^{-\text{pH}} \\ &= 10^{-(2.00)} \\ &= 1.00 \times 10^{-2} \text{ mol/L} \end{aligned}$$

$$\frac{1.0 \times 10^{-2} \text{ mol}}{\text{L}} \times \frac{0.500 \text{ L}}{1} \times \frac{36.46 \text{ g}}{1 \text{ mol}} = 0.18 \text{ g}$$

ex: Calculate the $[\text{H}_3\text{O}^+(\text{aq})]$ in a solution made from 2.50 g $\text{Ba}(\text{OH})_2$ dissolved to make 450 mL of solution at 25°C.

Find $[\text{OH}^-]$, then use $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}$ to determine $[\text{H}_3\text{O}^+]$ (don't forget 2 OH^- per mol $\text{Ba}(\text{OH})_2$)

$$\frac{2.50 \text{ g Ba}(\text{OH})_2}{171.35 \text{ g}} \times \frac{1 \text{ mol}}{1 \text{ mol Ba}(\text{OH})_2} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba}(\text{OH})_2} \times \frac{1}{0.450 \text{ L}} = 0.0648 \text{ mol / L OH}^-$$

$$1.00 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{[\text{OH}^-]}$$

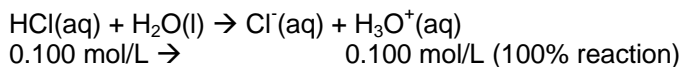
$$= \frac{1.00 \times 10^{-14}}{0.0648}$$

$$= 1.54 \times 10^{-13} \text{ mol/L}$$

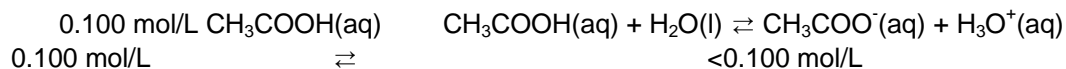
Calculations Involving Weak Acids and Weak Bases

- pH and pOH calculations can be performed for weak acids and bases as well, but must include provisions for the equilibrium that is set up because of the strength of the acid/base

ex: 0.100 mol/L HCl(aq)



$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+(\text{aq})] \\ &= 1.000 \end{aligned}$$



$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+(\text{aq})] \\ &= \text{much more than 1.000 since not all of the acid dissociates} \end{aligned}$$

** weak acids will have a higher pH than a strong acid at the same concentration

** \uparrow acid strength $\rightarrow \uparrow$ Ka $\rightarrow \downarrow$ pH

Calculations involving weak acids (or bases) include $[H_3O^+(aq)]$, $[OH^-(aq)]$, K_a (or K_b), pH and pOH

ex: In a 0.100 mol/L solution of propanoic acid, the hydronium ion concentration is found to be 1.16×10^{-3} mol/L. Calculate the ionization constant (K_a , or K_{eq}) for propanoic acid.

Solve this the way you would solve a K_{eq} problem – you need [eq] – use an ICE chart!
 - since they gave us $[H_3O^+(aq)]$, we can fill in [I], and [E], then work backwards to fill in the rest of the [E]s

This should remind you of equilibrium – there is nothing new, just now it's always acids and bases

	$C_2H_5COOH(aq) + H_2O(l) \rightleftharpoons C_2H_5COO^-(aq) + H_3O^+(aq)$
[I]	0.100 mol/L 0 0
[C]	$-(1.16 \times 10^{-3})$ $+(1.16 \times 10^{-3})$ $+(1.16 \times 10^{-3})$
[E]	0.099 mol/L 1.16×10^{-3} mol/L 1.16×10^{-3} mol/L

Start here, then fill backwards

$$K_a = \frac{[C_2H_5COO^-][H_3O^+]}{[C_2H_5COOH]}$$

$$= \frac{(1.16 \times 10^{-3})(1.16 \times 10^{-3})}{(0.099)}$$

$$= 1.36 \times 10^{-5}$$

make sure to keep all of your digits even though you record just the ones that are significant

ex: What is the pH of a 0.100 mol/L solution of HCN(aq)?

to solve for pH, you need to solve for $[H_3O^+(aq)]$. This is just like the equilibrium unit, solve an ICE chart for x, this will give you $[H_3O^+(aq)]$, then you can use $pH = -\log [H_3O^+]$

	$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$
[I]	0.100 mol/L 0 0
[C]	-x +x +x
[E]	$(0.100 - x)$ mol/L x mol/L x mol/L

$$K_a = \frac{[CN^-][H_3O^+]}{[HCN]}$$

to solve this without the quadratic equation, the approximation must be used:

- $1000xK_a < 0.100$?
- $1000(6.2 \times 10^{-10}) < 0.100$
- therefore $x \approx 0$
- therefore $0.100 - x \approx 0.100$

$$6.2 \times 10^{-10} = \frac{(x)(x)}{(0.100 - x)}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.100}$$

$$6.2 \times 10^{-11} = x^2$$

$$7.9 \times 10^{-6} = x$$

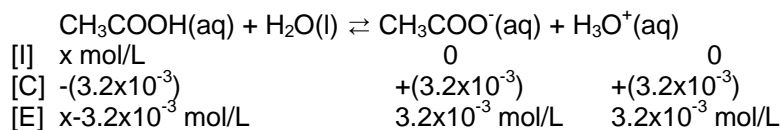
$$x = [H_3O^+(aq)] = 7.9 \times 10^{-6} \text{ mol/L}$$

$$\therefore pH = -\log [H_3O^+(aq)]$$

$$= -\log (7.9 \times 10^{-6})$$

$$= 5.10$$

ex: Determine the initial concentration of an acetic acid solution if its pH is 2.50.



this is another solve for x problem, but the only x will be with the initial concentration – from pH get $[\text{H}_3\text{O}^+(\text{aq})]$, and fill in the ICE chart

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.8 \times 10^{-5} = \frac{(3.2 \times 10^{-3})(3.2 \times 10^{-3})}{(x - 3.2 \times 10^{-3})}$$

Start here, then fill backwards

$$\begin{aligned} [\text{H}_3\text{O}^+(\text{aq})] &= 10^{-\text{pH}} \\ &= 10^{-2.50} \\ &= 3.2 \times 10^{-3} \text{ mol/L} \end{aligned}$$

$$1.8 \times 10^{-5}x - 5.7 \times 10^{-8} = 1.0 \times 10^{-5}$$

$$1.85 \times 10^{-5}x = 1.0 \times 10^{-5}$$

$$x = 0.56$$

$$\therefore [\text{CH}_3\text{COOH}(\text{aq})] = 0.56 \text{ mol/L}$$

Just another look at $[\text{H}_3\text{O}^+(\text{aq})]$, $[\text{OH}^-(\text{aq})]$

Don't you think it's rather strange that when an acid dissolves in water and donates its proton, it creates $\text{OH}^-(\text{aq})$?

(in case you've forgotten): $0.100 \text{ mol/L HCl}(\text{aq}) \rightarrow 0.100 \text{ mol/L H}_3\text{O}^+(\text{aq})$

and since $[\text{H}_3\text{O}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.00 \times 10^{-14}$

then $[\text{OH}^-(\text{aq})] = 1.00 \times 10^{-14} / 0.100 = 1.00 \times 10^{-13} \text{ mol/L}$

Really, acids don't create $\text{OH}^-(\text{aq})$, they just affect the equilibrium that already exists in the water where they are dissolved :



$$\text{(at } 25^\circ\text{C, in pure water, } [\text{H}_3\text{O}^+(\text{aq})] = [\text{OH}^-(\text{aq})] = \sqrt{1.00 \times 10^{-14}} = 1.00 \times 10^{-7} \text{ mol/L)}$$

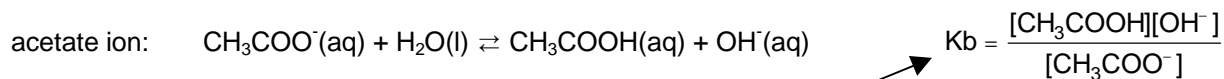
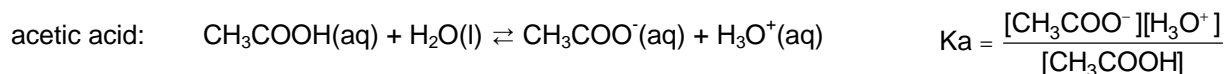
According to LeChatelier, if $\uparrow \text{H}_3\text{O}^+(\text{aq})$ (by adding an acid), then the equilibrium will shift \leftarrow , reducing the $[\text{OH}^-(\text{aq})]$ already present to maintain K_{eq} (K_a).

Since K_a is a constant (1.00×10^{-14}) at 25°C , then if one of either $\text{H}_3\text{O}^+(\text{aq})$ or $\text{OH}^-(\text{aq})$ is increased, then the other will decrease.

This equilibrium allows us to perform similar calculations for weak bases that we can for weak acids, even though our K_a table (on p.11) only has K_a values.

The Relationship between Ka and Kb

Looking at the weak acid, $\text{CH}_3\text{COOH}(\text{aq})$ and its conjugate base, $\text{CH}_3\text{COO}^-(\text{aq})$, we can examine the relationship between K_a and K_b .



we use K_b for bases instead of K_a

Although at first glance, the two don't seem related ($K_a \neq 1/K_b$), they are related through K_w (the equilibrium constant for water):

$$\frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} \times \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

So:

$$\boxed{K_a \times K_b = K_w}$$

don't forget this one

From this relationship, you can calculate K_b ($K_b = K_w/K_a$), where at 25°C , $K_w = 1.00 \times 10^{-14}$

ex: Calculate the K_b for $\text{NH}_3(\text{aq})$

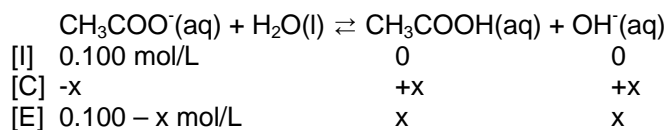
From p.11, K_a for $\text{NH}_4^+(\text{aq})$ is 5.8×10^{-10}

$$K_a \times K_b = K_w$$

$$\begin{aligned} K_b &= \frac{K_w}{K_a} \\ &= \frac{1.00 \times 10^{-14}}{5.8 \times 10^{-10}} \\ &= 1.7 \times 10^{-5} \end{aligned}$$

Once K_b is determined, all of the calculations performed for weak acids can be done for weak bases too.

ex: Calculate the pH of a 0.100 mol/L solution of $\text{NaCH}_3\text{COO}(\text{aq})$.



this is solved exactly the same way as if it were a weak acid, except that in the ICE chart, OH^- is present (instead of H_3O^+), we must find K_b (not just look it up), and when finding pH, we have to find pOH first.

$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.100 - x}$$

$$5.6 \times 10^{-10} = x^2$$

$$7.5 \times 10^{-6} = x$$

Approximation rule:
 $1000 \times K_b < 0.100$ (initial conc.)
 $\therefore x \approx 0$
 $\therefore 0.100 - x \approx 0.100$

$$x = [\text{OH}^-(\text{aq})] = 7.5 \times 10^{-6} \text{ mol/L}$$

$$\begin{aligned} \therefore \text{pOH} &= -\log [\text{OH}^-(\text{aq})] \\ &= -\log (7.5 \times 10^{-6}) \\ &= 5.13 \end{aligned}$$

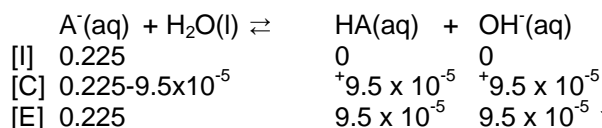
$$\text{pOH} + \text{pH} = 14.00$$

$$\therefore \text{pH} = 14.00 - 5.13 = 8.87$$

K_b was solved above using $K_a \times K_b = K_w$

ex: Calculate Kb for an unknown base if a 0.225 mol/L solution has a pH of 9.98.

let $A^-(aq)$ be the unknown base:



from pH, find pOH and then $OH^-(aq)$ so you can fill in part of the ICE chart & work backwards

$$pH + pOH = 14.00$$

$$pOH = 14.00 - 9.98 = 4.02$$

$$[OH^-(aq)] = 10^{-pOH}$$

$$[OH^-(aq)] = 10^{-4.02} = 9.5 \times 10^{-5} \text{ mol/L}$$

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

$$= \frac{(9.5 \times 10^{-5})^2}{(0.225)}$$

$$= 4.1 \times 10^{-8}$$

don't forget to keep all of the digits from the subtraction

Buffers

A buffer is a solution that will resist changes in pH when either an acid or a base is added. Buffers are used wherever a specific pH must be maintained – some examples are in the body (the blood buffers), in pools and hot tubs (to allow the chemicals to work properly inhibiting the growth of algae and other organisms) and fish tanks.

There are two basic types of buffers:

Solutions of strong concentrated acids (or strong concentrated bases)

- resist changes in pH when acids are added because they already have such a low pH
- resist changes in pH when (small amounts of) bases are added because they neutralize

Solutions that are able to neutralize both acids or bases

- there are two types of this variety of buffer – solutions of amphiprotic (amphoteric) substances (like $HCO_3^-(aq)$), or a solution of a weak acid and its conjugate base together ($CH_3COOH(aq)$ and $CH_3COO^-(aq)$ often in equal quantities).

These solutions behave as buffers because they are able to remove the excess ions that will lead to a change in pH, $H_3O^+(aq)$ or $OH^-(aq)$.

ex: In a buffer solution of $CH_3COOH(aq)$ and $CH_3COO^-(aq)$, $HCl(aq)$ was added. Why is the pH not changed noticeably?

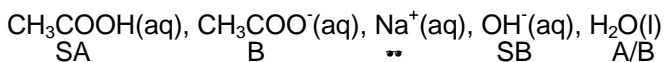
examine the situation using B-L equations: $CH_3COOH(aq)$, $CH_3COO^-(aq)$, $H_3O^+(aq)$, $Cl^-(aq)$, $H_2O(l)$

A
SB
SA
..
A/B



- since the forward reaction is favoured, most of the reactants (most of the $H_3O^+(aq)$) becomes products, and since most of the $H_3O^+(aq)$ is removed, the pH is not affected greatly.

ex: similarly, when NaOH(aq) is added:



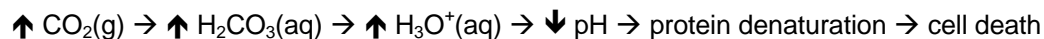
- since the forward reaction is favoured, most of the reactants (most of the OH⁻(aq)) becomes products, and since most of the OH⁻(aq) is removed, the pH is not affected greatly.

Buffers in the Body

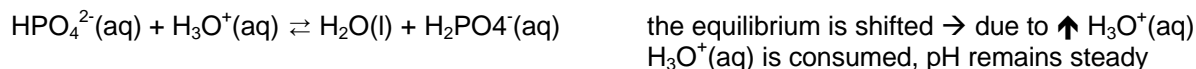
The body must maintain a pH of approximately 7.35 to keep enzymes and proteins from denaturing. If proteins and enzymes denature, they cannot function, and the body's cells (and eventually the body) die. The two buffers in the blood are HCO₃⁻(aq) and HPO₄²⁻(aq).

The acid that must be buffered is H₂CO₃(aq), which is produced naturally as CO₂(g) produced by the body's cells during cellular respiration is carried by the blood's plasma to be excreted in the lungs. The level of CO₂(g), and hence H₂CO₃(aq) is constantly changing as we take breaths and exhale. Our body's buffer system must be constantly adjusting to maintain a narrow pH range.

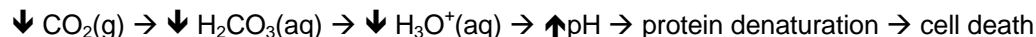
ex: as someone holds their breath (suffocates), what happens to maintain pH in the body?



the increase in H₃O⁺(aq) affects the following buffer system:



ex: if someone hyperventilates, what happens to maintain the pH in their body?

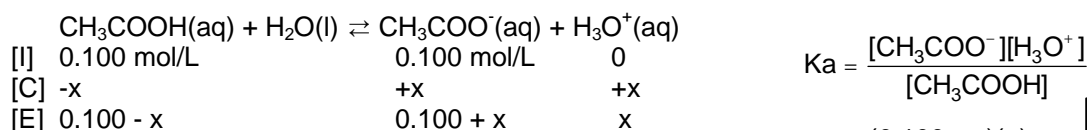


The decrease in H₃O⁺(aq) will cause the opposite shift in the equilibrium system described above.

Calculating the pH of a buffer system

Calculating the pH of a buffer is not unlike the other calculations. The only difference is that they are made with some acid and its conjugate base to start with (both product and reactant are present initially).

ex: (equal amounts of acid and conjugate base) Calculate the resulting pH of a buffer consisting of 0.100 mol/L each of CH₃COOH(aq) and its conjugate base CH₃COO⁻(aq).



$$x = [\text{H}_3\text{O}^+(\text{aq})] = 1.8 \times 10^{-5} \text{ mol/L}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}_3\text{O}^+(\text{aq})] \\ &= -\log (1.8 \times 10^{-5}) \\ &= 4.74 \end{aligned}$$

$$1.8 \times 10^{-5} = \frac{(0.100 - x)(x)}{(0.100 + x)}$$

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

Approx:- 1000K_a < 0.100
so x=0, 0.100-x=0.100,
and 0.100+x=0. Then,
the 0.100 on the top and
bottom cancel!!!

this only works if the initial concentrations are the same, if they are not, you have to complete the calculation, but still use the approximation