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Acid/Base Properties of Salts

Hiding in plain sight

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Recognizing Bases

- Sometimes it seems that all acids and bases are labeled with H^+ or OH^-
- Remember that a base is ANYTHING that can accept a proton
- The salt formed by the dissociation of an acid is the "conjugate base" of the acid. This must mean that the salt is a base, whether it has an OH^- or not.

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Pick a salt, any salt

- How about sodium acetate, $NaOAc$? An excellent choice.
- Sodium acetate is an ionic solid.
- Ionic solids dissociate in aqueous solution
- Aqueous $NaOAc$ will exist as anions and cations.

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NaOAc in aqueous solution

$\text{NaOAc}_{(aq)} \rightarrow \text{Na}^+_{(aq)} + \text{OAc}^-_{(aq)}$

So what can we say about Na^+ and OAc^- in aqueous solution?

We need to think "backwards". How did NaOAc get the Na^+ in the first place? (Or, more accurately, what is one way it could have gotten it?)

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Salts are products of acid/base reactions

You might recall that the reaction of an acid and a base yields a salt and water:

$\text{NaOH} + \text{HOAc} \rightarrow \text{H}_2\text{O} + \text{NaOAc}$

Keep in mind, this is the "net reaction", it doesn't give any detailed information on mechanisms.

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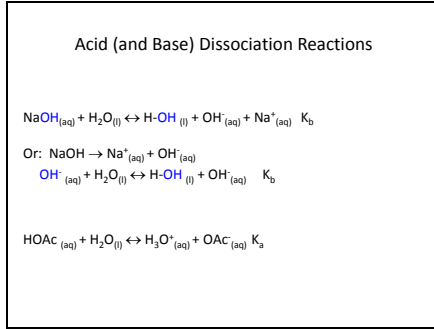
A more detailed reaction

Think of a titration:

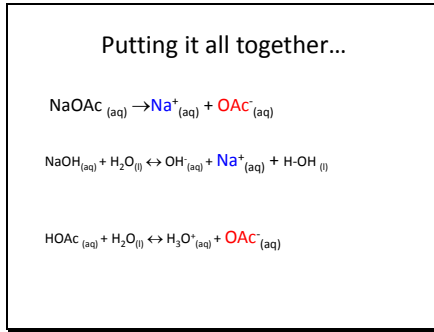
$\text{NaOH}_{(aq)} + \text{HOAc}_{(aq)} \leftrightarrow \text{H}_2\text{O}_{(l)} + \text{NaOAc}_{(aq)}$

But what are $\text{NaOH}_{(aq)}$ and $\text{HOAc}_{(aq)}$?

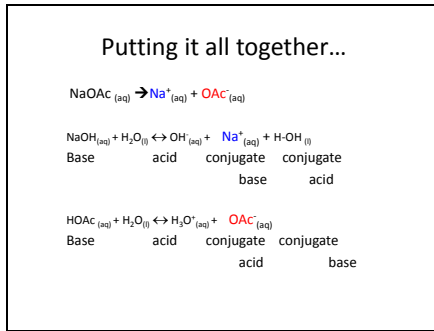
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In short...

We have the conjugate acid (Na^+) of a base (NaOH), and the conjugate base (OAc^-) of an acid (HOAc).

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You can always tell...

Take the anion and add an H^+ , that'll tell you what acid the anion "came from".

Take the cation and either add an OH^- or take away an H^+ and that will tell you what base the cation came from.

Ignore anything strong – it won't, it CAN'T go back!!!

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K_b becomes K_a as K_a becomes K_b

$\text{NaOAc}_{(aq)} \rightarrow \text{Na}^+_{(aq)} + \text{OAc}^-_{(aq)}$

$\text{Na}^+_{(aq)} + 2 \text{H}_2\text{O}_{(l)} \leftrightarrow \text{NaOH}_{(aq)} + \text{H}_3\text{O}^+_{(aq)}$
(or, if you prefer)

$\text{Na}^+_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{NaOH}_{(aq)} + \text{H}^+_{(aq)} \quad K_a = K_w / K_b$

$\text{OAc}^-_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{OH}^-_{(aq)} + \text{HOAc}_{(aq)} \quad K_b = K_w / K_a$

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But one is strong and the other weak.

$$\text{NaOH}_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{H-OH}_{(l)} + \text{OH}^-_{(aq)} + \text{Na}^+_{(aq)} \quad K_b = \infty$$
$$\text{Na}^+_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{NaOH}_{(aq)} + \text{H}^+_{(aq)} \quad K_a = K_w / K_b = 0$$
$$\text{HOAc}_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{H}_3\text{O}^+_{(aq)} + \text{OAc}^-_{(aq)} \quad K_a = 1.8 \times 10^{-5}$$
$$\text{OAc}^-_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{OH}^-_{(aq)} + \text{HOAc}_{(aq)} \quad K_b = K_w / K_a = 10^{-14} / 1.8 \times 10^{-5} = 5.56 \times 10^{-10}$$

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Net Result

NaOAc gives rise to a single equilibrium reaction that must be considered:

$$\text{OAc}^-_{(aq)} + \text{H}_2\text{O}_{(l)} \leftrightarrow \text{OH}^-_{(aq)} + \text{HOAc}_{(aq)}$$
$$K_b = 5.56 \times 10^{-10}$$

The salt is a base!!!!

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Sample problem

What is the pH of a 0.100 M solution of NaOAc?

What do we need?

Balanced equation

What is it?

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$\text{NaOAc (aq)} \rightarrow \text{Na}^+ \text{(aq)} + \text{OAc}^- \text{(aq)}$

Where does Na^+ "come from"?

NaOH

Where does the $\text{OAc}^- \text{(aq)}$ "come from"?

HOAc

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$\text{NaOAc (aq)} \rightarrow \text{Na}^+ \text{(aq)} + \text{OAc}^- \text{(aq)}$

~~NaOH~~
Strong Base

HOAc
Weak Acid

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$\text{OAc}^- \text{(aq)} + \text{H}_2\text{O (l)} \leftrightarrow \text{HOAc (aq)} + \text{OH}^- \text{(aq)}$

$\text{OAc}^- \text{(aq)} + \text{H}_2\text{O (l)} \leftrightarrow \text{HOAc (aq)} + \text{OH}^- \text{(aq)}$

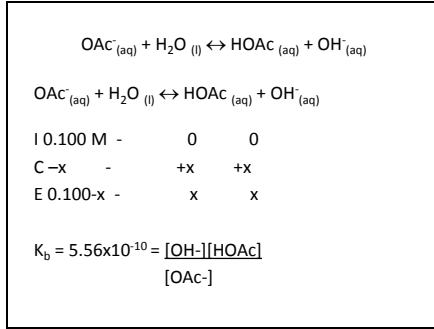
I	0.100 M	-	0	0
C	-x	-	+x	+x
E	0.100-x	-	x	x

Now all we need is K!

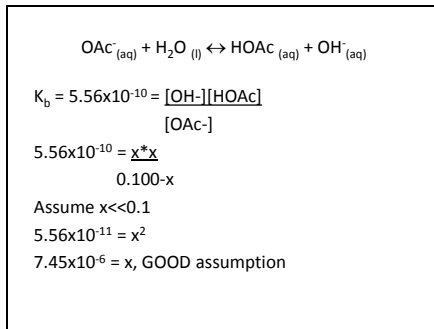
$K_b(\text{OAc}^-) = K_w/K_a(\text{HOAc})$

$= 1 \times 10^{-14} / 1.8 \times 10^{-5} = 5.56 \times 10^{-10}$

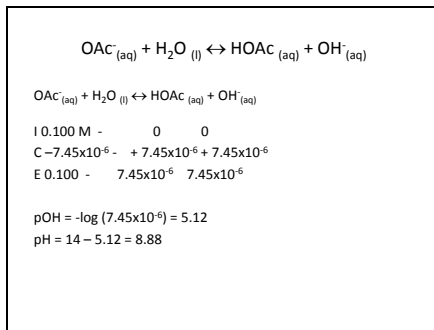
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More fun with salts...

What can we say about:
NaCl?
 $\text{NaCl} = \text{Na}^+ + \text{Cl}^-$
It came from...
Cation + OH⁻ = NaOH
Anion + H⁺ = HCl
So....

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If you can't find K_a ...

...maybe it's a strong acid.
Or if you do find it, it is HUGE!

Strong acids:
HCl $K_a \sim 10^6$
HBr
HI
HNO₃
HClO₄
H₂SO₄ (diprotic – more on this next week)

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If you can't find a K_b ...

...maybe it's a strong base. (Or K_b is huge.)

List of strong bases:
LiOH
NaOH $K_b \sim 10^8$
KOH
Sr(OH)₂
Ca(OH)₂
Ba(OH)₂

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What kind of acid/base is it?

NaOH?
STRONG base

HCl?
STRONG acid

The salt is...
NEUTRAL we "ignore both"

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More fun with salts...

What can we say about:
KCl

Came from:
KOH and HCl

Also neutral!

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More fun with salts...

What can we say about:
 $\text{Ca}(\text{OAc})_2$

It came from:
 $\text{Ca}(\text{OH})_2$ and HOAc

Only the HOAc matters, it's weak.

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More fun with salts...

What can we say about:
NH₄Cl

It came from?
NH₃ (or NH₄OH) and HCl

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More fun with salts...

What can we say about:
NH₄OAc

It came from:
NH₃ and HOAc

BOTH WEAK!!!

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If they are both weak...

Look at the K. Biggest K wins.

Although for NH₄OAc..

$K_b(\text{NH}_3) = K_{a2} = \frac{1.00 \times 10^{-14}}{1.76 \times 10^{-5}} = 5.68 \times 10^{-10}$
 $K_b(\text{NH}_3) = 1.76 \times 10^{-5}$

$K_b(\text{HOAc}) = K_{a2} = \frac{1.00 \times 10^{-14}}{1.79 \times 10^{-5}} = 5.59 \times 10^{-10}$
 $K_b(\text{HOAc}) = 1.79 \times 10^{-5}$

NH₄OAc ends up being neutral!

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If they don't cancel...???

We won't worry about the exact calculation (you'd need to do both simultaneous equilibria), but you can tell whether it is acidic or basic:

NH_4IO_3 (ammonium iodate)

It comes from:
 NH_3 and HIO_3

$K_b(\text{NH}_3) = 1.76 \times 10^{-5}$
 $K_a(\text{HIO}_3) = 1.7 \times 10^{-1}$

HIO_3 is a better acid than NH_3 is a base. Which means that NH_4^+ is a better acid than IO_3^- is a base.

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If they don't cancel...???

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A solution of NH_4IO_3 would be acidic.

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Another little problem

What is the pH of a 0.250 M solution of $\text{Ca}(\text{IO}_3)_2$?

Where do we start?

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Split the salt into ions

$\text{Ca}(\text{IO}_3)_2 \rightarrow \text{Ca}^{2+} + 2 \text{IO}_3^-$

Where did the ions come from?

$\text{Ca}(\text{OH})_2$ and HIO_3

Which are...?

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$\text{Ca}(\text{OH})_2$ is strong

HIO_3 is weak.

Ignore the Ca^{2+}

Now, we need to...

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Write a balanced equation!

$\text{IO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{HIO}_3(\text{aq}) + \text{OH}^-(\text{aq})$

Once I have a balanced equation:
ICE chart
K equation

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ICE ICE BABY ICE ICE

$\text{IO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{HIO}_3(\text{aq}) + \text{OH}^-(\text{aq})$

I		-		
C	-x	-	+x	+x
E		-		

What else?

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Another little problem

What is the pH of a 0.250 M solution of $\text{Ca}(\text{IO}_3)_2$?

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ICE ICE BABY ICE ICE

$\text{IO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{HIO}_3(\text{aq}) + \text{OH}^-(\text{aq})$

I	0.500 M	-	0	0
C	-x	-	+x	+x
E	0.500-x	-	x	x

Now K

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$$K = \frac{[\text{HIO}_3][\text{OH}^-]}{[\text{IO}_3^-]}$$

This is the K_b of IO_3^-

$$K_b(\text{IO}_3^-) = K_w/K_a$$
$$K_a(\text{HIO}_3) = 1.7 \times 10^{-1}$$
$$K_b = 1 \times 10^{-14} / 1.7 \times 10^{-1} = 5.88 \times 10^{-14}$$

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$$5.88 \times 10^{-14} = \frac{[\text{HIO}_3][\text{OH}^-]}{[\text{IO}_3^-]}$$
$$5.88 \times 10^{-14} = \frac{[x][x]}{[0.5-x]}$$

How do we solve this?

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Assume $x \ll 0.5$

$$5.88 \times 10^{-14} = \frac{[x][x]}{[0.5-x]} \approx \frac{[x][x]}{0.5}$$
$$2.941 \times 10^{-14} = x^2$$
$$1.715 \times 10^{-7} = x$$

Good assumption?
YES, MA'AM!

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ICE ICE BABY ICE ICE

$\text{IO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{HIO}_3(\text{aq}) + \text{OH}^-(\text{aq})$

I	0.500 M	-	0	0
C	-1.715×10^{-7}	-	$+1.715 \times 10^{-7}$	$+1.715 \times 10^{-7}$
E	0.500	-	1.715×10^{-7}	1.715×10^{-7}

How do we finish?

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Another little problem

What is the pH of a 0.250 M solution of $\text{Ca}(\text{IO}_3)_2$?

$[\text{OH}^-] = 1.715 \times 10^{-7}$
 $\text{pOH} = -\log(1.715 \times 10^{-7}) = 6.77$
 $\text{pH} = 14 - \text{pOH} = 14 - 6.77 = 7.23$

Make sense?
You bet! Strongish acid (HIO_3), very weak conjugate. Almost neutral.
