**Weak Acids and Bases**Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

*AP Chemistry Lecture Outline*

Video

1603

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**Weak Acids**

-- most acids are weak

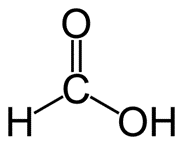
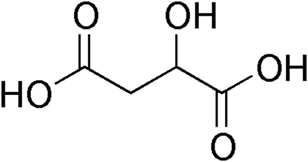
-- For a weak acid HX... HX(aq) H+(aq) + X–(aq)

-- acid-dissociation constant Ka =

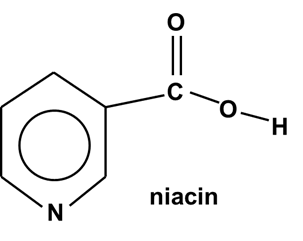
large Ka:

small Ka:

The % of a weak acid that is ionized is given by the equation:

For organic acids (containing only C, H, and O),

the “donated” H was connected to...



EX. A 0.020 M niacin solution has pH 3.26. (a) What % of the acid is ionized?

(b) What is Ka?

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EX. If Ka for niacin is 1.6 x 10–5, find the pH of a 0.010 M niacin solution.

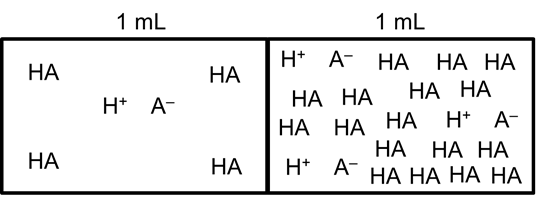
The conc. of a weak acid affects its % ionization.

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As [acid] conc. , % ion. \_\_\_\_; as [acid] conc. , % ion. \_\_\_\_.



-- more \_\_\_\_\_\_\_\_\_\_\_ -- more \_\_\_\_\_\_\_\_\_\_\_

-- \_\_\_\_ ionized = \_\_\_\_ -- \_\_\_\_ ionized = \_\_\_\_

(i.e., \_\_\_\_\_\_\_ % of (i.e., \_\_\_\_\_\_\_ % of

H+s popped off) H+s popped off)

-- has \_\_\_\_\_\_ [H+], -- has \_\_\_\_\_\_ [H+],

(i.e., \_\_\_\_\_ per \_\_\_\_\_) (i.e., \_\_\_\_\_ per \_\_\_\_\_)

and thus... and thus...

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EX. Calculate the % of HF molecules ionized in

– and the pH of – a 0.10 M HF solution. (Ka = 6.8 x 10–4)

EX. Calculate the % of HF molecules ionized in – and the pH of – a 0.010 M HF solution.

(Ka = 6.8 x 10–4)

Polyprotic acids – like sulfurous acid, H2SO3 – have more than one ionizable H.

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H2SO3(aq) H+(aq) + HSO3–(aq) (Ka1 = 1.7 x 10–2)

HSO3–(aq) H+(aq) + SO32–(aq) (Ka2 = 6.4 x 10–8)

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-- Usually, Ka2 is at least 1000X smaller than Ka1. In such cases, one can calculate [H+]

and pH based only on Ka1 (i.e., ignore Ka2 and pretend you have a monoprotic acid).

EX. Find the pH of a 0.0037 M carbonic acid solution. (Ka1 = 4.3 x 10–7, Ka2 = 5.6 x 10–11)

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**Weak Bases**

weak base + H2O conjugate acid + OH– Kb =

Weak bases are often nitrogen-containing

molecules (“amines”) or anions.

**..**

H

H

H

+

Example: H–N–CH3 (aq) + H2O(l) H–N–CH3 (aq) + OH–(aq)

EX. What is the equilibrium concentration of ammonia in a solution of pH 9.35? Ammonia’s

Kb = 1.8 x 10–5.

**The Ka / Kb Relationship for Conjugates**

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(1) NH3

(2) NH4+

Ka = Kb =

Then Ka x Kb =

For an acid and its conj. base:

Because they are easily calculated from Ka Just like pH = –log [H+],

values, reference tables are often “short” on Kb values.

EX. For hydrofluoric acid, Ka = 6.8 x 10–4.

a. Write the formula of the conjugate base.

b. Write the equation for which Kb applies.

c. Write the equation for which Ka applies.

d. Find pKa, pKb, and Kb.

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(6:30)

Ions related to weak acids or bases can themselves exhibit acidic or basic properties.

which is related The ion would ...and would

One such to \_\_\_\_, which tend to thus exhibit

ion is... is a ‘weak’. \_\_\_\_\_ H+... \_\_\_\_\_ properties.

CO32–

HCO3–

ClO–

ClO2–

NH4+

CH3COO–

\*\* NOTE: The ions we are dealing with here MUST be related to weaks, NOT strongs. Thus,

ions such as Cl– or NO3– or K+ or Sr2+ do NOT exhibit acidic or basic properties.

Because weak-related ions transfer H+s (as mentioned above, strong-related

ions DO NOT), solutions of salts containing ‘weak’ ions can be acidic or basic.

For example, consider adding a spoonful of the salt sodium hypochlorite to water.

The salt immediately does this...

The strong-related ion has no effect on H+ transfer, but the

weak-related ion DOES, in this way...

We expect a solution of sodium hypochlorite to have a pH...

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EX. What mass of solid potassium fluoride is required to make 5.0 L of a pH 8.35

solution?

To predict the pH of a salt solution, look at the formula of the salt.

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For salts with ions related to a...

(1) ...strong base and a weak acid, e.g.,

we predict a pH...

(2) ...weak base and a strong acid, e.g.,

we predict a pH...

(3) ...strong base and a strong acid, e.g.,

we predict a pH...

(4) ...weak base and a weak acid

--

e.g.,

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EX. Would we predict an NH4CN soln to be acidic or basic?

What related info is found in the ref. tables?

But we have \_\_\_\_\_, not \_\_\_\_, and \_\_\_\_\_ would \_\_\_\_\_ a proton, so we need its \_\_\_.

And we have \_\_\_\_\_, not \_\_\_\_, and \_\_\_\_\_ would \_\_\_\_\_ a proton, so we need its \_\_\_.

Anions that still have ionizable protons (e.g., HSO3–) are amphoteric.

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Equation for Ka:

Equation for Kb:

EX. Does KHSO3 form an acidic or a basic solution in water?

From a reference table, we find... (a) H2SO3 H+ + HSO3–

(b) HSO3– H+ + SO32–

*General Factors Affecting Acid Strength*

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1. To transfer H+, acid must have the polarity...

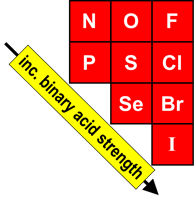
-- ionic hydrides don’t

-- C–H bonds don’t

As polarity increases, acid strength *generally* increases. BUT...

2. ...if H–X bond is too strong.. (e.g., )

3. Stable (i.e., weak) conjugate bases (e.g., ) indicate a...

For binary acids:

(1) In a group, bond strength is the determining factor.

As we go down a group... H2S

...charge separation \_\_\_,

so bond strength \_\_\_

and acid strength \_\_\_. H2Se

(2) In a period, bond polarity governs.

As we go L to R across a period... H2O HF

...bond polarity \_\_\_ and acid strength \_\_\_.

oxyacids:

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(8:24)

Consider H2SO4 (an acid) and Ca(OH)2 (a base).

O

H–O–S–O–H H–O–Ca–O–H

O

The EN between S and O is The EN between Ca and O is

\_\_\_\_\_\_\_\_\_\_ than that between O \_\_\_\_\_\_\_\_\_\_ than that between O

and H. This means that... and H. This means that...

When comparing related **Anything that draws e– density AWAY from**

oxyacids, remember this... **the “donate-able” H makes for a stronger acid.**

Related-Oxyacids *For oxyacids with the same # of O atoms, acid strength*

Comparison #1: *increases with increasing electronegativity of the central atom.*

e.g., HClO

HBrO H2SeO3

HIO H2SO3

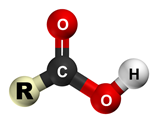
Related-Oxyacids *For oxyacids with the same central atom, acid strength increases*

Comparison #2: *as the # of oxygens attached to the central atom increases.*

e.g., HBrO HIO3

HBrO2 HIO2

Related-Oxyacids This comparison deals with carboxylic acids, which

 Comparison #3:contain the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

-- these are the largest category of organic acids

-- acid strength increases from adding add’l high-EN atoms

e.g., CH3COOH

CHF2COOH

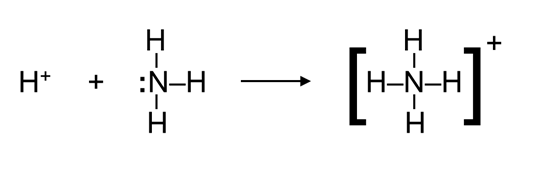
CF3COOH

**Lewis Acids and Lewis Bases**

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The Lewis definitions greatly broaden the range of acids because many

species other than H-containing ones can accept an e– pair.

Another example:

The simple term “acid” suggests that we are referring to an Arrhenius or a

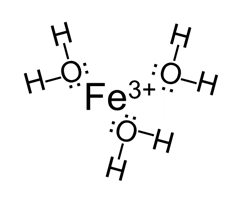
Bronsted-Lowry acid, i.e., an H-containing substance in an aqueous solution.

If you are referring to a Lewis acid, then use the term “Lewis acid.”

Substances with an incomplete octet (e.g., BF3) or ones having

vacant orbitals (e.g., Fe3+) can function as Lewis acids.

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-- The (+) charge attracts (i.e., accepts) the lone pairs of e–

on the O of a water molecule. This process is hydration:

the “glomming” of H2O m’cules onto metal ions.

Cation size and cation charge determine the extent to which the pH is affected.

Fe3+

K+

-- short distance -- larger distance

-- strong interaction -- weaker interaction

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