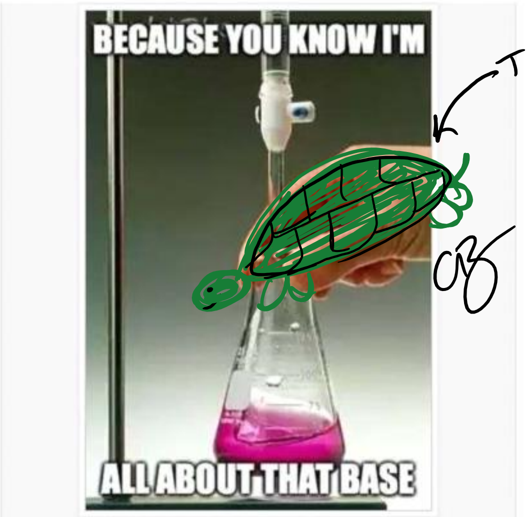
Chemistry 12

Unit V



Titration Turtle

Acid/Base II

Name:____________ Block:_____

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Chemistry 12 Acid/Base II ~ Helpful Equations

1. Strong acid in water: 100% dissociation

$HNO_3 + H_2O \Rightarrow H_3O^+ + NO_3^-$
.10M .10M .10M pH = -log(.10) = 1.00

Diprotic acids H_2SO_4 in water:

First proton is strong: $H_2SO_4 + H_2O \Rightarrow H_3O^+ + HSO_4^-$ 100%
.10M .10M .10M

Second proton is weak: $HSO_4^- + H_2O \Leftrightarrow H_3O^+ + SO_4^{2-}$ <5%
.10M <.10M <0.10M

* 2. Weak acid in water: not 100% dissociation ~<5%

$H_3PO_4 + H_2O \Leftrightarrow H_3O^+ + H_2PO_4^-$
.10M 0.0279M 0.0279M
pH = -log(0.0279) = 1.56

3. Strong base (hydroxide base) in water: 100% dissociation (water not in reaction; just a dissociation) eg. OH⁻ bases, NH_2^-, O^{-2}

$Sr(OH)_2 \rightarrow Sr^{2+} + 2OH^-$ pOH = -log(.20) = 0.70
.10M .10M .20M pH = 13.30

* 4. Weak base in water: not 100% dissociation

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
.10M 1.34×10^{-3}M 1.34×10^{-3}M
pOH = -log(1.34×10^{-3}) = 2.87 pH = 11.13

5. ANY reaction that involves a STRONG acid or base goes 100% to completion. So a weak acid with a strong base is 100% due to the strong base. A strong acid with a weak base is 100% due to the strong acid. ionizing 100%.

strong a/b will continuously drive the FWD rxn of the weak a/b

Example: $H_3PO_4 + 3NaOH \Rightarrow 3H_2O + Na_3PO_4$
weak strong

added continuously.

triprotic

The OH^- ions take all three protons off of each H_3PO_4 molecule, such as in a titration. If H_3PO_4 was merely in water, only one proton would come off at less than 100% like #2 earlier. (but will still react to completion 100%)

6. Weak acid and weak base:

$\underbrace{\underset{a}{NH_4^+} + \underset{b}{SO_4^{2-}}}_{\text{Reactants favored.}} \Leftrightarrow \underset{b}{NH_3} + \underset{a}{HSO_4^-}$ (side with weak acid is favoured)

RVS rxn driven

$HSO_4^- > NH_4^+$
stronger ∴ give up more H^+

I) Weak Acid Equilibrium and K_a

How do strong acids behave in water?

- dissociate 100% to create H_3O^+
- one-way rxn →

Write an equation for, and give the $[H_3O^+]$ and pH for a 0.0010M solution of HCl.

$$\underset{0.0010M}{HCl} + H_2O \longrightarrow \underset{0.0010M}{[H_3O^+]} + Cl^-$$

$pH = -\log(0.0010M)$

$\boxed{pH = 3.0}$ strong ā

How do weak acids behave in water?

- dissociate into ions much less than 100% ~5%.
- equilibrium rxns ⇌

Write an equation for a 0.010M HF solution.

$$\underset{0.010M}{HF} + H_2O \rightleftharpoons [H_3O^+] + F^-$$

? ... much less <<< 0.010M

It is not as simple to find the pH of a 0.010M solution of a weak acid as you must first have information on the extent of dissociation for the weak acid in question. K_a, the weak acid equilibrium constant, helps to determine this information and provides a means to solve **weak acid** problems. K_a is a K_{eq} $= \frac{[P]}{[R]}$ for weak acids, so it is simply an equilibrium constant, and therefore all K_{eq} rules apply.

(aq) ✓ ~~(s)~~ (g) ✓ ~~(l)~~

H^+ F^- H^+

$$HF_{(aq)} + \cancel{H_2O_{(l)}} \Leftrightarrow \underbrace{H_3O^+_{(aq)} + F^-_{(aq)}}_{[conc.]\ are\ equal} \quad K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{[H_3O^+]^2}{[HF]} = K_a$$

dissociating @ 1:1 ratio

weak acid

Notice the K_a values ___decrease___ as you go down the table because the acids are getting progressively ___weaker___ (creating ___less___ H_3O^+ in solution).

↑K_a = stronger weak acid.

In Chemistry 12, we work on quanititative problems that involve weak acids and weak bases.

$$\text{strong acid} = K_a = \frac{[H_3O^+]^2}{\emptyset}$$

diss. 100% ∴ no reactants = "undefined"

Weak acid quanititative problems can be broken into three types.

Type 1 Problems: Finding pH of a weak acid solution

Example: For the reaction:

$$H_2S_{(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+_{(aq)} + HS^-_{(aq)}$$

Calculate the $[H_3O^+]$ and pH if the $[H_2S]_i = 0.050M$.

R	$H_2S + H_2O \rightleftharpoons$	H_3O^+	$+ HS^-$
I	0.050M	0M	0M
C	$-x$	$+x$	$+x$
E	$0.050M - x$	x	

products always start @ ∅

*let $x = \Delta[H_3O^+]$

Look up Ka in table

$$Ka = \frac{[H_3O^+]^2_{eq}}{[H_2S]_{eq}} = \frac{x^2}{0.050M} = 9.1\times10^{-8}$$

s.f.

*assume that $0.050M - x \cong 0.050M$

$$\therefore x = \sqrt{(9.1\times10^{-8})(0.50M)}$$

$$x = 6.745\times10^{-5}$$

Check validity of assumption:

$$\frac{[H_3O^+]_{eq}}{[H_2S]_i} = \frac{6.7\times10^{-5}}{0.050} = 0.13\%$$

b/c it is less than 5% = VALID!

$\boxed{\therefore [H_3O^+] = 6.7\times10^{-5}M}$

$pH = -\log(6.745....)$

$\boxed{pH = 4.17}$

Example:

Calculate the pH of a 0.450M solution of $H_2PO_4^-$.

**Consider only the first proton dissociating from a polyprotic weak acid, as the dissociation of the second proton is negligible compared to the first.*

R	$H_2PO_4^- + H_2O \rightleftharpoons$	H_3O^+	$+ HPO_4^{2-}$
I	0.450M	0	0
C	$-x$	$+x$	$+x$
E	$0.450 - x \cong 0.450M$ *assume*	x	

Let $x = \Delta[H_3O^+]$

$$Ka = \frac{[H_3O^+]^2}{[H_2PO_4^-]} \qquad 6.2\times10^{-8} = \frac{x^2}{0.450} \qquad \therefore x = \sqrt{(6.2\times10^{-8})(0.450)}$$

$$x = 1.67\times10^{-4}$$

$\boxed{\therefore [H_3O^+] = 1.7\times10^{-4}M}$

$pH = -\log(1.67\times10^{-4})$

$\boxed{pH = 3.78}$

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→ metal(+) + non-metal(-) ⇒ ionic compound.

dissociate 100%.

Sometimes when salts dissolve in water, one of the ions can act as a weak acid in solution.

Example: What is the pH of a 0.100M NH_4Cl solution?

$$NH_4Cl_{(s)} \xrightarrow{H_2O} NH_4^+{}_{(aq)} + Cl^-{}_{(aq)}$$

weak acid

spectator ion ⇒ does nothing (aq)

? weak base

strongest weak acid

expect an acidic pH

K	$NH_4^+ + H_2O \rightleftharpoons$	NH_3	$+ H_3O^+$
I	0.100M	0	0
C	$-x$	$+x$	$+x$
			x

let $x = \Delta[H_3O^+]$

$$Ka = \frac{[H_3O^+]^2}{[NH_4^+]} \qquad 5.6\times10^{-10} = \frac{x^2}{0.100M}$$

$$x = \sqrt{(5.6\times10^{-10})(0.100M)}$$

$$x = 7.483\times10^{-6}$$

$x = \sqrt{(5.6\times10^{-10})(0.100M)}$

$x = 7.483\times10^{-6}$

$pH = -\log(7.483....)$

$\boxed{[H_3O^+] = 7.5\times10^{-6}M \quad pH = 5.13}$

C	$-x$		$+x$	$+x$
E	$0.100M - x \cong 0.100M$ *assumption*			x

let $x = \Delta[H_3O^+]$

Type 2 Problems: Calculating the Initial Concentration of a Weak Acid

Example: What $[Fe(H_2O)_6^{3+}]_i$ would be required to produce a pH of 4.120 ? → $[H_3O^+]_{eq}$

R	$Fe(H_2O)_6^{3+} + H_2O \rightleftharpoons$	H_3O^+	$+ Fe(H_2O)_5(OH)^{2+}$
I	x	(3) 0 M	0 M
C	-7.586×10^{-5}	$+7.586\times10^{-5}$	+
E	(4) $(x - 7.586\times10^{-5})$	$\boxed{7.586\times10^{-5}M}$	

(2) $[H_3O^+] = 10^{\wedge}(-4.120)$

$\boxed{= 7.586\times10^{-5}M}$

(1) Let $x = [Fe(H_2O)_6^{3+}]_i$

(5) $K_a = \dfrac{[H_3O^+]^2}{[Fe(H_2O)_6^{3+}]}$

$6.0\times10^{-3} = \dfrac{(7.586\times10^{-5})^2}{(x - 7.586\times10^{-5})} \times (\;)$

$6.0\times10^{-3}x - 4.552\times10^{-7} = 5.75\times10^{-9}$ $+4.552\times10^{-7}$

6.0×10^{-3}

$\boxed{x = 7.68\times10^{-5}M}$

Assignment 1:

1. Calculate the pH of a 0.50M solution of H_3BO_3.
2. Calculate the pH of a 0.235M solution of NaH_2PO_4.

Hebden p.128 #33 & p. 152 #79, 76, 78

Type 3 Problems: Finding the K_a of an Unknown Weak Acid

→ diprotic = "H_2A"

→ monoprotic = "HA" $H^+ A^-$

Example: A 0.20M solution ($[HA]_i$) of the weak acid, HA, has a pH of 1.32. ─▷ $[H_3O^+]_{eq}$.

Calculate the K_a of the weak acid and use this to identify it.

$K_a = \dfrac{[H_3O^+]^2}{[HA]_{eq}}$

R	$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons$	$H_3O^+_{(aq)}$	$+ A^-$
I	0.20M	(4) 0	0
ΔC	$-4.786\times10^{-2}M$	$+4.78\times10^{-2}M$	
E	0.152	$4.786\times10^{-2}M$	

(5) calc. $[HA]_{eq}$

(1) known:
$[HA]_i = 0.20M$
products = ∅ M @ start

(2) pH → $[H_3O^+]$

$[H_3O^+] = 10^{\wedge}(-1.32)$ ← 2 s.f. in [conc.]

(3) $[H_3O^+] = 4.7863\times10^{-2}M$

(6) $\therefore K_a = \dfrac{(4.786\times10^{-2})^2}{0.152}$

$K_a = 1.5\times10^{-2}$ ⟸ (7) Look up in K_a Table and match to an acid.

$\therefore$ unknown acid is H_2SO_3

Example: A 2.00M diprotic acid has a pH of 0.50. Calculate the K_a value.

Students try this one on own

R	$H_2A_{(aq)} + H_2O_{(l)} \rightleftharpoons$	H_3O^+	$+ HA^-$
I	2.00M	∅	∅

knowns:
2.00M

R	$H_2A_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+ + HA$		
I	2.00M	∅	∅
C	-0.316M	+0.3162M	
E	1.6838	0.3162M	

knowns:
2.00M
pH = 0.50

$pH = -\log[H_3O^+]$

$\therefore [H_3O^+] = 10^{\wedge}(-0.50)$

$= 0.3162M$

2sf.

$$Ka = \frac{[H_3O^+]^2}{[H_2A]} = \frac{(0.3162)^2}{(1.6838)}$$

$$\underbrace{Ka = 5.9 \times 10^{-2}}$$

unknown acid = $H_2C_2O_4$ (oxalic acid)

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