Calculating the pH of a weak acid solution

$HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$ $K_a = 7.1 \times 10^{-4}$

	HF	H^+	F^-	
I	0.50M	0	0	$[H^+]$ initial = 0.50M
C	$-X$	$+X$	$+X$	pH at equilibrium?
E	$0.5-X$	X	X	

$$K_a = \frac{[H^+][F^-]}{[HF]} = \frac{X \cdot X}{(0.5-X)} = 7.1 \times 10^{-4}$$

$$\frac{X^2}{0.5-X} = 7.1 \times 10^{-4}$$

Quadratic equation to solve for X

$X = 0.018487788$

$pH = -\log(0.018487788) = 1.73306808$

pH = 1.73 Round pH to 2 decimal places

$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

What is $[H_3O^+] =$ _____ M @ at 25°C?

$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

$\sqrt{x^2} = \sqrt{1 \times 10^{-14}}$

$x = 1.0 \times 10^{-7}$

$pH = -\log_{10}[H_3O^+]$

$pH = -\log(1.0 \times 10^{-7})$

$pH = 7.$

The pH concept

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

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

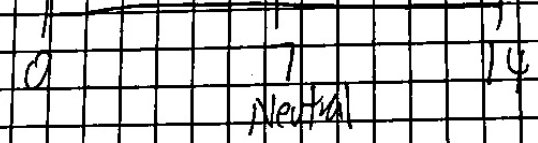
P Notation

$p(x) = \log(x)$

$pOH = -\log[OH^-]$

$pH + pOH = 14.00$

Why use pH?



$[H^+] = 1.0 \times 10^{-7} M$

☆ pH allow us to refer to very small concentrations using numbers from 0-14

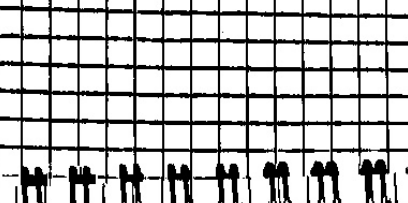
Water Auto ionization
self

$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-$

$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$

Pure liquid.



$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

	temp (°C)	K_w	
	0	1.1×10^{-15}	$K < 1$ Reactant is Favored
Body temp	25	1.0×10^{-14}	
	37	2.5×10^{-14}	K value increases

The acid-Base Properties of water

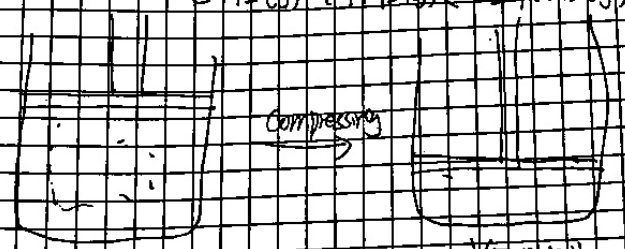
$[H^+] = [OH^-]$	$[H^+] > [OH^-]$	$[H^+] < [OH^-]$
Neutral pH = 7	Acidic pH	Basic pH
☆ Pure H_2O or Aqueous solution of a compound that is not an acid or a Base	AQ solution of an Acid $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$	Not Pure H_2O AQ. Solution of A Base $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-$

$NaCl_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

$\boxed{pH = 7}$

Changing the Volume in a Gaseous equilibrium

$3H_2(g) + 1N_2(g) \rightleftharpoons 2NH_3(g)$



V = 1.0 L
at
Equilibrium
P_{total} ↓
$PV = nRT$
V and T constant
P ∝ n.

V = 0.5 L

Q: (1) the equilibrium shifts to the left (Reactants) or the Right (Products)?

(Want to get Back to the Pressure)

to decrease # no. of Gas and thus decrease Pressure.

Chapter 16 - acid/Base equilibrium

Strong Acids

H_2SO_4, HCl, HBr, HI, HNO_3, $HClO_4$

100% dissociation (Not an equilibrium)

Strong bases

NaOH, LiOH, KOH, $Ca(OH)_2$

100% dissociation

Not an equilibrium.

Weak acids = all other acids

H_2O, CH_3OH, HF, CH_3CO_2H

intermediate dissociation

$HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$

Weak Bases (Non Hydroxide Bases)

NH_3, $CH_3CO_2^-$..

$NH_{3(aq)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$

Reaction Quotient, Q vs equilibrium Constant (K)

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

$$Q = \frac{[HI]^2}{[H_2][I_2]}$$ at any con. concentrations

Not necessary equilibrium

$$K = \frac{[HI]^2}{[H_2][I_2]}$$ at eq. con. c

$Q = K$ Only at equilibrium

$Q < K$ Reaction will shift Toward products

$Q > K$ Reaction ~ ~ ~ Reactants

$$K_c = \frac{(0.156)^2}{(0.022)(0.022)} = 49.37914884 = 49.4$$

$K_c = 4.94 \times 10^{-1}$

at equilibrium, then add some HI, Now [HI] = 0.200M

$$Q = \frac{(0.200)^2}{(0.022)(0.022)} = 126.8 = 127$$

What Does this mean? $Q > K$

equilibrium

Shifts to Reactant.

C. $2H_2(g) + 2I_2(g) \rightleftharpoons 4HI(g)$

* multiply rxn by 2
  square K

$K_c = \frac{[HI]^4}{[H_2]^2[I_2]^2} = X^2$

* multiply rxn by 3 ...
  Cube K.

le chatelier's Principle

If an outside influence upsets an equilibrium the system undergoes a change in a direction that counteracts the disturbing influence and, if possible, returns the system to equilibrium

$A + B = C \qquad K_c = \frac{[C]}{[A][B]}$

What are the outside influence?

1). add or remove a reactant or a product

* add a Reactant = equilibrium shifts towards the products
* add a product = eq. shifts to Reactants.

Remove a product = EQ. shift to product

2). Change the volume of the container in a Gaseous equilibrium. Decrease V, increase P ...

3). Change the temperature (K depends on T)

Homogeneous equilibria — All substances are in the same phase (same)

Heterogeneous equilibria — Substances are different (more than one phase)

$CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

$K_c = \frac{[CaO][CO_2]}{[CaCO_3]}$

$K_c = [CO_2]$ ← equilibrium concentration

$K_p = P_{CO_2}$ ← pressure

Acetic Acid

$CH_3CO_2H_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3CO_{2(aq)}^-$ (acetate)

$K_c = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H][H_2O]}$ ← This is not going to change significantly

☆ Pure liquids and pure solids are omitted from the equilibrium law

$K_c = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]}$

Manipulation of equilibrium laws.

- Reverse the Reaction.
- multiply the Reaction by a factor.

A. $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$ $K_c = \frac{[HI]^2}{[H_2][I_2]} = X$

B. $2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$ $K_c = \frac{[H_2][I_2]}{[HI]^2} = \frac{1}{X} = X'$

Equilibrium law

$K = \frac{[B]}{[A]} > 1$ $\quad$ $K = \frac{[B]}{[A]} < 1$

↑ Equilibrium constant

"product B favored"

☆ Can $K < 0$? (negative) No.

☆ No such thing has a negative concentration.

The Equilibrium Constant (K):

☆ A numberical value (No units)

found by experiement. What do we measure?

$N_2O_4(g) \rightleftharpoons 2NO_2(g)$ $\quad$ $K = 0.0465$

$K = 4.65 \times 10^{-2}$

$K = \frac{[NO_2]^2}{[N_2O_4]} = 4.65 \times 10^{-2}$

Equilibrium law.

The concentration of substances

The law of mass action

$aA + bB \rightleftharpoons cC + dD$

lowercase = Coefficients

Upper (A,B,C,D) = formulas

$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

↑ comes from experiments.

K_c vs K_p

K_c → concentration $[\ ]$ $\quad$ K_p → pressure partial pressure (atm)

$M = \frac{mol}{L}$

$K_c \neq K_p$

$K_p = K_c(0.0821T)^{\Delta n}$ ← change in # mol of Gas

Derived from $PV = nRT$.

$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$

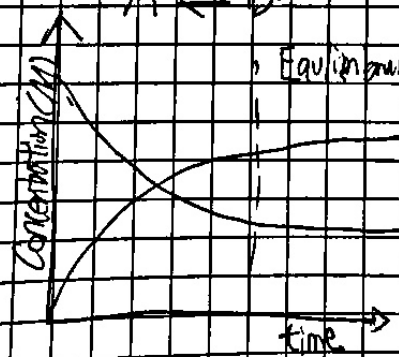
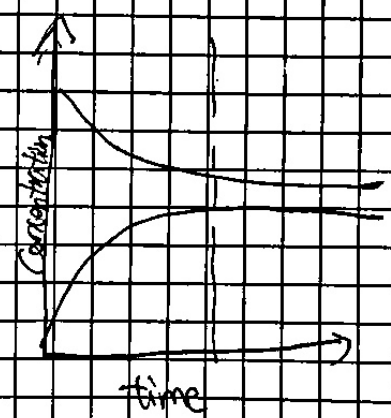
Triple Bond = 1σ + 2π

Chapter 15 – equilibrium

Physical (phase change)	Chemical (chemical equation)
$H_2O(l) \rightleftharpoons H_2O(g)$	$N_2O_4(g) \rightleftharpoons 2NO_2(g)$
$\rightleftharpoons$ or $\leftrightharpoons$	Reactant → Product
The process is happening in Both directions.	

Equilibrium

☆ forward rate = Reverse Rate.

☆ Change is still occuring, But the Reactant and Product Concentrations do not Change over time.

$A \rightleftharpoons B$