

# C12 - 0.0 - Methods 12 Rates/Eq/Solubility

LABS

Ripple Effects

Reaction Kinetics	Speed $aq^{\pm} > g \& l > s$	Reactant Nature (Chem) Amount/Concentration Energy/Energy Orientation/Geometry SA/V(Open?)/Size/pH Temp/Pressure(g) Catalyst (1stR/next*P)	Reaction Mechanism (Steps) <div style="border: 1px solid black; padding: 5px; margin: 5px;"> <math>2NO_2 \rightarrow NO + NO_3</math>  <math>CO + NO_3 \rightarrow NO_2 + CO_2</math>  <math>NO_2 + CO \rightarrow NO + CO_2</math> </div>	Reaction Intermediaries (Activated Complex) Overall Equation
Rates/Factors $\uparrow \downarrow \propto$ Takes Energy to break a bond Enthalpy -Exo/Endothermic -Activation Energy - $E_a$ - $\Delta H$ Le Chatelier's Principle - Nature Undoes*		Allotropes - Forms Spectators	Rate Determining Step - Slowest Step $rate = \frac{units}{time}$ Average (Secant) vs Instantaneous (Tangent)	
			You can't measure the concentration of a pure liquid or solid. Rates : Moles Reactants used up = Moles Products are made (Coefficients*).	

Equilibrium (Rates) : $K_f = K_r$	f - Forward, r - Reverse	$\frac{[Products]}{[Reactants]} = \frac{K_{fi}}{K_{ri}}$ $i - initial$	ICE Tables <div style="border: 1px solid black; padding: 5px; margin: 5px;"> <table> <tr> <td><math>\frac{mol}{L}</math></td><td><math>2@_{(g)}</math></td><td><math>+ @_{2(l)}</math></td><td><math>\rightleftharpoons 2@_{2(aq)}</math></td></tr> <tr> <td>Start</td><td>0</td><td>0</td><td>#</td></tr> <tr> <td><math>+ \Delta</math></td><td></td><td></td><td></td></tr> <tr> <td>= Eq</td><td>0.##</td><td></td><td></td></tr> </table> </div>	$\frac{mol}{L}$	$2@_{(g)}$	$+ @_{2(l)}$	$\rightleftharpoons 2@_{2(aq)}$	Start	0	0	#	$+ \Delta$				= Eq	0.##		
$\frac{mol}{L}$	$2@_{(g)}$	$+ @_{2(l)}$	$\rightleftharpoons 2@_{2(aq)}$																
Start	0	0	#																
$+ \Delta$																			
= Eq	0.##																		
Equilibrium Constant - $K_{eq}$	([ ]'s Constant at Eq)	$K_{eq} > 1$ - [products] > [reactants] $K_{eq} = 1$ - [products] = [reactants] $K_{eq} < 1$ - [products] < [reactants]	Mole Ratio - 2 : 1 : 2																
$K_{eq} = \frac{[C]^c \times [D]^d \times \dots}{[A]^a \times [B]^b \times \dots} \times \frac{[Products]}{[Reactants]}$																			
$aA + bB + \dots \rightleftharpoons cC + dD + \dots$	Pure solids/liquids are not included in $K_{eq}$																		

$Q = \text{Reaction Quotient (Trial K)}$	$Q = K_{eq} ; Eq$	$Q < K_{eq}$ Shift to Products	$Q > K_{eq}$ Shift to Reactants
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Moles Reactants used up = Moles Products are made

Only Temp can affect  $K_{eq}$ .

5% Rule

Solids and Pure Liquids exempt.

Assume  $x \ll \#$

$\downarrow$  Temp shift to side with heat (Larger hurdle to cross)

$\uparrow$  [Reactant]  $\rightarrow$   $\uparrow$  Forward Rate  
Pressure (g) constant at Eq  
Mass of (s) constant at Eq

Thermodynamics Laws

- 1) Energy cannot be created or destroyed.
- 2) For a spontaneous process, entropy of the universe increases.
- 3) A perfect crystal at zero Kelvin has zero entropy.

...

$S : \text{Entropy} \left( \frac{J}{K} \right)$   $+\Delta S$  favours products  
 $\Delta H$  favours Heat side

Spontaneous reaction occurs when

$\Delta S > 0$  &  $\Delta H < 0$

$\uparrow$  Moles  $\rightarrow$   $\uparrow$  Entropy

More Entropy Shifts ???

Solubility

Solubility Product Constant

Conductivity

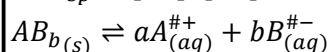
20 drops = 1 mL

Ionic/Molecular Solutions

$$K_{sp} = [A^+]^a [B^-]^b$$

-More soluble

$K_{sp}$  (Table)



$-\uparrow [ ]$

-More ions

Reactant is a Solid

Solubility Table In Water (Ability/(In) to be soluble) Solubility Product Table (Higher #, Higher Solubility)

Concentration

Precipitate :  $Q < K_{sp}$  Cannot Form

$Q = K_{sp}$  Barely saturated solution formed

$Q > K_{sp}$  Will Form

Altering Solubility

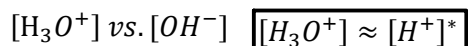
# C12 - 0.0 - Methods 12 Acids/Electro

*pH* - Level of acidity or alkalinity of a solution.

Acids, Bases & Salts

Indicators Table

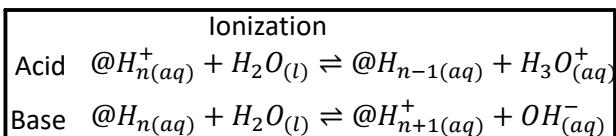
Strength Table - Higher #, Higher Strength of Acid (To give an  $H^+$ )



hydronium vs hydroxide

$$K_{eq} = \frac{K_a(\text{reactant acid})}{K_a(\text{product acid})}$$

Acid - Donor  $p^+$   
Base - Acceptor  $p^+$



As *pH* increases,  $[H_3O^+]$  decreases. (Less Acidic)  $\uparrow K_a$   
As *pOH* increases,  $[OH^-]$  decreases. (More Basic)  $\uparrow K_b$

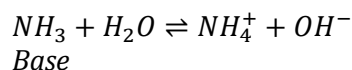
Stronger Acid favours other side.

$\downarrow pH$ ,  $\uparrow$  Ionization,  $\uparrow$  Conductivity

% Dissociation =  $\frac{[H^+]}{[@]}$       % Ionization =  $\frac{[OH^-]}{[@]}$

Amphiprotic anions compare  $K_a$  &  $K_b$

An Acid if placed in water, causes the  $[H_3O^+]$  to increase and the  $[OH^-]$  to decrease.  
A Base if placed in water, causes the  $[OH^-]$  to increase and the  $[H_3O^+]$  to decrease.



Stronger an Acid, the weaker it's conjugate base (And Vice Versa)

Amphiprotic

Arrhenius

Bronsted Lowrey

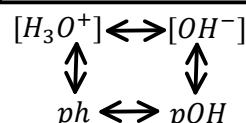
Lewis

Hydrolysis

Titrations

Buffers

$$K_w = [H_3O^+] \times [OH^-]$$



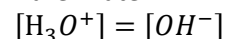
$$pH = -\log_{10}[H_3O^+]$$

$$pOH = -\log_{10}[OH^-]$$

$$pK_w = pH + pOH = 14$$

$$pK_w = -\log[K_w]$$

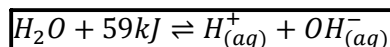
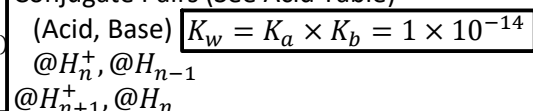
Pure Water\*



Ionization Constants

$K_a$  : Acid,  $K_b$  : Base

Conjugate Pairs (See Acid Table)



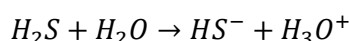
$\uparrow$ Temp, Shifts Right

$\uparrow [H_3O^+]$  &  $\uparrow [OH^-]$

Still Neutral  $[H_3O^+] = [OH^-]$

If  $\uparrow [H^+]$ ,  $\downarrow pH = -\log[H^+]$  below 7\*

Still Neutral  $[H^+] = [OH^-]$



$\uparrow$  Polarity H - A  $\rightarrow \uparrow$  Acidity\*

Weak Acid - Ion bond strongly to a proton.

Electrochemistry

Reduction Table - (Higher #, Higher Ability for itself to reduce)

Voltage

Oxidization/#/ $\Delta ON$

Major  
OHe

Agents

Anode

Cathode

Batteries

Lead-Acid

Cell

Redox

Cathode

Reducing Happens

Zinc-Carbon

Alkaline Dry-Cell

Predicting Spontaneity

$Cell\ Voltage = |E^o - E^o|$

Fuel Cell

Stronger Acid favours  
other side

Higher Acidity - Higher Conductivity

Reduction Potential

-Ability\* to be reduced

(Strength of Oxidization Agent)

History\* - Oxygen is good  
at removing electrons.

Platinum is Inert\*

Half-Reactions

