C12 - 0.0 - Methods 12 Rates/Eq/Solubility

LABS

Ripple Effects

Reaction Kinetics

Speed
$aq^{\pm}$>g&l>s

Rates/Factors
↑↓∝

Takes Energy to break a bond

Enthalpy
-Exo/Endothermic
-Activation Energy - E_a
-ΔH

Le Chatelier's Principle -
Nature Undoes*

Reactant Nature (Chem)
Amount/Concentration
Energy/Energy
Orientation/Geometry
SA/V(Open?)/Size/pH
Temp/Pressure(g)
Catalyst (1stR/next*P)

Allotropes - Forms
Spectators

Reaction Mechanism (Steps)

$$2NO_2 \rightarrow NO + NO_3$$
$$CO + NO_3 \rightarrow NO_2 + CO_2$$
$$NO_2 + CO \rightarrow NO + CO_2$$

Reaction Intermediaries (Activated Complex)

Overall Equation

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) : $\boxed{K_f = K_r}$ f - Forward, r - Reverse

$$\frac{[Products]}{[Reactants]} = \frac{K_{fi}}{K_{ri}}$$
i − initial

Equilibrium Constant - K_{eq} ([ ]'s Constant at Eq)

$$K_{eq} = \frac{[C]^c \times [D]^d \times \cdots}{[A]^a \times [B]^b \times \cdots} \boxed{\frac{[Products]}{[Reactants]}}$$

$K_{eq} > 1$ - [products]>[reactants]
$K_{eq} = 1$ - [products]=[reactants]
$K_{eq} < 1$ - [products]<[reactants]

$\boxed{aA + bB + \cdots \rightleftharpoons cC + dD + \cdots}$ Pure solids/liquids are not included in K_{eq}

ICE Tables

$\frac{mol}{L}$	$2@_{(g)}$ +	$@_{2(l)}$ ⇌	$2@_{2(aq)}$
Start	0	0	#
$+\Delta$			
$= Eq$	$0.\#\#$		

$\boxed{Mole\ Ratio\ -\ 2:\ 1:\ 2}$

$Q = Reaction\ Quotient\ (Trial\ K)$ $\boxed{Q = K_{eq}\ ; Eq}$ $Q < K_{eq}$ Shift to Products ↑ $\frac{[Products]}{[Reactants]}$ ↓ $Q > K_{eq}$ Shift to Reactants

Moles Reactants used up = Moles Products are made

↓Temp shift to side with heat (Larger hurdle to cross)

Only Temp can affect K_{eq}.
Solids and Pure Liquids exempt.

5% $Rule$
$Assume\ x \ll \#$

↑[Reactant]→↑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.
...

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

Spontaneous reaction occurs when
$\Delta S > 0$ & $\Delta H < 0$

↑Moles→↑Entropy

More Entropy Shifts ???

Solubility

Ionic/Molecular Solutions

K_{sp} (Table)

Solubility Product Constant
$$K_{sp} = [A^+]^a[B^-]^b$$
$$AB_{b(s)} \rightleftharpoons aA^{\#+}_{(aq)} + bB^{\#-}_{(aq)}$$

Reactant is a Solid

Conductivity
-More soluble
-↑[ ]
-More ions

$20\ drops = 1\ mL$

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

Concentration

Altering Solubility

Precipitate : $Q < K_{sp}$ Cannot Form | $\boxed{Q = K_{sp}}$ Barely saturated solution formed | $Q > K_{sp}$ Will Form

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^+)

$[H_3O^+]$ $vs.$ $[OH^-]$ $\boxed{[H_3O^+] \approx [H^+]^*}$

hydronium vs hydroxide

Amphiprotic
Arrhenius
Bronsted Lowrey
Lewis

Hydrolysis

Titrations

Buffers

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

$[H_3O^+] \leftrightarrow [OH^-]$
$\updownarrow \quad \updownarrow$
$ph \leftrightarrow pOH$

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

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

$\boxed{pK_w = pH + pOH = 14}$

$pK_w = -\log[K_w]$

Pure Water*
$[H_3O^+] = [OH^-]$

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

$$HReact + Prod^{-1} \rightleftharpoons React^- + HProd$$

Ionization Constants
K_a : Acid , K_a : Base

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

Ionization	
Acid	$@H^+_{n(aq)} + H_2O_{(l)} \rightleftharpoons @H_{n-1(aq)} + H_3O^+_{(aq)}$
Base	$@H_{n(aq)} + H_2O_{(l)} \rightleftharpoons @H^+_{n+1(aq)} + OH^-_{(aq)}$

Conjugate Pairs (See Acid Table)
(Acid, Base) $\boxed{K_w = K_a \times K_b = 1 \times 10^{-14}}$
$@H^+_n, @H_{n-1}$
$@H^+_{n+1}, @H_n$

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

$H_2O + 59kJ \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

$\uparrow$Temp, Shifts Right
$\uparrow [H_3O^+]$ & $\uparrow [OH^-]$
Still Neutral $[H_3O^+] = [OH^-]$
If $\uparrow [H^+]$,$\downarrow pH = -\log[H^+]$ below 7*
Still Neutral $[H^+] = [OH^-]$

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

$H_2S + H_2O \rightarrow HS^- + H_3O^+$

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

Amphiprotic anions compare K_a & K_b

Weak Acid - Ion bond strongly to a proton.

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.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$
$Base$

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

Electrochemistry

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

Voltage

Oxidization/#/ΔON

Major
OHe

Agents

Anode
Cathode

Batteries
Lead-Acid
Zinc-Carbon
Alkaline Dry-Cell
Fuel Cell

Cell

Redox

Cathode
Reducing Happens

Predicting Spontaneity

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

Reduction Potential
-Ability* to be reduced
(Strength of Oxidization Agent)

Stronger Acid favours other side

Higher Acidity - Higher Conductivity

History* - Oxygen is good at removing electrons.

Platinum is Inert*

Half-Reactions
Anode - $Zn \rightarrow Zn^{2+} + 2e^-$
Cathode - $Pb^{2+} + 2e^- \rightarrow Pb$