



Introductory Chemistry I

MASS, MOLES, & CONCENTRATIONS

NUMBER OF MOLES OF A

$$n_A = \frac{\text{Mass, } m_A}{\text{Molar mass, } M_A} = \frac{\# \text{ of particles}}{6.022 \times 10^{23}}$$

MOLE FRACTION OF A

$$X_A = \frac{\text{Moles of A}}{\text{Total moles}} = \frac{n_A}{\sum n_i}$$

MASS FRACTION OF A

$$m_{fA} = \frac{\text{Mass of A}}{\text{Total mass}} = \frac{m_A}{\sum m_i}$$

PARTS PER MILLION

$$ppm = \frac{\text{Mass of substance}}{\text{Total mass}} \cdot 10^6$$

MOLARITY

$$[M] = \frac{\text{Moles dissolved solute}}{\text{Liters solution}}$$

MOLALITY

$$m = \frac{\text{Moles dissolved solute}}{\text{Kilograms solvent}}$$

CONVERSIONS

$$1 \text{ ppm} = 10^{-4} \text{ mass percent} = 10^{-6} \text{ mass fraction}$$

$$X_A = \frac{n_A}{\sum n_i} \Rightarrow \frac{n_A M_A}{\sum n_i M_i} = m_{fA} \quad \text{where: } n_A M_A = m_A, \sum n_i M_i = \sum m_i$$

$$m_{fA} = \frac{m_A}{\sum m_i} \Rightarrow \frac{m_A / M_A}{\sum m_i / M_i} = X_A \quad \text{where: } m_A / M_A = n_A, \sum m_i / M_i = \sum n_i$$

$$\text{Molarity} \times \text{Liters solution} / \text{Total moles} = X_A$$

$$\frac{\text{Molarity}}{\text{Mass fraction of solvent} \times \text{Density of solution (kg/L)}} = m$$

THE GASEOUS STATE

TEMPERATURE

$$^{\circ}\text{C (Celsius)} = 5/9(^{\circ}\text{F} - 32)$$

$$^{\circ}\text{F (Fahrenheit)} = (9/5)^{\circ}\text{C} + 32$$

$$\text{K (Kelvin)} = ^{\circ}\text{C} + 273.15$$

$$^{\circ}\text{R (Rankin)} = ^{\circ}\text{F} + 459.67$$

Examples (K)

10 ⁸	-	H-Bomb
2000	-	Bunsen burner (CH ₄ /O ₂)
1000	-	Cigarette
373	-	Boiling point of H ₂ O
310	-	Body temperature
273	-	Freezing point of H ₂ O
90	-	Boiling point of O ₂
4.2	-	Boiling point of He
0	-	Absolute zero

GAS MIXTURES

Dalton's Law of Partial Pressures

- Partial pressure, p_i , = pressure exerted by individual component at T and V of mixture

$$\sum p_i = P_{\text{total}} \quad (T, V \text{ are constant})$$

Gas 1	Gas 2	Gas 1 & 2
1 L	1 L	1 L
T, P_1	T, P_2	T, P_{Total}

Average Molar Mass of Mixture

$$M_{\text{avg}} = \sum n_i \cdot M_i / n_{\text{tot}}$$

Amagat's Law of Partial Volumes

- Partial volume, V_i , = volume occupied by individual component at T and P of mixture

$$\sum V_i = V_{\text{total}} \quad (T, P \text{ are constant})$$

Gas 1	Gas 2	Gas 1 & 2
1 L	0.8 L	1.8 L
T, P	T, P	T, P

Density of Ideal Mixture

$$d_{\text{mixture}} = P \cdot M_{\text{avg}} / RT$$

PRESSURE

$$1 \text{ atm} = 101325 \text{ Pa}$$

$$= 1.01325 \text{ bar}$$

$$= 14.6959 \text{ psi}$$

$$= 760 \text{ mm Hg}$$

$$= 760 \text{ Torr}$$

Examples (atm)

10 ⁷	-	Center of sun
10 ⁶	-	Center of earth
500-1000	-	NH ₃ synthesis
150	-	Gas cylinder
50-75	-	CH ₄ pipeline
1	-	Atmosphere
10 ⁻⁶	-	Vapor pressure of Hg
10 ⁻¹²	-	Laboratory vacuum
100 molecules/cm ³	-	(outer space)

MOLECULAR PROPERTIES

Average Molecular Velocity

$$u = (3RT/M)^{1/2}$$

$$u_{N_2} (25^{\circ}\text{C}) \cong 500 \text{ m/s}$$

of Impacts per Second on 1 m²

$$C = 1/4(N/V)u \quad (N = \# \text{ of molecules})$$

$$C_{N_2} \cong 10^{28} \text{ collisions/m}^2 \cdot \text{s} \quad (\text{at } 1 \text{ atm, } 25^{\circ}\text{C})$$

Average Distance Between Molecules = Mean Free Path

$$\lambda = [\pi D_m^2 \cdot (N/V)]^{-1}$$

$$D_m = \text{diameter of molecule}$$

$$\lambda_{N_2} (1 \text{ atm}) \cong 10^{-7} \text{ m}$$

Diffusion Coefficient

$$D = 1/3u\lambda$$

$$D_{N_2} (1 \text{ atm, } 25^{\circ}\text{C}) \cong 0.17 \text{ cm}^2/\text{s}$$

GRAHAM'S LAW OF DIFFUSION/EFFUSION

$$\frac{\text{Rate of Transport of 1}}{\text{Rate of Transport of 2}} = \sqrt{\frac{\text{Molar Mass of 2}}{\text{Molar Mass of 1}}}$$

SINGLE GASES

1 Boyle's Law $PV = \text{Constant}$ (at constant T, n)

2 Law of Charles & Gay-Lussac $V/T = \text{Constant}$ (at constant P, n)

3 Avogadro's Hypothesis At constant T and P , equal volumes of different gases contain the same number of particles
 $V/n = \text{Constant}$ (at constant T, P)

4 Ideal Gas Law $PV = nRT$ (at low P , high T)

5 Molar Volume of an Ideal Gas $V_{\text{Molar}} = 22.414 \text{ L}$ at STP (STP = 0°C, 1 atm)

6 Density of an Ideal Gas $d = PM/RT$

7 Gas Constant, R $R = 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}$
 $= 82.0578 \text{ cm}^3 \cdot \text{atm} / \text{mol} \cdot \text{K}$
 $= 8.3145 \text{ J} / \text{mol} \cdot \text{K}$
 $= 8.3145 \text{ L} \cdot \text{kPa} / \text{mol} \cdot \text{K}$
 $= 1.9859 \text{ cal}(\text{I.T.}) / \text{mol} \cdot \text{K}$
 $= 0.73024 \text{ ft}^3 \cdot \text{atm} / \text{lb} \cdot \text{mol} \cdot ^{\circ}\text{R}$

VAN DER WAALS EQUATION FOR REAL GASES

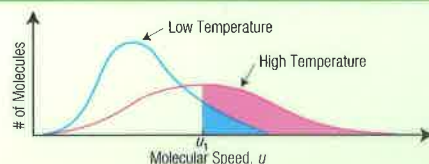
$$\left(P + \frac{a n^2}{V^2} \right) (V - nb) = nRT$$

$\overbrace{\left(P + \frac{a n^2}{V^2} \right)}$ pressure correction term
 $\underbrace{(V - nb)}$ volume correction term

a = measure of attractive force

b = measure of repulsive force or excluded volume

MAXWELL'S VELOCITY DISTRIBUTION OF GASEOUS MOLECULES



- As temperature increases, the average molecular speed increases, causing the curve to shift to the right and become broader

THE LIQUID & SOLID STATES

PROPERTIES OF LIQUIDS & CRYSTALLINE SOLIDS

	Orderedness	Diffraction Pattern	Rigidity	Molecular Movement	Separation Energy
Liquids	Low	Diffuse	Yielding	Yes	Low
Solids	High	Sharp	Rigid	No	Low to high

COMMON CRYSTAL STRUCTURES



Elements
O, F, P_{white},
In, Sn_{white},
Po

Simple Cubic



Alkali Metals
Li, Na, K, Rb
Ferrous Metals
Fe, Cr, Mn, V, Mo

Body-Centered Cubic



Noble Metals
Ag, Au, Pt, Pd, Rh
Other Metals
Al, Ca, Sr, Ni, Cu

Close-Packed Cubic (Face-Centered)



Hydrogen Halides
HCl, HBr, HI
Metals
Be, Mg, Zn

Close-Packed Hexagonal

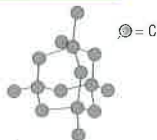
CRYSTAL TYPES & PROPERTIES

Bond Type	Ionic	Covalent	Molecular	Metallic
Particles	Cations & anions	Atoms	Molecules or atoms	Atoms
Hardness	Hard, brittle	Hard	Soft	Soft to hard
Melting	High	High	Low	Low to high
Conduction*	Poor	Poor**	Poor	Good
Examples	NaCl	Diamond	Water	Li, Fe

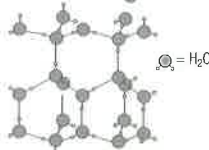
*Heat and electrical conduction **Except diamond products

EXAMPLES

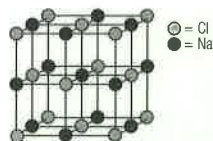
Diamond



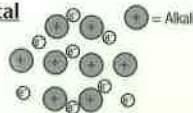
Water



Sodium Chloride

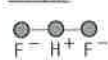


Alkali Metal

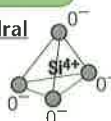


LOCAL PACKING ARRANGEMENTS

Linear



Tetrahedral



Triangular



Coordination Number

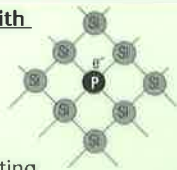
Coordination Number	Cation/Anion Radius
2	0 - 0.155
3	0.155 - 0.225
4	0.225 - 0.414

- At the lower limit of the radius ratio, cations and anions touch each other
- As the ratio increases, the structure opens up
- At the upper limit, transition to the next higher coordination number occurs
- For ratios of 0.414 - 0.732, the structure becomes octahedral; beyond this range, the structure becomes cubic (coordination number = 8)

SEMI-CONDUCTORS: SILICON CHIPS

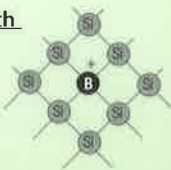
Silicon Doped with Phosphorus (P)

- P, the donor impurity, has one more valence electron than Si, thus creating a "negative center"
- Hence, the term **n** (for negative)-type semiconductor



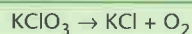
Silicon Doped with Boron (B)

- B, the acceptor impurity, has one less valence electron than Si, thus creating a "positive hole"
- Hence, the term **p** (for positive)-type semiconductor

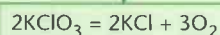
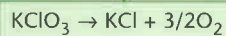


CHEMICAL EQUATIONS

BALANCING SIMPLE CHEMICAL EQUATIONS



- First balance oxygen
- Then balance potassium and chlorine
- Finally, go back and re-balance oxygen



Tip: When balancing simple chemical equations, start with elements that occur the least number of times

CALCULATING MASS, MOLES & CONCENTRATIONS

Principle: Ratio of number of moles = ratio of number of particles

Irreversible Reaction

	Zn	+	2SO ₂	=	ZnS ₂ O ₄
Mass (g)	x		y		1000
Moles	x/65.4		y/64		1000/193.4
Mole Ratios	$\frac{x/65.4}{1000/193.4} = 1:1$	\Rightarrow	x = 338 g Zn		
	$\frac{y/64}{1000/193.4} = 2:1$	\Rightarrow	y = 662 g SO ₂		

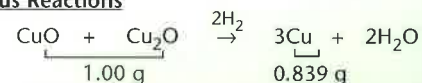
Irreversible Reaction with Limiting Reagent

	Zn	+	2SO ₂	=	ZnS ₂ O ₄
Initial Mass (g)	130.8		128		0
Initial Moles	$\frac{130.8}{65.4} = 2$		$\frac{128}{64} = 2$		0
Moles Left Over	1		0 (limiting)		1

Reversible Reaction

	CaCO ₃	\rightleftharpoons	CaO	+	CO ₂
Initial Moles	2		0		0
Reacted Moles	-0.5		+0.5		+0.5
Moles Left Over	1.5		0.5		0.5

Simultaneous Reactions



	CuO	+	H ₂	=	Cu	+	H ₂ O
Mass (g)	x				y		
Moles	x/79.5				y/63.5		
Reaction 2	Cu ₂ O	+	H ₂	=	2Cu	+	H ₂ O
Mass (g)	1-x				0.839-y		
Moles	$1-x/143$				$(0.839-y)/63.5$		
Mole Ratios	$\frac{x/79.5}{y/63.5} = 1:1$				x = 0.545 g CuO		
	$\frac{(1-x)/143}{(0.839-y)/63.5} = 1:2$				1-x = 0.455 g Cu ₂ O		

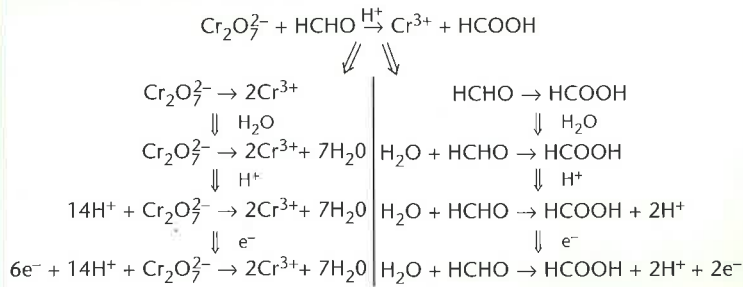
Titration

	5Fe ²⁺	+	MnO ₄ ⁻	\rightarrow	Mn ²⁺	+	5Fe ³⁺
Volume (mL)	20.8		16				
Molarity	x		0.130				
Moles	$\frac{20.8 \times x}{10^3}$		$\frac{16.0 \times 0.130}{10^3}$				
Mole Ratios	$\frac{20.8 \times x}{16.0 \times 0.130} = 5:1$				\Rightarrow	x = 0.50 M	

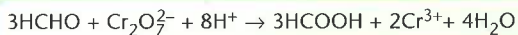
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BALANCING REDOX EQUATIONS

Principle: Split reaction into oxidation/reduction half-reactions and balance each in turn with H₂O, H⁺, and electrons (e⁻) • Recombine half-reaction equations algebraically so as to cancel e⁻s (electrons are conserved in redox reactions)



• Multiply right side by 3 (electrons must cancel) and add:



CHEMICAL EQUILIBRIA

GENERAL CHARACTERISTICS

• The same principles apply as for Phase Equilibria • In addition:

- (I) The concentrations at equilibrium stand in a fixed ratio called the **Equilibrium Constant, K_c**:

$$aA + bB \rightleftharpoons cC + dD \quad K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}_{eq}$$
- (II) The concentrations at the start of a reaction stand in a similar constant ratio called the **Reaction Quotient, Q**:

$$aA + bB \rightleftharpoons cC + dD \quad Q_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}_{initial}$$
- (III) If Q_c < K_c then the reaction proceeds from left to right
 If Q_c = K_c then the system is at equilibrium
 If Q_c > K_c then the reaction proceeds from right to left

THE EQUILIBRIUM CONSTANT, K

- (I) For gaseous reactions, K is often expressed in terms of partial pressures: $K_p = \frac{[P_C]^c [P_D]^d}{[P_A]^a [P_B]^b}$
- (II) The relation between K_p and K_c is given by:

$$K_p = K_c (RT)^{\Delta n}$$

$$\Delta n = (c + d) - (a + b)$$
- (III) In heterogeneous reactions, solid concentrations are constant and can be included in the constant, K:

$$\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2$$

$$K_c = [\text{CO}_2] \quad K_p = P_{\text{CO}_2}$$
- (IV) Unknown K's can be calculated from known K's of component reactions:

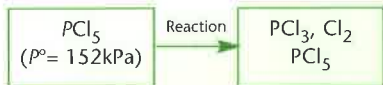
$$2\text{NO} + \text{O}_2 \rightleftharpoons 2\text{NO}_2 \quad K_1 = [\text{NO}_2]^2 / [\text{NO}]^2 [\text{O}_2]$$

$$2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4 \quad K_2 = [\text{N}_2\text{O}_4] / [\text{NO}_2]^2$$

$$2\text{NO} + \text{O}_2 \rightleftharpoons \text{N}_2\text{O}_4 \quad K_x = K_1 \cdot K_2$$

CALCULATIONS USING K

- Calculate p's and P_{Tot} at equilibrium and the % yield
- K_p = 116 (T = constant)



	PCl_5	\rightleftharpoons	PCl_3	$+$	Cl_2
Initial	152		0		0
Reaction	-X		+X		+X
At Eq	152 - X		+X		+X

$$K_p = 116 = \frac{X^2}{152 - X} \Rightarrow X^2 + 116X - (152)(116) = 0$$

Solve for X, P_{Tot}, and % yield

$$X = P_{\text{PCl}_3} = P_{\text{Cl}_2} = 87.3 \text{ kPa}$$

$$P_{\text{Tot}} = (152 - 87.3) + 2 \cdot 87.3 = 239.3 \text{ kPa}$$

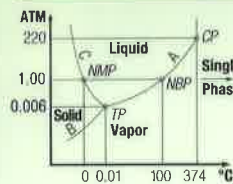
$$\% \text{ yield} = \frac{87.3}{152} (100) = 57.4\%$$

PHASE EQUILIBRIA

GENERAL CHARACTERISTICS

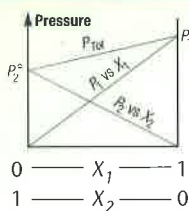
- (I) Temperature, pressure, composition, and other properties of the system do not change with time
- (II) Equilibrium results when the rates of two opposing processes (e.g., freezing/melting) become equal
- (III) Properties of a particular equilibrium state are independent of how that state was reached
- (IV) **Le Chatelier's Principle:** A system at equilibrium reacts to a change in a variable by minimizing that change

SINGLE COMPONENT SYSTEMS: THE WATER PHASE DIAGRAM



- A..... Boiling point or vapor pressure curve
- B..... Sublimation or solid vapor pressure curve
- C..... Melting point or freezing point curve
- NBP/...Boiling/melting point at NMP "normal" pressure of 1 atm
- CP..... Critical Point
- TP.....Triple Point

TWO COMPONENT SYSTEMS: VAPOR-LIQUID EQUILIBRIA



Raoult's Law of Ideal Solutions

- The partial pressure, P_i, of a component in the vapor phase equals mole fraction, X_i, in the liquid phase multiplied by the pure component vapor pressure, P_i^o:

$$P_1 = X_1 \cdot P_1^o$$

$$P_2 = X_2 \cdot P_2^o$$

CALCULATIONS USING RAULT'S LAW

- Given P₁^o, P₂^o, X₁, at constant T, calculate X₂, P₁, P₂, P_{Tot}, and the vapor phase mole fractions, Y₁ and Y₂

Solve X₂ $X_1 + X_2 = 1 \Rightarrow X_2 = 1 - X_1$

Solve P₁, P₂, P_{Tot} $P_1 = X_1 P_1^o \Rightarrow P_2 = X_2 P_2^o \Rightarrow P_{\text{Tot}} = P_1 + P_2$
Raoult's Law Dalton's Law

Apply Ideal Gas Laws $P_1 = n_1 RT/V \Rightarrow P_2 = n_2 RT/V \Rightarrow P_{\text{Tot}} = n_{\text{Tot}} RT/V$
Ideal Gas Laws

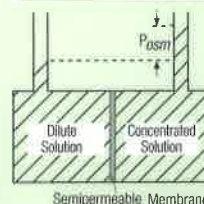
Solve for Y₁, Y₂ $Y_1 = P_1/P_{\text{Tot}} \Rightarrow Y_1 + Y_2 = 1 \Rightarrow Y_2 = 1 - Y_1$

COLLIGATIVE PROPERTIES

- Properties of phase equilibria which depend on the concentration of a non-volatile solute

Dissolved Substance	Changes Depend on m, M
Lowers the vapor pressure of solvent	$\Delta P_v = K_v m$
Raises the boiling point of solvent	$\Delta T_b = K_b m$
Lowers the freezing point of solvent	$\Delta T_f = K_f m$
Responsible for osmotic pressure	$P_{osm} = MRT$

OSMOTIC PRESSURE



- Water permeates from left to right across the semipermeable membrane until P_{osrn} stops the flow
- In reverse osmosis, extra pressure is applied on the concentrated side of the membrane, forcing water to flow from right to left

- An example of reverse osmosis is desalinated seawater

IONIC EQUILIBRIA (THE SOLUBILITY PRODUCT, K_{sp})

SPARINGLY SOLUBLE SOLIDS

- K_{sp} = Equilibrium constant of sparingly soluble ionic substances
- $$\text{Mg(OH)}_2(s) \rightleftharpoons \text{Mg}^{2+} + 2\text{OH}^-$$
- $$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 1.2 \times 10^{-11}$$

CALCULATING ION CONCENTRATIONS

$$\text{Mg(OH)}_2 \rightleftharpoons \text{Mg}^{2+} + 2\text{OH}^-$$

$$[\text{Mg}^{2+}]/[\text{OH}^-] = 1/2 \quad \therefore [\text{Mg}^{2+}] = \frac{1}{2}[\text{OH}^-]$$

$$K_{sp} = \frac{1}{2}(\text{OH}^-)^3 = 1.2 \times 10^{-11}$$

$$\therefore [\text{OH}^-] = 2.9 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

$$[\text{Mg}^{2+}] = \frac{1}{2}[\text{OH}^-] = 1.45 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

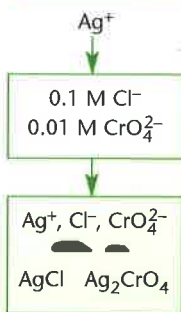
CALCULATING SOLUBILITY, s

- s = Mass of dissolved solids (g/L)
- $$\{\text{Mg(OH)}_2\}_{\text{dissolved}} \rightarrow [\text{Mg}^{2+}] = 1.45 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$
- $$s = 1.45 \times 10^{-4} \times \text{Molar Mass} = 8.5 \times 10^{-3} \frac{\text{g}}{\text{L}}$$

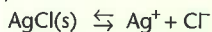
COMMON ION EFFECT

- An increase in the ion concentrations reduces the solubility (Le Chatelier)
- Solubility of Mg(OH)_2 in 1 M NaOH:
- $$[\text{Mg}^{2+}][1]^2 = 1.2 \times 10^{-11} \Rightarrow s = 7 \times 10^{-10} \frac{\text{g}}{\text{L}}$$

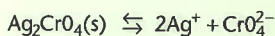
SELECTIVE PRECIPITATION



Example: Pour solution of silver ions (i.e., AgNO_3) into a beaker containing chloride and chromium anions • Which will precipitate first, AgCl or Ag_2CrO_4 ? • How complete is the separation? • Possible equilibria are:



$$K_{sp} = 2.8 \times 10^{-10}$$



$$K_{sp} = 1.9 \times 10^{-12}$$

- $[\text{Ag}^+]$ required for precipitation to start

$$[\text{Ag}^+]_{\text{AgCl}} = \frac{K_{sp}}{[\text{Cl}^-]} = \frac{2.8 \times 10^{-10}}{0.1} = 2.8 \times 10^{-9} \quad \text{AgCl precipitates first}$$

$$[\text{Ag}^+]_{\text{Ag}_2\text{CrO}_4} = \left\{ \frac{K_{sp}}{[\text{CrO}_4^{2-}]} \right\}^{1/2} = \left\{ \frac{1.9 \times 10^{-12}}{0.01} \right\}^{1/2} = 1.4 \times 10^{-5}$$

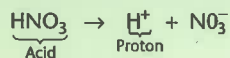
- $[\text{Cl}^-]$ remaining when Ag_2CrO_4 starts to precipitate

$$[\text{Cl}^-]_{\text{rem}} = \frac{K_{sp}}{[\text{Ag}^+]} = \frac{2.8 \times 10^{-10}}{1.4 \times 10^{-5}} = 2 \times 10^{-5} \Rightarrow 0.002\%$$

IONIC EQUILIBRIA (ACIDS & BASES)

LOWRY-BRONSTED DEFINITIONS

Bronsted Acid = Proton Donor



Bronsted Base = Proton Acceptor



ION PRODUCT OF WATER, K_w

- K_w = Dissociation equilibrium constant of water
- $$2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$$
- or
- $$\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$$
- $$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14} \quad (\text{at } 25^\circ\text{C})$$
- $$K_w = [\text{H}^+][\text{OH}^-] = 10^{-14} \quad (\text{at } 25^\circ\text{C})$$

THE pH SCALE

$$\text{pH} = -\log[\text{H}^+] \quad \text{pOH} = -\log[\text{OH}^-]$$

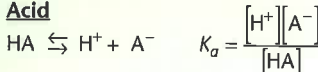
$$\text{pH} + \text{pOH} = 14 \quad (\text{at } 25^\circ\text{C})$$

- For pure water at 25°C :

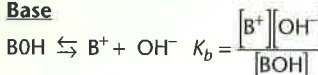
$$\text{pH} = \text{pOH} = 7 \quad \text{and} \quad [\text{H}^+] = [\text{OH}^-] = 10^{-7} \frac{\text{mol}}{\text{L}}$$

DISASSOCIATION CONSTANTS OF WEAK ACIDS & BASES

Acid



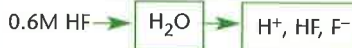
Base



Indicators are Weak Acids or Bases



CALCULATING ION CONCENTRATIONS WITH K_a



	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$	
Initial:	0.6	0
Reaction:	-X	+X
At Eq:	0.6 - X	+X

$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]} = \frac{X^2}{0.6 - X} = 7.1 \times 10^{-4}$$

$$X^2 + 7.1 \times 10^{-4}X - 4.32 \times 10^{-4} = 0$$

$$X = [\text{H}^+] = [\text{F}^-] = 0.021 \frac{\text{mol}}{\text{L}}$$

SOLUTIONS OF A SALT & ITS PARENT WEAK ACID OR BASE: BUFFERS

- Buffer solutions minimize changes in pH caused by addition of an acid or base
- Their pH can be calculated from the Henderson-Hasselbach equation

$$\text{pH} = \log \frac{[\text{A}^-]}{[\text{HA}]} - \log K_a$$

$$\text{pOH} = \log \frac{[\text{B}^+]}{[\text{BOH}]} - \log K_b$$

pH OF SOME FLUIDS

HCl (1M)	0.0
Vinegar	3.0
Milk	6.5
Pure water	7.0
Blood	7.4 ± 0.05
NaOH (1M)	14.0
Gastric juice	1 - 2
Lemon juice	2.4
Urine	4.8 - 7.5
Saliva	6.4 - 6.9

STRONG ACIDS & BASES (TOTAL DISSOCIATION)

HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	Ca(OH)_2
H_2SO_4	Ba(OH)_2

WEAK ACIDS & BASES (PARTIAL DISSOCIATION)

HF	NH_3
HNO_2	CH_3NH_2
HCN	$\text{C}_2\text{H}_5\text{NH}_2$
HCOOH	$\text{CO(NH}_2)_2$
CH_3COOH	$\text{C}_6\text{H}_5\text{N}$
Aspirin	Caffeine

SOLUTIONS OF SALTS OF A WEAK ACID OR BASE: HYDROLYSIS

- Salts of weak acids and bases initially dissociate completely, then hydrolyze ("split") water, resulting in a basic or acidic solution
- The hydrolysis constant can be calculated from the component K 's (i.e., K_a or K_b) and K_w



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