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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

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

MOLARITY

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

MOLALITY

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

CONVERSIONS

1ppm = 10^4 mass percent = 10^6 mass fraction

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

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

Molarity × Liters solution/Total moles = X_A

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

THE GASEOUS STATE

TEMPERATURE

°C (Celsius)	=	5/9(°F – 32)
°F (Fahrenheit)	=	(9/5)°C + 32
K (Kelvin)	=	°C + 273.15
°R (Rankin)	=	°F + 459.67

Examples (K)

10^8	–	H-Bomb
2000	–	Bunsen burner (CH_4/O_2)
1000	–	Cigarette
373	–	Boiling point of H_2O
310	–	Body temperature
273	–	Freezing point of H_2O
90	–	Boiling point of O_2
4.2	–	Boiling point of He
0	–	Absolute zero

PRESSURE

1 atm = 101325 Pa
= 1.01325 bar
= 14.6959 psi
= 760 mm Hg
= 760 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)

SINGLE GASES

❶ Boyle's Law	PV = Constant (at constant T, n)
❷ Law of Charles & Gay-Lussac	V/T = Constant (at constant P, n)
❸ Avogadro's Hypothesis	At constant T and P, equal volumes of different gases contain the same number of particles V/n = Constant (at constant T, P)
❹ Ideal Gas Law	$PV = nRT$ (at low P, high T)
❺ Molar Volume of an Ideal Gas	V_{Molar} = 22.414 L at STP (STP = 0°C, 1 atm)
❻ Density of an Ideal Gas	$d = PM/RT$
❼ Gas Constant, R	R = 0.08206 L•atm/mol•K = 82.0578 cm^3•atm/mol•K = 8.3145 J/mol•K = 8.3145 L•kPa/mol•K = 1.9859 cal(I.T.)/mol•K = 0.73024 ft^3•atm/lb-mol•°R

GAS MIXTURES

Dalton's Law of Partial Pressures

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

$\sum pi = P_{total}$ (T, V 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

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

Amagat's Law of Partial Volumes

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

$\sum V_i = V_{total}$ (T, P 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

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

MOLECULAR PROPERTIES

Average Molecular Velocity

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

u_{N_2} (25°C) ≅ 500 m/s

# of Impacts per Second on 1 m^2

$C = 1/4(N/V)u$ (N = # of molecules)

$C_{N_2} \cong 10^{28}$ collisions/m^2•s (at 1 atm, 25°C)

Average Distance Between Molecules = Mean Free Path

$\lambda = [\pi D^2_m \bullet (N/V)]^{-1}$

D_m = diameter of molecule

λ_{N_2} (1 atm) ≅ 10^{-7} m

Diffusion Coefficient

$D = 1/3u\lambda$

D_{N_2} (1 atm, 25°C) ≅ 0.17 cm^2/s

GRAHAM'S LAW OF DIFFUSION/EFFUSION

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

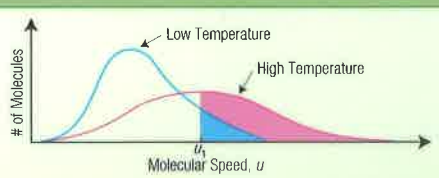
VAN DER WAALS EQUATION FOR REAL GASES

$$\underbrace{\left(P + an^2/V^2\right)}_{\text{pressure correction term}}\underbrace{(V - nb)}_{\text{volume correction term}} = nRT$$

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

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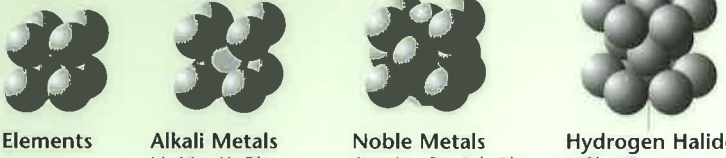
THE LIQUID & SOLID STATES

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

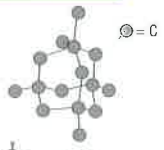
CRYSTAL TYPES & PROPERTIES

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

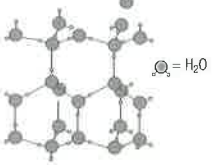
*Heat and electrical conduction **Except diamond products

EXAMPLES

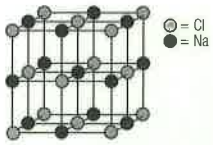
Diamond



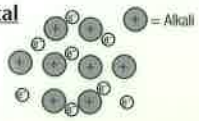
Water



Sodium Chloride



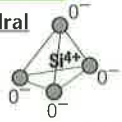
Alkali Metal



LOCAL PACKING ARRANGEMENTS

Linear

Tetrahedral



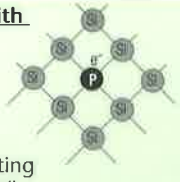
Triangular

Coordination Number	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)

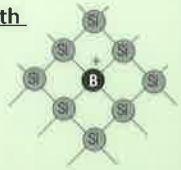
SEMI-CONDUCTORS: SILICON CHIPS

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

$KClO_3 \rightarrow KCl + O_2$

$\downarrow$

$KClO_3 \rightarrow KCl + 3/2O_2$

$\downarrow$

$2KClO_3 = 2KCl + 3O_2$

- 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_2O_4
Mass (g)	x		y		1000
Moles	$x/65.4$		$y/64$		1000/193.4

Mole Ratios

$$\frac{x/65.4}{1000/193.4} = 1{:}1 \Rightarrow x = 338 \text{ g Zn}$$

$$\frac{y/64}{1000/193.4} = 2{:}1 \Rightarrow y = 662 \text{ g } SO_2$$

Irreversible Reaction with Limiting Reagent

	Zn	+	$2SO_2$	=	ZnS_2O_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

	$CaCO_3$	$\rightleftarrows$	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

$$\underbrace{CuO}_{1.00\text{ g}} + Cu_2O \xrightarrow{2H_2} \underbrace{3Cu}_{0.839\text{ g}} + 2H_2O$$

Reaction 1	CuO	+	H_2	=	Cu	+	H_2O
Mass (g)	x				y		
Moles	$x/79.5$				$y/63.5$		

Reaction 2	Cu_2O	+	H_2	=	2Cu	+	H_2O
Mass (g)	$1-x$				$0.839-y$		
Moles	$1-x/143$				$(0.839-y)/63.5$		

Mole Ratios

$$\frac{x/79.5}{y/63.5} = 1{:}1$$

$$\frac{(1-x)/143}{(0.839-y)/63.5} = 1{:}2$$

$x = 0.545$ g CuO
$1-x = 0.455$ g Cu_2O

Titration

	$5Fe^{2+}$	+	MnO_4^-	→	Mn^{2+}	+	$5Fe^{3+}$
Volume (mL)	20.8		16				
Molarity	x		0.130				
Moles	$\frac{20.8\,x}{10^3}$		$\frac{16.0\times0.130}{10^3}$				

Mole Ratios

$$\frac{20.8\,x}{16.0\times0.130} = 5{:}1 \Rightarrow x = 0.50\text{ M}$$

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BALANCING REDOX EQUATIONS

Principle: Split reaction into oxidation/reduction half-reactions and balance each in turn with H_2O, H^+, and electrons (e^-) • Recombine half-reaction equations algebraically so as to cancel e's (electrons are conserved in redox reactions)

$$Cr_2O_7^{2-} + HCHO \xrightarrow{H^+} Cr^{3+} + HCOOH$$

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$	$HCHO \rightarrow HCOOH$
$\Downarrow H_2O$	$\Downarrow H_2O$
$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$	$H_2O + HCHO \rightarrow HCOOH$
$\Downarrow H^+$	$\Downarrow H^+$
$14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$	$H_2O + HCHO \rightarrow HCOOH + 2H^+$
$\Downarrow e^-$	$\Downarrow e^-$
$6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$	$H_2O + HCHO \rightarrow HCOOH + 2H^+ + 2e^-$

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

$$3HCHO + Cr_2O_7^{2-} + 8H^+ \rightarrow 3HCOOH + 2Cr^{3+} + 4H_2O$$

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**:

$$aA + bB \leftrightarrows cC + dD \qquad K_c = \left\{\frac{[C]^c[D]^d}{[A]^a[B]^b}\right\}_{eq}$$

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

(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:

$CaCO_3(s) \leftrightarrows CaO(s) + CO_2$

$K_c = [CO_2]$ $\qquad K_p = P_{CO_2}$

(IV) Unknown K's can be calculated from known K's of component reactions:

$2NO + O_2 \leftrightarrows 2NO_2 \quad K_1 = [NO_2]^2/[NO]^2[O_2]$

$2NO_2 \leftrightarrows N_2O_4 \quad K_2 = [N_2O_4]/[NO_2]^2$

$2NO + O_2 \leftrightarrows N_2O_4 \quad K_x = K_1 \cdot K_2$

CALCULATIONS USING K

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

$$K_p = 116 = \frac{X^2}{152 - X} \Rightarrow$$

$$X^2 + 116X - (152)(116) = 0$$

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

$$X = P_{PCl_3} = P_{Cl_2} = 87.3 \text{ kPa}$$

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

PCl_5 (P° = 152kPa) → Reaction → PCl_3, Cl_2, PCl_5

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

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

PHASE EQUILIBRIA

GENERAL CHARACTERISTICS

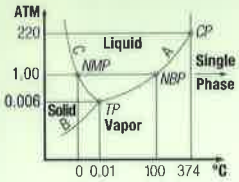
(I) Temperature, pressure, composition, and other properties of the system do not change with time

(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

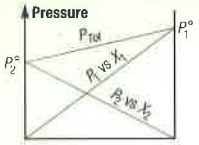
(IV) Le Chatelier's Principle: A system at equilibrium reacts to a change in a variable by minimizing that change

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/NMP...Boiling/melting point at "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_1, in the liquid phase multiplied by the pure component vapor pressure, P_i°: $P_1 = X_1 \bullet P_1^\circ$
$P_2 = X_2 \bullet P_2^\circ$

CALCULATIONS USING RAOULT'S LAW

• Given P_1°, P_2° X_1, at constant T, calculate X_2, P_1, P_2, P_{Tot}, and the vapor phase mole fractions, Y_1 and Y_2

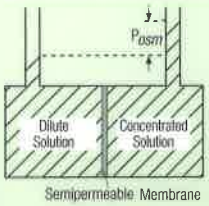
Solve X_2	$X_1 + X_2 = 1 \Rightarrow \quad X_2 = 1 - X_1$
Solve P_1, P_2, P_{Tot}	$\underbrace{P_1 = X_1P_1^\circ \Rightarrow P_2 = X_2P_2^\circ}_{\text{Raoult's Law}} \Rightarrow \underbrace{P_{Tot} = P_1 + P_2}_{\text{Dalton's Law}}$
Apply Ideal Gas Laws	$\underbrace{P_1 = n_1RT/V \Rightarrow P_2 = n_2RT/V \Rightarrow P_{Tot} = n_{Tot}RT/V}_{\text{Ideal Gas Laws}}$
Solve for Y_1, Y_2	$Y_1 = P_1/P_{Tot} \Rightarrow \quad Y_1 + Y_2 = 1 \Rightarrow \quad Y_2 = 1 - Y_1$

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_{osm} 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

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

SPARINGLY SOLUBLE SOLIDS

- K_{sp} = Equilibrium constant of sparingly soluble ionic substances

$$Mg(OH)_2(s) \leftrightarrows Mg^{2+} + 2OH^-$$

$$K_{sp} = [Mg^{2+}][OH^-]^2 = 1.2 \times 10^{-11}$$

CALCULATING ION CONCENTRATIONS

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

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

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

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

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

CALCULATING SOLUBILITY, *S*

- s = Mass of dissolved solids (g/L)

$$\{Mg(OH)_2\}_{dissolved} \rightarrow [Mg^{2+}] = 1.45 \times 10^{-4} \tfrac{mol}{L}$$

$$s = 1.45 \times 10^{-4} \times \text{Molar Mass} = 8.5 \times 10^{-3} \tfrac{g}{L}$$

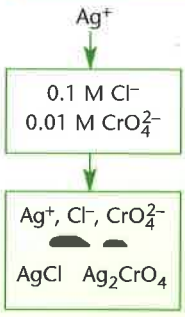
COMMON ION EFFECT

- An increase in the ion concentrations reduces the solubility (Le Chatelier)

Solubility of $Mg(OH)_2$ in 1 M NaOH:

$$[Mg^{2+}][1]^2 = 1.2 \times 10^{-11} \Rightarrow s = 7 \times 10^{-10} \tfrac{g}{L}$$

SELECTIVE PRECIPITATION



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

$$AgCl(s) \leftrightarrows Ag^+ + Cl^-$$

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

$$Ag_2CrO_4(s) \leftrightarrows 2Ag^+ + CrO_4^{2-}$$

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

- [Ag+] required for precipitation to start

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

$$[Ag^+]_{AgCrO_4} = \left\{\frac{K_{sp}}{[CrO_4^{2-}]}\right\}^{1/2} = \left\{\frac{1.9 \times 10^{-12}}{0.01}\right\}^{1/2} = 1.4 \times 10^{-5}$$

- [Cl⁻] remaining when Ag_2CrO_4 starts to precipitate

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

IONIC EQUILIBRIA (ACIDS & BASES)

LOWRY-BRONSTED DEFINITIONS

Bronsted Acid = Proton Donor

$$\underbrace{HNO_3}_{\text{Acid}} \rightarrow \underbrace{H^+}_{\text{Proton}} + NO_3^-$$

Bronsted Base = Proton Acceptor

$$\underbrace{OH^-}_{\text{Base}} + \underbrace{H^+}_{\text{Proton}} \rightarrow H_2O$$

ION PRODUCT OF WATER, K_W

- K_w = Dissociation equilibrium constant of water

$$2H_2O \leftrightarrows H_3O^+ + OH^- \qquad K_w = [H_3O^+][OH^-] = 10^{-14} \text{ (at 25°C)}$$

or

$$H_2O \leftrightarrows H^+ + OH^- \qquad K_w = [H^+][OH^-] = 10^{-14} \text{ (at 25°C)}$$

THE pH SCALE

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

$$pH + pOH = 14 \text{ (at 25°C)}$$

- For pure water at 25°C:

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

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

DISASSOCIATION CONSTANTS OF WEAK ACIDS & BASES

Acid

$$HA \leftrightarrows H^+ + A^- \qquad K_a = \frac{[H^+][A^-]}{[HA]}$$

Base

$$BOH \leftrightarrows B^+ + OH^- \qquad K_b = \frac{[B^+][OH^-]}{[BOH]}$$

Indicators are Weak Acids or Bases

$$\underbrace{HIn}_{\text{color 1}} \leftrightarrows H^+ + \underbrace{In^-}_{\text{color 2}} \quad \text{(weak acid)}$$

$$\underbrace{InOH}_{\text{color 1}} \leftrightarrows OH^- + \underbrace{In^+}_{\text{color 2}} \quad \text{(weak base)}$$

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	$C_2H_5NH_2$
HCOOH	$CO(NH_2)_2$
CH_3COOH	C_6H_5N
Aspirin	Caffeine

CALCULATING ION CONCENTRATIONS WITH K_a

0.6M HF → H_2O → H^+, HF, F^-

	HF	⇆	H^+	+	F^-
Initial:	0.6		0		0
Reaction:	$-X$		$+X$		$+X$
At Eq:	$0.6-X$		$+X$		$+X$

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

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

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

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

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

$$pH = \log\frac{[A^-]}{[HA]} - \log K_a \qquad pOH = \log\frac{[B^+]}{[BOH]} - \log K_b$$

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