

Chemistry !

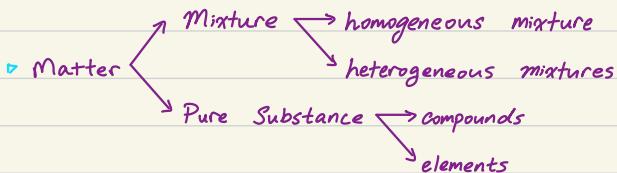
Chapter 1: Introduction

Study of Chemistry

▷ study of matter & change

↳ how energy is involved in change

Classification of Matter



▷ mixture: combination of two or more substances in which substances retain their distinct identities

▷ pure substance: matter has definite or constant composition & distinct properties

▷ a mixture can be separated into pure substances by physical methods

▷ heterogeneous mixture: uniform composition

▷ homogeneous mixture: nonuniform composition

▷ element: substance that cannot be broken down by chemical means

▷ compounds: two or more elements united in fixed proportions

	Mixture	Pure Substance	
		Compound	Element
Observable level (macroscopically)	<ul style="list-style-type: none">• Can be separated by physical means• Composition varies (not constant or definite)	<ul style="list-style-type: none">• Cannot be separated by physical means• Can be broken down into simpler substances with definite proportions• Distinct properties	<ul style="list-style-type: none">• Cannot be broken down further• Distinct properties
Particle level (microscopically)	<ul style="list-style-type: none">• More than one type of particles• No fixed ratio between particles	<ul style="list-style-type: none">• Identical particles (molecules) or• Fixed ratio between smallest particles (atoms)	Identical smallest particles (atoms)

Chapter 2: Atomic Theory & Periodic Table

Subatomic Particles

	mass	charge
▷ electron:	9.1×10^{-31}	-1.6×10^{-19}
▷ proton:	1.67×10^{-24}	1.6×10^{-19}
▷ neutron:	1.67×10^{-24}	0

Atomic Number

mass number $\rightarrow A$  $n \leftarrow$ charge
atomic number $\rightarrow Z$  element symbol



- ▷ mass number: number of protons & neutrons
- ▷ atomic number: number of protons
- ▷ isotope: atoms of same element w/ different number of neutrons in nucleus
- ▷ natural abundance: percent appearance in real life
- ▷ relative atomic mass: \sum mass number * natural abundance

Periodic Table

- ▷ 1A: alkali metal
- ▷ 2A: alkaline earth metal
- ▷ 7A: halogen
- ▷ 8A: noble gas

Metal vs Non-Metal

- | ▷ metal elements | nonmetal elements |
|-----------------------------------|-------------------------------------|
| conductor | insulators |
| electrons move between atoms | electrons only move within atom |
| nucleus attracts electrons weakly | nucleus attracts electrons strongly |

Ions

- ▷ generally between metal & nonmetal
- ▷ electron is "borrowed" from one atom to another
- ▷ cation: positively charged, loose electron
- ▷ anion: negatively charged, gain electron

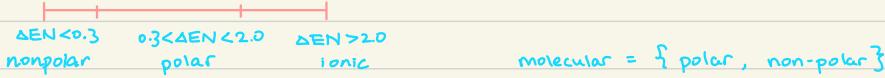
Charges

- ▷ 1A +1 charge, 2A +2 ...
- ▷ 7A -1 charge, 6A -2 ...
- ▷ elements in middle can have multiple charges
- ▷ memorize: Ag^+ , Zn^{2+} , Cd^{2+}



Nonmetal + Nonmetal

- ▷ nonpolar - nuclei of both atoms attract each other's electrons
 - some electrons spend time between atoms
 - no electron transfer
- ▷ polar: - no electron transfer
 - electrons spend time between nuclei, closer to Cl



- ▷ nonmetal - nonmetal definitely molecular
- ▷ metal - nonmetal likely ionic

Crystalline Solids

- ▷ molecular - strong attraction between atoms within molecule
 - weak attraction between molecules
- ▷ ionic - strong attraction between (+) & (-) ions, but not individual molecules

Naming Binary Compounds

▷ rules:

- molecular:
 - prefix { mono, di, tri, tetra, penta, hexa }
 - second element "-ide" ending
- ionic:
 - metal keeps element name
 - nonmetal "-ide" ending
- ionic w/ various charges
 - roman numeral indicates charge of cation
 - nonmetal "-ide" ending



▷ polyatomic ions: cluster of atoms that gains or loses electrons

- ammonium: NH_4^+ +1
- hydroxide: OH^- -1
- nitrate: NO_3^- -1
- carbonate: CO_3^{2-} -2
- sulfate: SO_4^{2-} -2
- phosphate: PO_4^{3-} -3
- others don't need to memorize



Hydrate

- ▷ compound that have specific amount of water molecules crystallized in the lattice
- ▷ number shows ratio between water molecule & 1 formula unit
- ▷ anhydrous: substance w/ no water

Chapter 3: Stoichiometry

General

- ▷ law of definite proportions: a given chemical compound always contains its component elements by a fixed ratio
- ▷ law of mass conservation: in a chemical reaction:
$$\text{total mass of product} = \text{total mass of reactants}$$
- ▷ Avogadro's Law: equal volumes of all gases, at the same temperature and pressure, have the same number of molecules
- ▷ mole:
 - 1 mole = $\frac{1}{12}$ number of carbon atoms in 12 grams of carbon-12 isotope
 - $1 \text{ mole} = 6.022 \times 10^{23} \text{ atoms}$
- ▷ molar mass: mass of 1 mole of any particle in grams ($\times \text{ gram}/1 \text{ mole}$)
 - ↳ weighed average of isotope
- ▷ atomic mass unit: • (amu), formula mass, mass ratio
 - "how many times as massive as hydrogen"

Question Archetypes

- ▷ % composition from formula
- ▷ % composition by molar mass
- ▷ determine empirical formula from % composition \rightarrow assume 100g
- ▷ empirical formula \rightarrow molecular formula

Chemical Reactions

- ▷ macroscopically: the process of forming new substances/products from original substances
- ▷ microscopically: rearrangement of atoms within molecules to form product molecules
- ▷ all atoms in product must come from atom in reactant
 - ↳ reduce to lowest ratio
- ▷ limiting reactant: reactant that runs out first
- ▷ reaction yield: theoretical: amount of product calculated from limiting reactant
actual: actually collected
$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretically}} \times 100\%$$

Chapter 4: Reactions in Aqueous Solutions

General

- ▷ solution: a homogeneous mixture of two or more substances
- ▷ solute: the substance in smaller amount
- ▷ solvent: the substance in larger amount
- ▷ aqueous solution: water as solvent
- ▷ electrolyte: free moving ions
- ▷ nonelectrolyte: free moving neutral molecules
- ▷ total ionic equation: equation broken down into ions
- ▷ net ionic equation: total ionic equation but remove ions not making new substance

Redox Reaction

- ▷ eg $4\text{Fe(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Fe}_2\text{O}_3\text{(s)}$
- Fe become (+) ion
- Fe is oxidized
- Fe is reducing agent
- O gains electrons, becoming (-) ion
- O is reduced
- O is oxidizing agent

Chapter 6 Energy

▷ energy can be transferred through heating (q), working (w), or radiating

Chemical Potential Energy

▷ exothermic reaction: energy leave reaction (eg. $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$ + energy)

▷ endothermic reaction: energy gained in reaction (eg. energy + $\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)$)

▷ internal energy: sum of kinetic & potential energies of all particles in system

$$\hookrightarrow \Delta U = q + w$$

\hookrightarrow heating \hookrightarrow working

▷ for gases: work = - pressure · change in volume w = PΔV

▷ under constant volume: $\Delta U = q_v$

▷ under constant pressure: $\Delta H = \Delta U + P\Delta V = q_p$
or $\Delta U = q_p - P\Delta V$

\hookrightarrow at constant P : energy transferred via heating = enthalpy
enthalpy = $\Delta H_{\text{reaction}} = H_{\text{product}} - H_{\text{reaction}}$

▷ finding ΔH : $\Delta q_{\text{sys}} = -\Delta q_{\text{surr}}$

$$q = S \Delta T \rightarrow \Delta \text{temperature}$$

\hookrightarrow specific heat \hookrightarrow mass

\hookrightarrow amount of energy need to raise 1°C of 1 g of substance

▷ $P\Delta V = \Delta n RT$: change in n can cause PV work to be done

Calorimetry

- use external liquid (water) to test ΔT to find q_{rxn}

- $q_{\text{rxn}} > 0$ endothermic $q_{\text{rxn}} < 0$ exothermic

- $\Delta H_{\text{rxn}} = \frac{q_{\text{rxn}}}{\text{#mole}}$ ← limiting reactant

Enthalpy Change of Reactions

- ▷ ΔH in going from particular set of reactants to products is the same regardless of whether reaction is carried out in one step or in a series of steps
- ▷ ΔH_f° is the change in enthalpy for formation of 1 mole of a compound from its elements
 - ↳ gas 1 atom, 1 M at 1 atm, pure liquid/solid

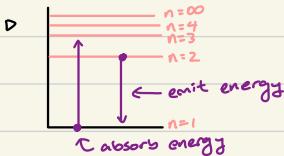
Chapter 7 Electron Structure of Atoms

Waves

- ▷ wavelength = $\frac{\text{velocity}}{\text{frequency}}$ $\lambda = \frac{v}{f}$
- ▷ for electromagnetic waves $\lambda = \frac{\text{speed of light}}{f}$
- ▷ energy can only emitted in quantized form (multiples of elemental quantity)
↳ $E = h\nu$ (energy = planck's constant * frequency)
- ▷ emission of electrons from metal surfaces depends on frequency, not intensity of incident radiation

Electron Energy

- ▷ energy of electron is also quantized \rightarrow only certain energy levels are allowed
- ▷ electron emits energy (photons) when changing energy level



- ▷ electrons behave like standing waves around nucleus $\lambda = \frac{h}{mv}$
- ▷ atomic orbital : approximately where an electron can be found around the nucleus
- ▷ electron density : the probability an electron will be found in a particular region of an atom

Quantum Numbers

- ▷ principle quantum number (n) - describes energy of orbital & average distance of electron from nucleus
 - $n \uparrow$ energy \uparrow distance \uparrow
- ▷ angular momentum quantum number (l) - describes shape of orbitals
 - $0: s: 1, 1: p: 3, 2: d: 5, 3: f: 7$ complex
- ▷ magnetic quantum number: describe orientation of orbitals
- ▷ spin quantum number: describe spin of orbitals

Orbital Energies

- ▷ degenerate: all orbitals with same n have same energy
- ▷ for same n , $s < p < d < f$
- ▷ electron configuration: how electrons are distributed among various atomic orbitals
 - ↳ three rules: (1) electrons fill lowest level first ($1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow \dots$)
 - (2) each orbital can hold two electrons with opposite signs
 - (3) in same level, electrons prefer empty orbitals
- ▷ the spdf orbitals are reflected in periodic table structure
- ▷ noble gas notation can reduce writing (e.g. Mg = $1s^2 2s^2 2p^6 3s^2 = [\text{Ne}] 3s^2$)
- ▷ exception: $[\text{Ar}] 4s^1 3d^5$ not $[\text{Ar}] 4s^2 d^4$
 $[\text{Ar}] 4s^1 3d^{10}$ not $[\text{Ar}] 4s^2 d^9$
- ▷ ions in transition metals have ns removed first (e.g. Mn = $[\text{Ar}] 4s^2 3d^5$ $\text{Mn}^{2+}: [\text{Ar}] 3d^5$)

paramagnetic: attracted by magnetic field, has unpaired electron or transition metal

dimagnetic: repelled by magnetic field, no unpaired electrons

Chapter 8 Atomic Radius & Ionization Energy

Effective Nuclear Charge

- ▷ Z_{eff} is the net attraction on an electron accounting for attraction from nuclear charge and repulsive effects of the other electrons
- ▷ $Z_{\text{eff}} = Z - \sigma$ = nuclear charge attraction - shielding constant
- ▷ Key forces effecting Z_{eff} :
 - (1) attraction from nucleus, more proton more attraction ✗
 - (2) repulsion from core electron to valence electron
 - (3) repulsion between valence electrons
- ▷ Z_{eff} increases in a period, $Z \uparrow$ but σ increases less
- ▷ Z_{eff} increases in a group, but not as much as Z increases
- ▷ atomic radius decreases along period, but increases along group
- ▷ for ions, losing electrons makes atom smaller, gaining electron makes atom bigger
- ▷ isoelectronic: same number of electrons

Ionization Energy

- ▷ minimum energy required to remove an electron from a gaseous atom/ion in its ground state
- ▷ the higher the energy level the electron is at, the lower the IE
- ▷ core electrons have much higher IE than valence electrons
- ▷ generally, IE increases along period, and decreases along group
- ▷ exceptions: $2A > 3A$, $5A > 6A$

Chapter 9 Covalent Bond

Ionic vs Covalent

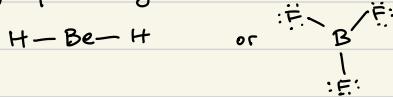
- ▷ ionic bonds "give" electrons while covalent bonds "share" electrons
- ▷ ionic: high EN + low EN covalent: high EN + high EN
- ▷ EN > 2.0 covalent ; 0.3 < EN < 2.0 polar covalent ; EN < 0.3 nonpolar

Lewis Dot Symbol

- ▷ ionic structures have two ions
- ▷ covalent bond has bonding pairs
- ▷ process:
 - (1) count total number of valence electrons
 - (2) put least electronegative (EN) atom in middle (other than H)
 - (3) draw single bond from central atom to other atoms
 - (4) put enough lone pairs on outer atoms to fulfill octet
 - (5) check octet violations and total valence electrons in structure
 - (6) create octet on central atom, add lone pairs or shift in double/triple bonds

▷ ions need to be labeled, e.g. $\text{CO}_3^{2-} \rightarrow \left[\begin{array}{c} :\ddot{\text{O}}: \\ | \\ \text{C} - \ddot{\text{O}}: \\ || \\ :\ddot{\text{O}}: \end{array} \right]^{2-}$

- ▷ charge of ions count as valence electrons
- ▷ exceptions: (1) group 3A hydrides & halides have incomplete octet



(2) NO_2 $:\ddot{\text{O}}=\text{N}=\ddot{\text{O}}:$ odd electron, highly reactive



Chapter 10 Molecular Geometry

Molecular Geometry

- ▷ each lone pair counts as 1 electron domain
- ▷ each single bond, double bond or triple bond counts as 1 domain
- ▷ shapes: name domains draw angles

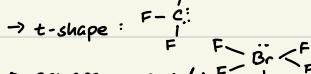
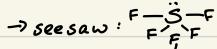
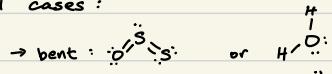
linear	2		180°
trigonal planar	3		120°
tetrahedral	4		109.5°
trigonal bipyramidal	5		120° and 90°
octahedral	6		90°

- ▷ repulsion force: 2 lone pairs \rightarrow lone pair vs bonding \rightarrow bonding pairs

- ▷ Valence Shell Electron pair Repulsion (VSEPR) notation:

$\rightarrow A\text{B}_n\text{E}_m \rightarrow$ n number of lone pair domains
 \underbrace{\hspace{1cm}}_{\text{central atom}} \underbrace{\hspace{1cm}}_m \text{ number of bond domains}

- ▷ Special cases:



Dipole Moments (polarity)

- ▷ ideas:
 - polarity of bond bond can be characterized by vector (dipole moment)
 - two bond dipole vectors have equal length but opposite directions cancel out
 - if all bonds are nonpolar, molecule is nonpolar
 - if shape is highly symmetric, molecule is nonpolar

Valence Bond Theory

- ▷ h-h bond formation
- ▷ covalent bonding is the overlap of atomic orbitals
 - ↳ where bonding electrons appear w/ highest probability
- ▷ σ -bond: head on head overlap : 
- ▷ π -bond: side by side overlap : 
- ▷ Single bonds are σ bonds
 - double → one σ one π
 - triple → one σ two π

Chapter 12 Intermolecular Forces

Three Phases

- ▷ gas : low density, no volume & shape
particles far apart, constant random motion, little attraction
 - ▷ liquid : high density, definite volume, no shape
particles close together, constant random motion, strong attraction
 - ▷ solid : high density, definite volume, no definite shape
particles close together, vibrating in place, orderly arrangement, strongest attraction

Attraction Between Particles

- ▷ **Dipole - Dipole Attraction**: • electrostatic attraction between partially charged ends of polar molecules
 - ▷ **Ion Dipole Attraction**: • dissolving ionic compounds in water
 - ▷ induced dipole: • $\text{(+)} \quad \begin{array}{c} \delta- \\ \uparrow \\ \text{ion induced dipole} \end{array} \quad \text{or} \quad \begin{array}{c} (\delta-\delta+) \\ \uparrow \\ \text{dipole induced dipole} \end{array}$
 - ▷ polarizability: • the ease with which the electron distribution in nonpolar species can be distorted
 - increases with: (1) greater number of electrons
 - (2) more diffuse electron cloud
 - ▷ **Dispersion Force**: • attraction between temporary dipoles
 - exists in all species (neutral, charged, polar, nonpolar)
 - ▷ **Hydrogen Bonding**: • F, O, N very electronegative & small
 - lone pairs on F, O, N attract $\delta+$ H atoms
 - intermolecular attraction between molecules
 - ▷ strengths: dispersion \rightarrow dipole-dipole \rightarrow hydrogen bonding \rightarrow ion dipole \rightarrow ionic bonding
 - ↳ in order of increasing strength

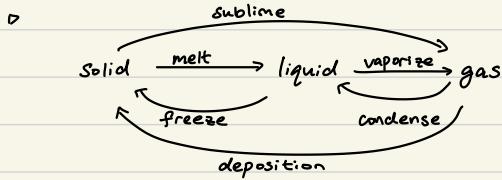
Surface Tension

- ▷ property of liquids
- ▷ imbalance in forces on outer molecules: inward force
- ▷ energy is needed to increase surface area
- ▷ units N/m or J/m²
- ▷ stronger intermolecular attraction → higher surface tension
 - ↳ higher viscosity

Bonding in Solids

	ionic crystals	molecular crystals	covalent crystals	metallic crystals
bond	ionic	intermolecular force	covalent	metallic
physical strength	hard	soft	hard	soft ~ hard
melting point	high	low	high	low ~ high
conductivity	poor	poor	poor	good
liquid conduct	low mp & soft	high mp, doesn't conduct	good conduct	

Phase Changes



- ▷ temperature change: $q = sm\Delta T$
- ▷ phase change: $q = \Delta H m$

Liquid Vapor Equilibrium

- ▷ equilibrium vapor pressure: pressure exerted by vapor at equilibrium
- ▷ rate of evaporation = rate of condensation
- ▷ weaker the intermolecular forces, higher the vapor pressure
- ▷ boiling point: temperature where vapor pressure of liquid = external pressure
- ▷ normal boiling point: boiling point when external pressure = 1 atm
- ▷ $\ln P = -\frac{\Delta H_{vap}}{RT} + C$
 - ↳ ΔH_{vap} : molar heat of vaporization: measure of IMF strength in liquid phase
 - unit: J/mol
- ▷ $\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$

Chapter 15 Chemical Equilibrium

Chemical Equilibrium

▷ equilibrium: the state where two opposite processes become equal

↳ chemical: reactants & products remain constant with time

▷ equilibrium constant: if $jA + kB \rightleftharpoons lC + mD$

$$\text{then } K = \frac{C^l D^m}{A^j B^k}$$

• if reverse: $-K$

• if multiply by n : K^n

▷ if: $K > 1$, more products, eq. rights

$K < 1$, more reactants, eq. left

▷ suppose: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ (all gas)

$$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3} \quad K_p = \frac{P_{NH_3}^2}{P_{N_2} P_{H_2}^3} \quad P \text{ is partial pressure}$$

$$K_c = K_p (RT)^{-\Delta n} \quad \Delta n = \underbrace{(l+m)}_{\text{product}} - \underbrace{(j+k)}_{\text{reactant}}$$

▷ H_2O not included in equilibrium constant expressions

▷ heterogeneous equilibrium: concentrations of pure solids & liquids are not included in expression

▷ let Q be value at moment: $Q = K$ equilibrium
 $Q > K$ make more reactants
 $Q < K$ make more products

▷ ICE diagram:



Initial 1.000 mol/L 2.000 mol/L 0

Change $-x$ $-x$ $+2x$

Equilibrium $1-x$ $2-x$ $2x$

then plug in to $K = \frac{(HI)^2}{(H_2)(I_2)} = \frac{(2x)^2}{(1-x)(2-x)}$ find x

▷ stress: changes in concentration, pressure & temperature can all affect stress

▷ energy + \rightsquigarrow \rightsquigarrow , reduce stress, endothermic

▷ \rightsquigarrow \rightsquigarrow + energy, exothermic

▷ if change in conditions, equilibrium will change to counteract stress

Chapter 16 Acids and Bases

Background

- ▷ Acid
 - substance that dissociates in water to produce hydrogen ions (H^+)
 - H^+ proton donor, H loses electron
- ▷ Base
 - substance that dissociates in water to produce hydroxide (OH^-)
 - H^+ proton acceptor, H gains?
- ▷ conjugate acid-base pair: two substances related to each other by donating & accepting single proton
 - ↳ method: determine reactant side first, taking is based

pH

- ▷ $pH = -\log[H^+]$
 - ↑ unit: mol/L or M
- ▷ $pH > 7$ is base, $pH < 7$ is acidic
- ▷ higher pH \rightarrow lower $[H^+]$
- ▷ $pH = 14 - pOH$
- ▷ strong acid \rightarrow large K_A , weak acid \rightarrow small K_A
- ▷ the weaker the acid, the stronger its conjugate base
dissociated acid at equilibrium
- ▷ percent dissociation: $\frac{\text{dissociated acid at equilibrium}}{\text{initial acid concentration}} \times 100\%$

Chapter 18 Thermodynamics

Energy Distribution in System

- ▷ some molecules have more energy than others
- ▷ equally likely for molecule to gain & lose energy
- ▷ states A & B, W_A & W_B number of ways for system to have state A & B respectively
- ▷ state A \rightarrow state B spontaneously only if $W_B > W_A$
- ▷ **entropy:** $S = k \ln(W)$

$$\Delta S = S_f - S_i = k \ln\left(\frac{W_f}{W_i}\right)$$

k = boltzmann constant

W_i, W_f number of microstates

- $\Delta S > 0$, $W_f > W_i$
- $\Delta S_1 + \Delta S_2 + \dots = \Delta S_{\text{total}}$

Thermal Entropy

- ▷ energy leave: less ways to distribute $\Delta S < 0$
- ▷ energy enter: more ways to distribute $\Delta S > 0$
- ▷ $\Delta S = \frac{q}{T}$

Structure Entropy

- ▷ arrangement of molecules \rightarrow phase change
- ▷ arrangement of atoms \rightarrow chemical reaction
- ▷ $\Delta S_{\text{System}} = \Delta S_{\text{th}} + \Delta S_{\text{st}}$