

CHEM 1412

General Chemistry II

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Ch. 13: Properties of solutions

solution - homogeneous mixture.

solute - minor component (is dissolved).

solvent - major component (retains phase).

Energy changes and solution formation:

ΔH_{soln} - enthalpy change when solution forms

if solvent is H_2O (aqueous), is ΔH_{hydr}

may be exothermic or endothermic

Solution formation is related to *spontaneity* and *disorder*. Exothermic processes with an increase in disorder tend to be spontaneous.

saturated solution - maximum amount of dissolved solute at a given temperature; is in equilibrium with undissolved solute.

unsaturated solution - contains less solute than saturated solution.

supersaturated solution - has more solute dissolved than a saturated solution at a given temperature.

Factors affecting solubility:

1. solute-solvent interaction - "like dissolves like"
2. pressure - solubility of gases

Henry's Law: $S_g = k_g P_g$ **S = solubility**

3. temperature (aqueous)

solids, liquids - usually related to ΔH
gases - inverse relationship

Concentration:

$$\text{mass \%} = \frac{\text{mass solute}}{\text{mass solution}} \times 100\%$$

- ppm, ppb

$$\text{mole fraction} = X_i = \frac{\text{moles}_i}{\text{total moles}}$$

$$\text{molarity} = M = \frac{\text{moles solute}}{\text{L solution}}$$

$$\text{molality} = m = \frac{\text{moles solute}}{\text{kg solvent}}$$

$$\text{normality} = N = \frac{\text{equivalents solute}}{\text{L solution}}$$

colligative properties - solution properties which depend upon only quantity of solute particles present

1. vapor pressure lowering

$$P_A = X_A P_A^\circ \quad P^\circ = \text{VP pure solvent}$$

2. boiling point elevation

$$\Delta T_b = k_b m \quad k_b = \text{BP elevation cst}$$

3. freezing point depression

$$\Delta T_f = k_f m \quad k_f = \text{FP depression cst}$$

4. osmotic pressure

$$\Pi = MRT \quad R = 0.0821 \frac{\text{L} \times \text{atm}}{\text{mole} \times \text{K}}$$

Determination of molar mass...

van't Hoff factor - ions / mole for electrolytes

colloidal dispersion (colloids) - small particles suspended in solvent, not true solution. These will be too small to settle out by gravity - usually 10 - 2000 Å^o in diameter.

- can be solid, liquid, or gas phase
- Tyndall effect - scattering of light

aqueous colloidal dispersions:

- hydrophilic - example: proteins
- hydrophobic - example: oils
(must be stabilized)

removal:

- coagulation (addition of electrolytes)
- semipermeable membranes

Ch. 14: Chemical Kinetics

kinetics - study of reaction rates (speed).

factors which affect rates:

1. state of reactants
2. concentration
3. temperature
4. catalysts - increase rate without undergoing a permanent change; an enzyme is a biological catalyst

$$\text{Rate} = - \frac{\Delta[\text{reactants}]}{\Delta t} \quad t = \text{time}$$

*Reaction rates are experimentally determined.

*rates may be averaged over time

*instantaneous rate is slope of line at tangent to curve at the time of interest

*The rate depends on how the rate expression is defined (stoichiometry).

rate law - mathematical expression relating rate to concentration.

form: $\text{Rate} = k[\text{reactants}]^n$ $n = \text{order}$
 $k = \text{rate cst}$

1st order: $\text{Rate} = k[X]$

2nd order: $\text{Rate} = k[X]^2$ or $\text{Rate} = k[X][Y]$

3rd order: $\text{Rate} = k[X][Y]^2$, etc

Determination of rate laws: experimental!

If rate depends on concentration of 2 or more reactants, use method of initial rates. To do this, vary concentration of only one reactant, while keeping others constant, and measure new rate. The rate change will be due to the effect of that one reactant only. Then, correlate the rate change to the concentration change to determine the order of that reactant in the rate law.

Note: if no effect, may be 0 order.

Rate Laws:

1st order rate law:

$$\text{Rate} = k[A]$$

integrated rate law:

$$\ln [A]_t/[A]_0 = -kt, \text{ or}$$

$$\ln [A]_t = -kt + \ln [A]_0$$

linear plot of $\ln[\text{react}]$ vs time

**at half-life, reactant concentration
is at 50% of initial, so**

$$t_{1/2} = 0.693 / k$$

2nd order rate law:

$$\text{Rate} = k[A]^2 \quad (\text{one reactant only})$$

integrated rate law:

$$1/[A] = kt + 1/[A]_0$$

linear plot of $1/[\text{react}]$ vs time

Collision model for reactions:

activation energy (E_a) = energy which reacting species must have to form transition state or activated complex.

-note: Reactants must have both E_a and correct orientation (steric factor) in order for reaction to occur.

Arrhenius equation:

$$k = Ae^{-E_a/RT} \quad A = \text{frequency factor}$$

reaction mechanism - detailed sequence of reaction steps. Each single step is an elementary step. The slowest step is the rate-determining step.

molecularity - number of molecules reacting in a step. (unimolecular, bimolecular, etc)

intermediate - species formed and used up; will not appear in rate law.

Note: Rate laws are experimentally determined. Mechanisms are theoretical, but must agree with experimental results. Requirements:

- 1. steps must add up to give correct reaction stoichiometry.**
- 2. resultant rate law must agree with experimental rate law.**

catalyst = substance which speeds up rate without undergoing permanent change.

homogeneous catalyst - in same phase.

heterogeneous catalyst - different phase.

enzyme - biological catalyst - "lock-and-key" model.

The substrate, the substance which undergoes reaction, binds at the active site. An inhibitor blocks the active site.

Chapter 15: Chemical Equilibrium

reversible reactions -

equilibrium - state when forward and reverse reaction rates become equal; concentrations will become constant.

Note: This is a dynamic process; the reaction has not stopped.



$$K = \text{equilibrium constant} = \frac{k_f}{k_r} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

if in gas phase, can use P for concentration

$$K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

$$K_p = K(RT)^{\Delta n} \quad n = \text{moles of gas}$$

heterogeneous equilibria - solids and pure liquids do not appear in the equilibrium constant expression. Reaction quotient = Q and is of same form as K , but not equilibrium concentrations.

**Use Q to predict direction of reaction:
If $Q > K$, form more reactants
If $Q < K$, form more products
If $Q = K$, at equilibrium**

Equilibrium problems:

- 1. Determine K from equilibrium conc.**
- 2. Determine equilibrium conc. from K .**

Le Chatelier's Principle: If a system at equilibrium is disturbed, it will react in such a way as to counteract the disturbance and return to equilibrium.

**addition/removal of reactant or product
change in pressure (gas only)
*change in temperature - changes K**

Chapter 16: Acid Base Equilibria

Definitions:

practical:

acid - tastes sour
- turns litmus red

base - tastes bitter
- slippery feel
- turns litmus blue

Arrhenius:

acid - H^+ donor
(note: H^+ = proton; in aqueous solution, will be as H_3O^+ - hydronium ion)

base - OH^- donor

Bronsted-Lowery:

acid - H^+ donor
base - H^+ acceptor

polyprotic acids - more than one H^+

strong acid - completely dissociated

HCl, HBr, HI, HNO₃, H₂SO₄, HClO₄

strong base - completely dissociated

NaOH, KOH

weak acid - incompletely dissociated

(equilibrium - K_a)

HF, HC₂H₃O₂, organic acids (-CO₂H)

weak base - incompletely dissociated

(equilibrium - K_b)

NH₃, organic bases (amines)

conjugate acids and bases - determine

strength of acids and bases (note:

leveling effect)

****water can act as either acid or base**

$$K_w = [\text{H}^+][\text{OH}^-] = \text{autoionization cst}$$
$$= 1 \times 10^{-14} \text{ at } 25^\circ \text{ C}$$

neutral solution: [H⁺] = [OH⁻] = 1 x 10⁻⁷ M

acid solution: [H⁺] > [OH⁻], [H⁺] > 1 x 10⁻⁷ M

base solution: [H⁺] < [OH⁻], [H⁺] < 1 x 10⁻⁷ M

pH Scale: $\text{pH} = -\log [\text{H}^+]$

neutral solution: $[\text{H}^+] = 1 \times 10^{-7} \text{ M}$, $\text{pH} = 7$

acidic solution: $[\text{H}^+] > 1 \times 10^{-7} \text{ M}$, $\text{pH} < 7$

basic solution: $[\text{H}^+] < 1 \times 10^{-7} \text{ M}$, $\text{pH} > 7$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$\text{pH} + \text{pOH} = 14$$

pH of strong acids: complete ionization

$[\text{H}^+] = \text{nominal conc. of acid}$

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

pH of strong bases: complete ionization

$[\text{OH}^-] = \text{nominal conc. of base}$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$\text{pH} = 14 - \text{pOH}$$

pH of weak acids: K_a (solve eq. problem)

% dissociation

if $[HA]/K_a > 10^3$, neglect $-x$ in denom.

$$[H^+] = (K_a M_a)^{1/2}$$

$$pH = -\log [H^+]$$

if not, must solve quadratic equation

Polyprotic acids:

If $K_{a1} / K_{a2} > 1000$, ignore 2nd ionization.

pH of weak bases: K_b (solve eq. problem)

if $[B]/K_b > 10^3$, neglect $-x$ in denom.

$$[OH^-] = (K_b M_b)^{1/2}$$

$$pOH = -\log [OH^-]$$

$$pH = 14 - pOH$$

if not, must solve quadratic equation

Hydrolysis - pH effect of salt on water.

Salt - pH controlled by strength of conjugate species

salt of strong acid + strong base = neutral

salt of strong acid + weak base = acidic

salt of weak acid + strong base = basic

salt of weak acid + weak base = ??

solve pH of salts like that of a weak base or weak acid, using conjugate species of "weak" component

Examples: NH_4Cl → acidic solution

KNO_2 → basic solution

$$** K_w = K_a K_b$$

Lewis definition of acids and bases:

base is e^- pair donor, acid is acceptor

Chapter 17: Additional Aqueous Equilibria

common ion effect - addition of a common ion (one produced by the dissociating species) will reduce the degree of dissociation of a compound.

buffer solution - solution which resists a drastic change in pH upon addition of small amounts of a strong acid or base.

These frequently consist of a weak acid and its salt or a weak base and its salt.

pH of a buffer solution:

a. weak acid and its salt:

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[H^+][\text{salt}]}{[\text{acid}]}$$

$$[H^+] = K_a \times [\text{acid}]/[\text{salt}]$$

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

note: minimal dilution effects

b. weak base and its salt:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]} = \frac{[\text{OH}^-][\text{salt}]}{[\text{base}]}$$

$$[\text{OH}^-] = K_b \times [\text{base}]/[\text{salt}]$$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$\text{pH} = 14 - \text{pOH}$$

titration - addition of a measured amount of a solution of known concentration to an unknown solution to determine its concentration.

equivalence point - occurs when equivalent amounts of reactants have been combined. In the case of an acid/base titration, this will result in the formation of a salt solution.

end point - detectable change to signal the equivalence point. This may be by a color change of an **indicator**, a substance added to the titration to give a color change when the equivalence point is reached.

titration curve - plot of pH (y-axis) vs. volume of titrant (x-axis).

- a. strong acid / strong base
- b. weak acid / strong base (det. of K_a)
- c. strong acid / weak base
- d. polyprotic acids

pH meter - device to directly determine pH of a solution potentiometrically.

Solubility Equilibria:

solubility - amount of a solute which will dissolve in a given volume of solvent at a given temperature.



$$K_{sp} = [M^+][X^-]$$

K_{sp} = **solubility product constant** (p. 1063)

- a. determination of K_{sp} from solubility
- b. determination of solubility from K_{sp}
- c. common ion effect on solubility

Qualitative analysis - identification of ions

Chapter 18: Chemistry of the Environment

atmosphere - gaseous portion of the earth's environment; provides essential gases and temperature regulation; divided into regions based on temperature changes with increasing altitude separated by -pauses (see p. 776).

troposphere - lowermost region of atmosphere; where weather occurs

stratosphere - second region; where ozone layer occurs

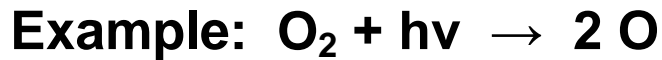
composition - varies with altitude

troposphere
78.08 % N₂
20.95 % O₂
0.93 % Ar
0.04 % CO₂

outer regions - low density, unstable species can exist; high energy solar radiation

reactions in outer atmospheric regions:

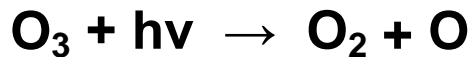
photodissociation - rupture of bond



Photoionization - ionization by photon

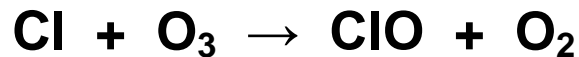
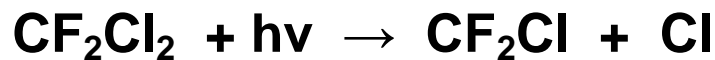


Ozone - main absorber of $\lambda = 240\text{-}310 \text{ nm}$



Ozone depletion - see p. 782-3.

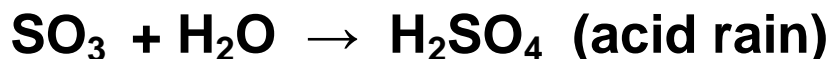
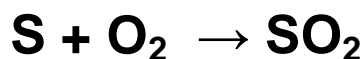
chlorofluorocarbons:



ClO acts as catalyst to further break down ozone by a series of reactions.

Chemistry of the troposphere - air pollution

S oxides - mostly from burning of coal with high S content



CO - blocks O₂ transport by hemoglobin

N oxides - mostly from automobile exhaust - photochemical smog

Complex process involving N oxides, hydrocarbons, and sunlight (see p. 787).

CO₂, H₂O, and global warming - CO₂ and H₂O are "greenhouse gases" which absorb IR radiation (see p. 789).

H₂O - treatment of fresh water (p. 796)

Green Chemistry (p. 798)

Chapter 19: Chemical Thermodynamics

spontaneous process - occurs without outside intervention; goes in the indicated direction. (not related to rate, but may be T dependent)

note: review Ch. 5 - Thermochemistry

1st Law of Thermodynamics - Law of conservation of energy. (E_{univ} is cst)

H = enthalpy (related to heat)

$\Delta H = H_{\text{final}} - H_{\text{initial}}$ (is a state function)

$\Delta H < 0$ = exothermic (usually spont)

$\Delta H > 0$ = endothermic

reversible process - change to system is made in such a way that system can be restored to original state by exactly reversing the change.

Example: system at equilibrium

irreversible process - system must take another path to return to original state.

There is a disorder factor involved in chemical and physical processes.

S = entropy (measure of disorder)

$\Delta S = S_{\text{final}} - S_{\text{initial}}$ (is a state function)

2nd Law of Thermodynamics - in any spontaneous process there is an increase in the entropy of the universe.

$$S_{\text{univ}} = S_{\text{system}} + S_{\text{surroundings}}$$

3rd Law of Thermodynamics - the entropy of a perfect crystal at 0 K is 0.

T affects magnitude of ΔS effect

G = Gibbs Free Energy (also a state function)

$$\Delta G = \Delta H - T\Delta S \text{ and } \Delta G = \Delta G_{\text{final}} - \Delta G_{\text{initial}}$$

**** If $\Delta G < 0$, the process is spontaneous.**

$$\Delta G = -RT \ln K$$

Can calculate ΔH , ΔS , and ΔG from data in tables. (Appendix C, pp. 1100-1102)

Chapter 20: Electrochemistry

electrochemistry - study of relationships between electricity and chemical reactions

oxidation-reduction (redox) reactions:

review: Chapter 4, pp 138-145

oxidation - loss of electrons

reduction - gain of electrons



Balancing redox equations:

Method of half-reactions

1. Identify species being oxidized and reduced and write as half-reactions in the direction given.
2. Balance each half-reaction by first balancing the element undergoing oxidation or reaction, then balance oxygen by adding H_2O , then balance hydrogen by adding H^+ , then balance charge by adding e^- .

3. Multiply each balanced half-reaction by a common denominator to balance e^- loss and gain.
4. Add half-reactions back together to obtain balanced equation. Check for smallest coefficients.
5. If reaction is not in acidic solution, add OH^- to each side to neutralize all H^+ .

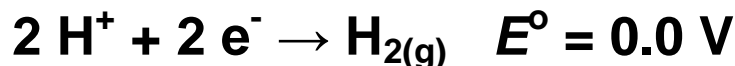
voltaic (galvanic) cell - uses energy from spontaneous redox reaction to generate electricity by transfer of electrons through external pathway

anode - electrode where oxidation occurs
cathode - electrode where reduction occurs

salt bridge - allows for passage of ions to maintain electrical neutrality

cell potential - driving force for electron flow; measured in volts; also called emf or electromotive force

standard reduction (half-cell) potential
 (E°_{red}) - potential of a half-reaction, written as a reduction, under standard conditions ($P = 1 \text{ atm}$, $[X] = 1 \text{ M}$), compared to the standard hydrogen electrode.



These can be found in tables. (p. 1064)

$$E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$$

If $E_{cell} > 0$, this is a spontaneous process
and the cell is a voltaic (galvanic) cell.

$$\Delta G^{\circ} = - nFE^{\circ}$$

n = moles e^- changed
 F = Faraday's cst
= 96,485 C/mole e^-
1 C = 1 amp x 1 sec

note: $\Delta G = - RT \ln K$, so this is one method of determining K

If $E_{cell} < 0$, this is a nonspontaneous process
and the cell is an electrolytic cell.

dependence of cell potential on concentration is given by the Nernst equation:

$$E = E^{\circ} - RT/nF \ln Q$$

At 25° C, this may be simplified to

$$E = E^{\circ} - 0.0592V/n \log Q,$$

so one can calculate E_{cell} if concentrations are known.

At equilibrium, $Q = K$ and $E_{cell} = 0.00 V$,

so this becomes $E^{\circ} = 0.0592V/n \log K$,

$$\text{or } \log K = nE^{\circ} / 0.0592.$$

battery - voltaic cell or series of voltaic cells.

Example: Pb storage battery

fuel cell - voltaic cell with constant potential because reactants are constantly added

Example: $H_2 + O_2$ (space shuttle)

corrosion - naturally occurring oxidation of metals

Example: rusting of iron

electrolysis - forcing a current through a cell to drive a nonspontaneous reaction to occur.

Example: production of Cl₂ and Na

quantitative electrolysis: time x current

$$\begin{aligned} 1 F &= 96,485 \text{ C/mole } e^- \\ &= 96,485 \text{ amp x sec} \end{aligned}$$

Chapter 21: Nuclear Chemistry

Review Ch 2, pp 49 - 50

nucleons - subatomic particles

atomic number = protons in nucleus
(defines identity of element)

mass number = sum of protons and
neutrons for a particular nuclide

isotope = atom with a particular
number of neutrons in nucleus

Example: ^1H , ^2H , ^3H

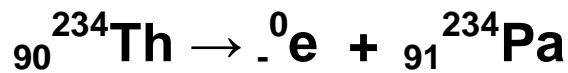
radioactivity - decay of an unstable
nucleus by emission of particles and/or
electromagnetic radiation. A nucleus
which is radioactive is called a
radionuclide or radioisotope).

Nuclear equations:

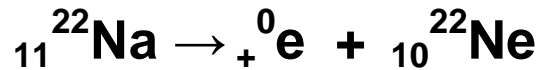
alpha particle (α) - ^4_2He nucleus



beta particle (β) - electron (e^-)

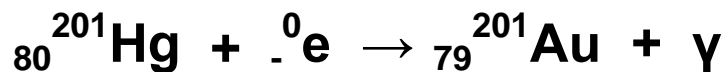


positron (β^+) - positive electron (e^+)



gamma radiation (γ) - high energy electromagnetic radiation (shorter λ than X-Rays)

electron capture - core electron captured by nucleus



nuclear stability - neutron/proton ratio; all nuclei with $Z > 84$ are radioactive.

nuclear transmutation - conversion of one element into another (decay series)

manmade elements - transuranium elements ($Z > 92$)

kinetics - radioactivity is 1st order

$$\ln N/N_0 = -kt$$

half-life - $t_{1/2} = \ln 2 / k = 0.693 / k$

isotopic dating - ${}_6^{14}\text{C}$ ($t_{1/2} = 5730 \text{ yr}$)

detection techniques

1. photographic film
2. Geiger counter
3. scintillation counter

Energy relationships: $E = mc^2$

mass defect - difference in mass between a nucleus and the sum of its nucleons

nuclear binding energy - energy required to decompose a nucleus into its nucleons

fission - splitting of heavy nuclei



nuclear energy - reactor types

fusion - combining of lighter nuclei into heavier nuclei



note: Radioactivity produces ionizing radiation which is harmful to living organisms. However, since radioactive isotopes exist for all elements and many occur naturally, we are constantly exposed to low levels of radioactivity.

effects of radiation - depend upon

- 1. E of radiation**
- 2. penetrating ability**
- 3. ionizing ability**
- 4. chemical properties of element
(biological concentration)**

somatic effects - during lifetime of organism (burns, cancer, etc)

genetic effects - evidenced in later generations (DNA alteration)

medical applications of radioactivity:

- 1. tracers**
- 2. cancer treatment**

Chapter 24: The Chemistry of Life: Organic and Biological Chemistry

organic chemistry - study of C-containing compounds (CO, CO₂, CO₃²⁻, CN⁻ are considered inorganic) – 16 x 10⁶ known

biological chemistry (biochemistry) - study of chemistry of living organisms

hydrocarbons - compounds containing C and H; obtained from petroleum by distillation processes

alkanes - HC compounds containing only C-C bonds (sp³ hybridized); also known as saturated hydrocarbons

CH₄ - methane

C₂H₆ - ethane

C₃H₈ - propane

C₄H₁₀ - butane

C₅H₁₂ - pentane

C₆H₁₄ - hexane

C₇H₁₆ - heptane

C₈H₁₈ - octane

C₉H₂₀ - nonane

C₁₀H₂₂ - decane

isomers - different compounds with same chemical formula

structural isomers - different bonding

optical isomers - mirror images (chiral compounds)

cycloalkanes - C_5H_{10} , C_6H_{12} , etc

naming alkanes - IUPAC Convention

- 1. longest continuous C-C chain is parent compound**
- 2. number C atoms so as to have smallest sum of numbers of side chains, group side chains and use prefix such as di-, tri-, etc (note: assign number to each side chain in the grouping)**
- 3. name side chains (alkyl groups) as to length of chains and specify location**
- 4. list side chains alphabetically**

unsaturated HC's - have one or more C=C or C≡C's.

alkenes - contain one or more C=C, sp²

geometric isomers - different bond angles (cis and trans)

naming alkenes - IUPAC Convention

1. longest continuous C-C chain containing double bond is parent compound
2. number C atoms starting at end closest to double bond, and specify location of side chains
3. name side chains (alkyl groups) as to length of chains, group side chains and use prefix as with alkanes
4. note if cis or trans and number location of double bond

alkynes - contain one or more C≡C, sp

naming - similar to alkenes

aromatics - contain planar ring structures with alternating single and double bonds

example: benzene (C₆H₆)

reactions – additions and substitutions

hydrocarbon derivatives (functional groups) - substitution of C or H by another element (see Table 24.6)

- 1. halogens**
- 2. oxygen**
- 3. nitrogen**

Biochemistry

polymer – large molecule composed of repeating units

Biopolymers:

proteins - chains of amino acids linked by amide bonds (polypeptides)

- 1. primary structure - amino acid sequence**
- 2. secondary structure - coiling into α -helix**
- 3. tertiary structure - folding into final shape**

denaturation - destruction of shape

enzyme function depends on shape (especially about active site)

carbohydrates (CH₂O) – polyhydroxy aldehydes and ketones

monosaccharides - simple sugars

ex: glucose, fructose

chiral molecules (optical isomers)

d- and *l*- isomers

cyclic structures

disaccharides - glycoside linkage

ex: sucrose, lactose

polysaccharides - α- or β- linkages

ex: starch, cellulose, glycogen

lipids - fats (saturated vs unsaturated)

triglycerides - ester of glycerol and long-chain fatty acid

soap - alkali metal salt of fatty acid

micelle

Nucleic acids -

DNA - deoxyribonucleic acid

- genetic information
- protein synthesis

RNA - ribonucleic acid

- transfers info. to ribosomes
- polymers of nucleotides containing
 1. 5-C sugar
 2. N-containing organic base
 3. phosphoric acid

DNA exists in form of double helix

Chapter 23: Transition Metals and Coordination Chemistry

lithosphere – solid portion of earth; man only accesses surface

mineral – naturally occurring solid inorganic compound

- most metals exist in this form
- **ore** – economically useful source of a metal (usually oxides, sulfides, carbonates)
- **gangue** – waste portion of ore

metallurgy – science and technology of extracting and preparing them for use

1. mining
2. concentrating ores
3. reducing metal
4. refining (purifying)
5. **alloying** – metallic material containing 2 or more elements (steel, brass, etc)

Transition Metals: d- and f- orbitals

- size - Lanthanide contraction
- magnetic properties

diamagnetism – all electrons paired

paramagnetism – one or more unpaired e-

ferromagnetism – unpaired e- interact with neighboring atoms

general properties:

- multiple oxidation states
- form colored compounds
- conductivity (electrical and thermal)
- flexibility - malleable and ductile

metallic bonding – delocalization of electrons (electron sea model)

band model (from mo theory)

coordination compounds: contain central metal atom or ion surrounded by ligands

ligands - electron pair donating species

monodentate – one donor atom

bidentate – two donor atoms

polydentate – more than 1 donor atom (form chelate complexes)

note-These may be cationic, anionic, or neutral species

coordination number - number of atoms directly bonded to the metal

geometries – related to coord. number; usually linear, sq. planar, tetrahedral, or octahedral

naming – see p. 1012

isomers – same chemical formula, different arrangement of atoms

structural isomers – different bonds

**linkage isomers – dif. bonding atom
coordination isomers – different cpd**

stereoisomers – same bonds, different spatial arrangement

**geometric isomers – cis and trans
optical isomers - enantiomers**

colors – see p. 1019

biological importance of chelates - ring structures formed by bonding of polydentate ligands to a metal

Ex: myoglobin, hemoglobin, chlorophyll, metalloenzymes

Chapter 22: Chemistry of the Nonmetals

- **descriptive chemistry - systematic study of the elements and their compounds**
- **1st row elements differ in properties due to size; form maximum of 8 electrons on central atom and form multiple bonds better**
- **review Periodic Properties (Ch. 7)**

Relative abundance in earth's crust:

O	49.2%	
Si	25.7%	(silicates)
Al	7.00%	
Fe	4.71%	
Ca	3.39%	
Na	2.63%	
K	2.40%	
Mg	1.93%	

Most abundant elements in living organisms:

C, H, O, N

Hydrogen

- most abundant element in universe
- 3 isotopes (^1H , ^2H , ^3H)

chemistry:

- lose electron (H^+)
- gain electron (H^-)
- share electron (covalent bond)

preparation:

- active metal + strong acid
- methane + steam
- electrolysis of water

uses - Haber process (NH_3)

hydride compounds

Group 8-A: the noble gases

- outer electron config = ns^2np^6 (He = $1s^2$)
- all are monoatomic gases
- Ar, 3rd most abundant atmospheric gas
- all very unreactive; however, some compounds of Xe and Kr have been made with F and O

Group 7-A: the halogens

- outer electron config = ns^2np^5
- all very reactive; found in nature only in reacted form
- elemental form is diatomic - electrolysis
- all of biological importance (except At)

F:

- most electronegative element
- only ox. states are -1 and 0
- very reactive
- F_2 very toxic
- HF

Cl:

- Cl_2 used for sterilizing water
- ox. states -1 - +7
- compounds - oxyacids

Br:

- elemental form is red-brown liquid

I:

- elemental form is purple-black solid

- interhalogen compounds

Group 6-A: the chalcogens

- outer electron config = ns^2np^4

O:

- 2 allotropes - O_2 and O_3 (ozone)



- $O_{2(g)}$ - paramagnetic (2 unpaired e^-)

- obtained from distillation of air

- also from heating of $KClO_3$

- used mostly in steel industry

- #3 industrial chemical

compounds:

- oxides

- peroxides

- superoxides

S:

- several allotropes - S_8

- Frasch process

- oxyacids - H_2SO_4 (ox. state = +6) #1

- H_2S (ox. state = -2)

- biological importance - amino acids
Se: essential trace element

Te: all compounds toxic

Po: radioactive

Group 5-A:

- outer electron config = ns^2np^3
- all except N can have expanded octets (MX_5)

N:

- N_2 - stable gas, triple bond
 - obtained from liquid air
- oxidation states: $-3 \rightarrow +5$
- NH_3 (ammonia) - Haber process (#2)
- N cycle
- N oxides (N_2O , NO , NO_2)
- oxyacids (HNO_2 , HNO_3)

P:

- obtained from phosphate minerals
several allotropes
- oxides (P_2O_3 and P_2O_5)
- oxoacids
- biological import - DNA, ATP, etc

Group 4A:

C - 3 allotropes

Diamond, graphite, fullerenes

inorganic C compounds

- oxides: CO and CO₂
- carbonic acid and carbonates
- cyanides

Si - obtained from SiO₂

- has diamond-type structure
- silicate minerals

Group 3-A

B - boranes

- boric acid