

Chemistry 1411

General Chemistry I

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Lecture Notes

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Ch. 1: Matter and Measurement

chemistry - study of the properties of materials and the changes which they undergo.

matter - physical material of the universe (anything which has mass and occupies space).

elements- simplest forms of matter (100+).

atom - smallest unit into which an element can be divided without losing its identity.

molecule - two or more atoms joined together in a specific shape.

States of matter:

gas - no set shape or volume; expands to fill container.

liquid - set volume, but not set shape.

solid - set shape and volume.

pure substance - matter with distinct properties and uniform composition

element - simplest substance. (Table 1.2)

compound - contains 2 or more elements

Law of constant composition (Law of definite proportions) - The elemental composition of a pure compound is always the same.

mixture - combination of 2 or more substances, each of which retains its own chemical identity.

homogeneous mixture - uniform composition (solution)

heterogeneous mixture - nonuniform composition.

Properties of matter:

physical properties - can be measured without involving chemical reactivity.

Examples: density, color, BP

chemical properties - involve reactivity.

Examples: flammability, toxicity

extensive properties - depend on amount of material present.

Examples: mass, volume

intensive properties - independent of amount of material present.

Examples: temperature, density

Changes in matter:

physical change - does not change chemical identity.

chemical change - changes chemical identity.

separation of mixtures - based on physical properties or chemical properties

physical - distillation or filtration
chemical - reactivity differences

Scientific Method

1. observation and question
2. **hypothesis** (tentative explanation)
3. test hypothesis (experimentation)
4. **theory** (tested hypothesis)
5. **law** (statement of generally expected behavior)

qualitative analysis - determination of identity of a substance or its components.

quantitative analysis - determination of relative amounts of components.

accuracy - agreement of experimental value with true value; measured by **error**.

precision - agreement among a series of experimental values; measured by **deviation**.

Units of Measurement - based on metric system

SI units (Table 1.4)

Prefixes (Table 1.5)

conversion factors:

length: 1 in = 2.54 cm

volume: 1 L = 1.06 qt

mass: 1 kg = 2.205 lb

temperature: K = °C + 273.15

note: temperatures are actually measured using the Celsius scale, which is based on the physical properties of water.

$$\text{°C} = 5/9 (\text{°F} - 32)$$

dimensional analysis - a mathematical problem solving approach using units to cancel each other out (also known as factor unit method).

uncertainty in measurement - experimental measurements will always be limited by the precision of the measuring device.

Significant Figures - all digits of a measured quantity *including the uncertain one*.

1. all nonzero numbers are significant
2. zeros between nonzero digits are always significant
3. zeros at beginning of a number are never significant
4. zeros at end of number *and after a decimal point* are always significant
5. zeros at end of number but before a decimal point are ???

Significant Figures in calculations:

multiplication / division: final answer has same number of sig figs as least number of starting numbers.

addition / subtraction: final answer has same number of decimals as least number in starting numbers.

Chapter 2: Atoms, Molecules, and Ions

The ancient Greeks believed matter to be infinitely divisible. Democritus (~ 400 BC) considered matter to be composed of small, indivisible particles, which he called atoms.

Atomic Theory: John Dalton (~ 1805)

Postulates:

1. Each element is composed of extremely small particles called atoms.
2. All atoms of a given element are identical. Atoms of different elements are different and have different properties.
3. Atoms of an element are not changed into different types of atoms by chemical reactions.
4. Compounds are formed when atoms of more than one element combine. A compound always has the same relative number and kind of atoms.

Discovery of atomic structure - read in book, pp 39-44.

Modern View of the Atom:

atomic mass units (1 amu = 1.66054×10^{-24} g)

**proton - (+ chg) charge = 1.602×10^{-19} C
mass = 1.0073 amu**

**neutron - (0 chg) charge = 0 C
mass = 1.0086 amu**

**electron - (- chg) charge = -1.602×10^{-19} C
mass = 5.486×10^{-4} amu**

nucleus - core of atom; contains protons and neutrons; $d \approx 10^{-14}$ m.

**electrons - located outside of nucleus.
Diameter of atom $\approx 10^{-10}$ m, so the atom is mostly empty space.**

**** In an atom, the number of protons is equal to the number of electrons.**

atomic number - number of protons in a nucleus; defines the element.

mass number - number of protons + neutrons.

isotopes - atoms of an element with differing numbers of neutrons. Some isotopes are unstable (radioactive).

atomic mass - avg. mass of element, considering isotopes and natural abundance.

Periodic Table

molecule - two or more atoms bound tightly together.

chemical formula - tells types and relative number of atoms in a molecule.

structural formula - shows bonding.

other representations: perspective (shows geometry), ball and stick, space-filling, etc.

ion - charged particle (protons \neq electrons)

anion - neg. ion (protons $<$ electrons)

cation - pos. ion (protons $>$ electrons)

ionic compound - formed by combination of cations and anions (+ and - charged ions), with relative numbers so that charge is zero.

Naming : cation before anion

cations:

1. **monoatomic** - (usually metals) take name of parent element
2. If parent atom can have more than one charge, put charge in Roman numerals in parenthesis.
3. **polyatomic** cations - have names ending in *-ium*.

anions:

1. **monoatomic** - (usually nonmetals) take name of parent element, drop ending, and add *-ide*.
2. **polyatomic** anions - may have names ending in *-ide*. Oxy anions may have names beginning with *hypo-* or *per-* and ending with *-ite* or *-ate*.
3. **hydrogen** added to anions -

Naming:

1. acids

2. molecular compounds

3. organic compounds

alkanes -

alkenes - multiple bonds

functional groups

Chapter 3: Stoichiometry

stoichiometry - study of quantitative relationships between chemical formulas and chemical equations.

Law of Conservation of Mass - total mass of substances present after a chemical reaction is the same as present before the reaction.

Chemical Equations - show products and reactants (states)

Reactants → Products

*balance by using **coefficients**

Types of reactions:

combination reactions -

2 reactants → 1 product

decomposition reactions -

1 reactant → 2 products

combustion reactions - (HC's in O₂)

atomic mass scale - based on $^{12}\text{C} = 12.0000$

average atomic mass = atomic weight

formula weight = sum of atomic weights in a formula unit. If formula unit is a molecule, this is a molecular weight.

% composition from formula weights -

mole - Avogadro's number (6.02×10^{23}) of formula units of a substance.

molar mass - formula weight expressed in grams.

empirical formula - simplest ratio of types of atoms in a molecule.

determination from % composition

quantitative aspects

theoretical yield

limiting reagent

% yield = (actual yield / theo. yield) x 100%

Chapter 4: Aqueous Reactions and Solution Stoichiometry

solution - homogeneous mixture

solute - minor component (dissolves)

solvent - major component (keeps phase)

aqueous solvent - water is solvent

ionic compounds - dissociate into ions in solution. These solutions will conduct electricity, so the compounds are also called **electrolytes**.

strong electrolytes - completely ionize

weak electrolytes - incompletely ionize; equilibrium between ions and parent

nonelectrolyte - does not ionize (molecular compound)

precipitation reaction - forms insoluble substance (precipitate)

solubility rules - Table 4.1

exchange (metathesis) reactions - anions and cations exchange partners.

molecular equation - shows reactants and products as whole species.

complete ionic equation - shows reactants and products as ions.

spectator ions - do not participate.

net ionic equation - ionic equation without spectator ions.

acid-base (neutralization) reactions:

acid - donates H^+ in aqueous solution

strong and weak acids

base - accepts H^+ in aqueous solution

strong and weak bases

acid + base \rightarrow salt + H_2O

Neutralization reactions:

net ionic equation is usually



but if base is not a hydroxide, may have gas formation (H_2S , CO_2 , etc)

Oxidation-reduction reactions (redox):

oxidation - loss of electrons by a substance

reduction - gain of electrons by a substance

oxidation numbers - electron tracking system (hypothetical charge)

1. ox. no. = 0 for atom in elemental form.
2. for monoatomic ion, ox. no. = charge.
3. metals are usually positive in compounds, and nonmetals are usually negative.
 - a. O is usually -2 in compounds (except in peroxides (O_2^{2-})).
 - b. H is +1 when with nonmetals and -1 with metals.
 - c. F is -1 in all compounds.
 - d. sum of ox. no's = chg on species.

oxidation of metals - displacement reactions

activity series - list of metals arranged in order of decreasing ease of oxidation. (Any metal on list will be oxidized by ions of metals below it).

Concentration of solutions:

$$\text{Molarity} = M = \frac{\text{moles solute}}{\text{solution volume in L}}$$

$$\text{moles solute} = M \times \text{volume (L)}$$

Dilution:

$$M_{\text{con}} V_{\text{con}} = M_{\text{dil}} V_{\text{dil}}$$

Solution stoichiometry:

titration - reaction of a measured quantity of an unknown with a measured amount of a solution of known concentration (standard solution). The equivalence point is when stoichiometric amounts of reagents have been combined. The end point is a detectable change to signal the equivalence point, such as a color change of an indicator.

Chapter 5: Thermochemistry

thermodynamics - study of energy and its transformations from one form to another.

thermochemistry - study of energy changes involved in chemical reactions.

kinetic energy - energy which a body has by virtue of being in motion.

$$E_k = 1/2 mv^2$$

potential energy - stored energy by virtue of position or composition.

energy units:

joule (J) is official S.I. unit

$$1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2 / \text{sec}^2$$

calorie (cal) = amount of heat required to raise T of 1 g H₂O by 1° C.

$$1 \text{ cal} = 4.184 \text{ J}$$

note: 1,000 cal = 1 kcal = 1 Calorie = 4.184 kJ

system - that portion of the universe under study.

surroundings - the rest of the universe.

closed system - can exchange energy, but not matter, with its surroundings.

energy - capacity to do work or transfer heat.

force – any push or pull exerted on an object.

work - product of a force and the distance over which it is applied.

$$w = F \times d$$

heat - flow of energy from a hotter body to a colder one.

First Law of Thermodynamics - Energy is neither created nor destroyed in chemical reactions; it is converted from one form to another. (Law of Conservation of Energy).

total energy = internal energy = $E_{\text{pot}} + E_{\text{kin}}$.

- cannot measure absolutely, but can measure changes

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

**- therefore, $\Delta E > 0$ = energy gain, and
 $\Delta E < 0$ = energy loss.**

$E = q + w$ (q = heat change, w = work done)

endothermic reaction - heat is absorbed.

exothermic reaction - heat is given off.

state function - property which depends only on final and initial states, is independent of path.

enthalpy (H) - state function involving heat changes in chemical reactions at constant pressure.

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

**note: $\Delta H < 0$ = exothermic reaction;
 $\Delta H > 0$ = endothermic reaction.**

calorimeter - apparatus used to measure heat flow. (Calorimetry is the measurement of heat flow)

heat capacity (C) - amount of heat required to raise T of an object 1° C.

molar heat capacity - amount of heat required to raise T of one mole of a substance 1° C.

specific heat - amount of heat required to raise T of one gram of a substance 1° C.

$$\text{sp heat} = \frac{\text{heat transferred}}{\text{mass} \times \Delta T} = \frac{q}{m \times \Delta T}$$

constant pressure calorimetry and constant volume calorimetry

$$q_{\text{lost}} = q_{\text{gained}}$$

Hess's Law - if a reaction is carried out in a series of steps, ΔH for the reaction will be equal to the sum of ΔH 's for each step.

If one knows the H of reactants and products, one can calculate $\Delta H_{\text{reaction}}$

ΔH_f° is known for many compounds and can be looked up in tables. (pp. 1059-1061)

note: ΔH_f° is the amount of energy (in kJ) required to form 1 mole of a substance from its elements in their standard states under standard conditions ($T = 25^\circ\text{C}$, $P = 1 \text{ atm}$).

Ch. 6: Electronic Structure of Atoms

Electromagnetic radiation - type of energy which has both electrical and magnetic components (ex - heat and light).

- requires no supporting medium; is transmitted readily through a vacuum
- frequently described in terms of a wave

ν = frequency (cycles per second)

λ = wavelength (distance between two equivalent points)

$$c = \nu\lambda \quad c = 3.00 \times 10^8 \text{ m/sec}$$

electromagnetic spectrum - large range of λ ; divisions based on methods required to generate radiation.

photoelectric effect – non-additive property of energy below a threshold level. This was explained by the German physicist Max Planck, who assumed that the energy was emitted in discrete amounts of energy, called quanta.

$$E = h\nu = hc/\lambda \quad h = \text{Planck's constant} \\ = 6.63 \times 10^{-34} \text{ J-sec}$$

Einstein used Planck's quantum theory to explain the photoelectric effect by considering radiant energy as a stream of packets of energy, which he called photons.

spectra - continuous vs line spectra.

Gas phase atoms emit line spectra. This led to the development of the Bohr theory of the atom. Bohr postulated that the electrons in an atom existed only in certain allowed orbits (n), and when an atom passed from one orbit to another, energy changed.

$$E_n = (-2.18 \times 10^{-18} \text{ J})(1 / n^2) \text{ for H}$$

ground state - lowest energy (n = 1).

excited state - higher energy level (n > 1).

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

Louis de Broglie - wave nature of matter.

Werner Heisenberg - uncertainty principle - one cannot simultaneously know the position and momentum of an electron.

Erwin Schrodinger - wave model of atom.

wave mechanics (quantum mechanics) - describes electron in atom in terms of a wave function (Ψ), which is a function of the energy and position of the electron.

Ψ^2 = probability of locating electron in a given position.

orbital - solution to wave function.

Solution of wave function requires three quantum numbers:

n = principal quantum number

l = azimuthal quantum number

m_l = magnetic quantum number

electron shell - all electrons with same n .

subshell - all electrons with same n and l .

An orbital is represented by the three quantum numbers:

n represents the distance of the electron from the nucleus.

$$n = 1, 2, 3, \dots$$

l indicates the shape of the orbital.

$$l = 0, 1, 2, \dots (n - 1)$$

$l = 0$ is an s orbital

$l = 1$ is a p orbital

$l = 2$ is a d orbital

$l = 3$ is an f orbital

m_l indicates the orientation of the orbital.

$$m_l = l, l - 1, l - 2, \dots 0 \dots - (l - 1)$$

$$l = 0, m_l = 0$$

$$l = 1, m_l = 1, 0, -1$$

$$l = 2, m_l = 2, 1, 0, -1, -2$$

$$l = 3, m_l = 3, 2, 1, 0, -1, -2, -3$$

electrons also have a magnetic moment, so

$m_s = \text{spin quantum number} = \pm 1/2$.

Pauli exclusion principle - no two electrons in an atom have all 4 quantum numbers the same. (max of 2 electrons / orbital)

orbital energy levels - increase with n and l

electron configuration - arrangement of electrons in orbitals in an atom.

Hund's rule - electrons in an atom will fill orbitals in order of increasing energy. When **degenerate orbitals** (those with the same energy) are filled, each will first fill singly, then will be filled by an electron with opposite spin.

core electrons - electrons in filled inner shells; are not involved in reactivity.

valence electrons - outer electrons involved in chemical reactions.

condensed electron configurations

Periodic table - elements are arranged in order of increasing atomic number. Elements in columns have same outer electron configurations.

representative elements - s and p orbitals

transition elements - d orbitals

f-block elements

Ch 7: Periodic Properties of the Elements

development of periodic table:

ancient times to mid-1800's

Mendeleev (Russia) and Meyer
(Germany) - properties vs atomic wts

Mosely (1913) - atomic number

G.N. Lewis - electron configuration

effective nuclear charge - shielding

periodic properties - vary in a predictable
manner

atomic radius - avg. distance from nucleus to
outermost electrons.

down group - increases (larger n)

across period - decreases (larger
effective nuclear charge)

ionic sizes:

cations - smaller than parent atom

anions - larger than parent atom

ionization energy - energy required to remove an electron from a gaseous atom or ion

down group - decreases (larger r and smaller effective nuclear charge)

across period - increases (smaller r and larger effective nuclear charge)

increases as successive electrons are removed

electron affinity - energy released when an electron is added to a gaseous atom or ion.

across period - increases (larger effective nuclear charge)

down group - not much change

metals and nonmetals

metals

form cations

have luster

malleable & ductile

conduct heat & elect

form basic oxides

nonmetals

anions

dull

brittle

insulators

acidic oxides

metalloids - intermediate (semiconductors)

Active Metals - group trends

Group 1A: Alkali Metals

- outer electron config. ns^1
- form +1 ions only
- soft metallic solids (occur as ions)
- biological importance
- can form O_2^{2-} and O_2^- (K,Rb,Cs)

Group 2A: Alkaline Earth Metals

- outer electron config. ns^2
- form +2 ions only
- biological importance
- Ca^{2+} and Mg^{2+} in hard water

Nonmetals - group trends

Hydrogen

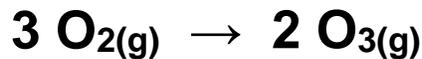
- H_2 gas (not in atmosphere)
- most abundant element in universe
- higher ionization energy than I-A's
- good fuel
- chemistry:
 - lose electron (H^+)
 - gain electron (H^-)

Group 6-A: the Oxygen Group

- outer electron config = ns^2np^4

O:

- elemental forms - O_2 and O_3 (ozone)



- strong oxidizing agent

S:

- several elemental forms - S_8

- good oxidizing agent

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

- H_2S (ox. state = -2)

- SO_2 pollutant

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

- all of biological importance (except At)

F:

- most reactive halogen
- only ox. states are -1 and 0
- F₂ very toxic (pale yellow gas)
- HF

Cl:

- Cl₂ used for sterilizing water
- ox. states -1 - +7
- compounds

Br:

- elemental form is red-brown liquid

I:

- elemental form is purple-black solid

Group 8-A: the noble gases

- outer electron config = ns^2np^6 (not He)
- all are monoatomic gases
- all very unreactive; however, some compounds of Xe and Kr have been made with F and O

He:

- 2nd most abundant element in universe (fusion in stars)

Ch. 8: Chemical Bonding - Basic Concepts

ionic bond - electrostatic attraction between oppositely charged ions.

covalent bond - electrons are shared between two atoms.

metallic bond - electrons are shared among many atoms.

electronegativity - ability of an atom in a molecule to attract electrons to itself.

- This is a **periodic property** which decreases down a group (smaller effective nuclear charge) and increases across a period. The most electronegative element is F, and O is second.

ionic bond - formed between two elements with large electronegativity difference, such as a metal and a nonmetal.

covalent bond - formed between two elements with small electronegativity difference, such as two nonmetals. If moderate difference, this is a **polar bond**, or a **polar covalent bond**.

Lewis dot structures - symbol of element showing valence electrons as dots.

ionic: NaCl, CaCl₂

note: no electron sharing

covalent:

pure covalent: H₂, F₂

polar covalent: H₂O, NH₄⁺

Drawing covalent Lewis structures:

- 1. add up electrons (check charges)**
- 2. show bonded atoms as - (- is one e⁻ pair)**
- 3. all atoms except H should have 8 e⁻ 's
(Lewis Octet Rule)**
- 4. use multiple bonds if necessary**
single bond - one e⁻ pair shared
double bond - two e⁻ pairs shared
triple bond - three e⁻ pairs shared

exceptions to Lewis octet rule

- 1. odd number of electrons**
- 2. less than 8 electrons**
- 3. more than 8 electrons**

polar covalent bond - covalent bond between atoms with difference in electronegativity.

Examples: Cl_2 , CCl_4 , KCl

If electrical charges of equal magnitude, but different sign, are separated by a distance, a dipole is established, which is measured by a dipole moment ($\mu = Qr$).

$Q = chg$, r is separation distance

If centers of positive and negative charge do not coincide, this is a polar molecule.

Examples: CCl_4 (nonpolar), CHCl_3 , H_2O

formal charge - electron bookkeeping in Lewis structures.

- 1. all nonbonding electrons are assigned to the atom on which they occur**
- 2. bonding electrons are split equally**
- 3. sum of formal charges adds up to charge on species**

If more than one Lewis structure is possible, the most likely will be the one with the lowest formal charges (and - charges on most electronegative atoms).

resonance structures - equivalent Lewis structures

Examples: CO_3^{2-} , C_6H_6 [a \leftrightarrow b]

Energetics of Bond Formation:

Ionic Bonds: electrostatic attraction between oppositely charged ions (no e^- sharing). In the solid state, the ions are held together in a geometric array called a crystal lattice. The lattice formation is an exothermic process.

lattice energy - energy required to separate one mole of an ionic solid into its gaseous ions. This will increase as ionic charges increase and ionic sizes decrease.

covalent bonds - one or more electron pairs shared. Bond length - distance between bonded atoms. Bond length decreases with number of shared pairs.

single > double > triple bond

bond strength - energy required to break a bond. Measured as bond enthalpy - ΔH to convert one mole of a substance into gaseous atoms. (shorter bond = stronger)

Chapter 9: Molecular Geometry and Bonding Theories

molecular shapes - geometric form of molecule (linear, tetrahedral, etc)

bond angle - angle formed by lines joining bonded atoms in a molecule. This will dictate the molecular shape.

Basic molecular shapes:

AB₂ - linear or bent

AB₃ - trigonal planar, trigonal pyramidal, or T-shaped

AB₄ - tetrahedral or square planar

AB₅ - trigonal bipyramid or sq. pyramid

AB₆ - octahedral

Valence Shell Electron Pair (VSEPR) Theory

This is a model used to predict molecular shapes based upon the premise that electron pairs will orient as far as spatially possible from each other. All electron pairs, both bonding and nonbonding, will be assigned to an electron domain, and the number of electron domains will dictate the bond angles, and thus the molecular shape. (see Tables 9.2 and 9.3)

Procedure to predict geometry:

1. Draw Lewis structure.
2. Determine *electron-domain geometry*.

Each nonbonded pair and each bond (single, double, or triple) counts as an electron domain. Arrange for maximum repulsion.

3. Determine *molecular geometry*.

Only consider arrangement of atoms.

polarity of molecules - depends on molecular geometry and polarity of bonds. In a **polar molecule** the centers of positive and negative charge do not coincide.

Valence Bond Theory

A chemical bond is formed when atomic orbitals of two atoms **overlap** (occupy same region of space).

σ (sigma) bond - two orbitals overlap along line joining the bonded atoms

π (pi) bond - two orbitals overlap at 90° to line joining the bonded atoms

hybridization - used to explain equality of bonding. Use correct number of unpaired electrons, then combine orbitals as needed to obtain needed number of bonds (average energies) Geometries will be the same as would be predicted from VSEPR Theory.

Examples: CH_4 , BCl_3 , BeF_2

multiple bonds: double bond is sp^2
triple bond is sp

Molecular Orbital Theory

Atomic orbitals combine to form molecular orbitals, which exist in quantized energy states. Molecular orbitals have many of the same characteristics as atomic orbitals, such as a maximum of 2 paired electrons per orbital.

To form molecular orbitals, combine the atomic orbitals. The average energy of the mo's must be the same as that of the parent atomic orbitals.

bonding mo - lower energy than parents
antibonding mo - higher energy than parents

rules for molecular orbital diagrams:

1. number of mo's formed is same as number of ao's input
2. atomic orbitals of similar energy combine most effectively
3. effectiveness of combination is proportional to overlap
4. maximum electrons / mo = 2, with paired spins
5. mo's of same energy will fill singly with one spin before pairing occurs.

bond order = 1/2 (bonding e⁻ - antibonding e⁻)

Examples: H₂ and He₂

Li₂ and Be₂

B₂ - Be₂

paramagnetism - one or more unpaired e⁻.

diamagnetism - all e⁻ paired.

heteronuclear diatomic molecules -

Chapter 10: Gases

characteristics of gases:

- expand to fill the container
- are compressible (only about 0.1% of volume is actually matter)
- common gases contain only nonmetals (atmosphere)
- form homogeneous solutions

vapor - substance usually in solid or liquid phase, but with low boiling point

pressure (P) - force/area

- atmospheric pressure (pressure of a column of atmosphere 1 m² in cross-section and extending to top of atmosphere = 1.01×10^5 N)
- SI unit is pascal (Pa); $1 \text{ Pa} = 1 \text{ N} / \text{m}^2$. Therefore, $P_{\text{atm}} = 1 \times 10^5 \text{ Pa} = 100 \text{ kPa}$.
- A common unit is the bar = 10^5 Pa

barometer - device used to measure P

- $P_{\text{std atm}} = 76 \text{ cm Hg}$
 $= 760 \text{ mm Hg} = 760 \text{ torr}$

manometer - measures P of closed systems

A gas is characterized by its V, T, P, and n (number of moles).

Gas Laws:

Boyle's Law: $PV = \text{constant}$, or $V = \text{cst} / P$

Charles's Law: $V = \text{cst} \times T$, or $V / T = \text{cst}$
Note: T is in K ($^{\circ}\text{C} + 273.15$)

Avogadro's Law: $V = n \times \text{cst}$

Ideal Gas Equation: $PV = nRT$
($R = 0.0821 \text{ L-atm} / \text{mole-K}$)

note: At ideal conditions ($P = 1 \text{ atm}$, $T = 0^{\circ}\text{C}$), 1 mole of gas will occupy 22.4L). This is called standard molar volume.

Calculations:

1. using ideal gas equation
2. change one or two conditions
3. gas densities and molar mass

Dalton's Law - $P_{\text{total}} = P_A + P_B + P_C + \dots$

$X_i = \text{mole fraction of gas}_i = n_i / n_{\text{total}}$

$P_i = X_i P_{\text{total}}$

Collecting gases under water:

$P_{\text{total}} = P_{\text{H}_2\text{O}} + P_{\text{gas}}$

Kinetic Molecular Theory:

1. gases consist of particles in constant random motion
2. volume of particles is negligible
3. attractive and repulsive forces between particles are negligible
4. collisions are elastic
5. K.E. is proportional to absolute temperature

diffusion - spread of a substance through space

effusion - escape of gas from a container

Graham's Law: $\frac{r_1}{r_2} = \left(\frac{MM_2}{MM_1} \right)^{1/2}$

ideal gas - obeys gas laws perfectly

non-ideal behavior: van der Waals eq:

$$(P + (n^2a / V^2))(V - nb) = nRT$$

- look up a and b in tables for specific gases

Chapter 11: Intermolecular Forces, Liquids, and Solids

Physical properties of a substance depend upon its state (solid, liquid, or gas).

Liquids and solids are often referred to as condensed phases.

Substances which are liquids at room temperature are almost all molecular in nature.

intermolecular forces - attractive forces between molecules.

The state of a substance at a given temperature is a result of the relative strength of the intermolecular forces compared to the kinetic energy.

Note: Intermolecular forces are much weaker than chemical bonds. Strength of these forces is indicated by MP and BP.

Intermolecular forces are electrostatic in nature.

Types of intermolecular forces:

1. ion-dipole forces - attractions between ion and end of a polar molecule
 - important in solution process
2. dipole-dipole forces - attraction between ends of polar molecules (weaker than ion-dipole forces)

hydrogen bonding - special type of dipole-dipole attraction between H bonded to an electronegative atom and an unshared electron pair on a nearby electronegative atom; can be inter- or intramolecular.

3. London dispersion forces - temporary induced dipoles; increase with size.

polarizability - ease with which the charge distribution about a molecule can be distorted.

- Comparison of intermolecular forces:

- Properties of liquids:

viscosity - resistance of a liquid to flow

surface tension - energy necessary to increase surface area of a liquid

cohesive forces - between like molecules

adhesive forces - between liquid and a surface, as in a meniscus or capillary action

phase changes -

melting-freezing (liquid ↔ solid)

vaporization-condensation (liquid ↔ gas)

sublimation-deposition (gas ↔ solid)

energy changes - ΔH_{fusion} and ΔH_{vap}

Heating curves:

evaporation - escape of molecules from the surface of a liquid (endothermic process)

volatile - substance which evaporates easily

vapor pressure - partial pressure of a vapor above a liquid at a certain temperature

boiling - when vapor pressure within a liquid equals external pressure

normal boiling point - boiling point of a liquid at 1 atm pressure.

critical temperature - highest temperature at which a substance can exist as a liquid. The pressure required for a liquid at this temperature is the **critical pressure**.

Phase diagrams:

Ch. 13: Solids

crystalline solid - well-defined, ordered, repeating arrangement

amorphous solid - no orderly structure

crystal lattice - three-dimensional representation of a crystalline solid

unit cell - smallest unit of a crystalline solid which can represent its structure. This is usually 6-sided and is described by edge lengths and angles.

Ex: cubic unit cell

close packing of spheres -

hexagonal and cubic

types of crystalline solids -

- 1. molecular** - atoms or molecules held together by dipole-dipole, London forces, or H-bonding. Ex: sugar
- 2. covalent** - atoms joined by covalent bonds. Ex: diamond and graphite
- 3. ionic** - ions with ionic bonding.
Ex: NaCl
- 4. metallic** - atoms held together by metallic bonds (delocalized electrons).
Ex: typical metals

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 H₂O (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 gasses

Henry's Law: $S_g = k_g P_g$ S = solub.

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

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

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

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

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

