KEY IDEAS:
- Rate Law
- Catalysts
- Reaction Mechanism

Chapter 13

$$\text{Avg rate} = \frac{-\Delta[A]}{\Delta t}$$

Bimolecular rxn:

$OH^- + H_3O^+ \rightarrow H_2O + H_2O$

Intermediates:

Things that cancel, don't make it to final rxn

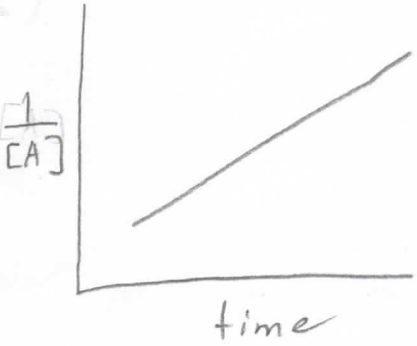
Unimolecular rxn:

$N_2O_4 \longrightarrow NO_2 + NO_2$

0 order

1^{st} order

2^{nd} order



$$t_{1/2} = \frac{[A]}{2k} \qquad t_{1/2} = \frac{\ln(2)}{k} \qquad t_{1/2} = \frac{1}{k[A]}$$

Arrhenius Equation

$$k = Ae^{-Ea/RT}$$

$$\ln k = \left(\frac{-Ea}{R}\right)\left(\frac{1}{T}\right) + \ln A$$

Mechanism

$N_2O_5 \rightleftharpoons NO_2 + NO_3$

$NO_2 + NO_3 \rightarrow NO + NO_2 + O_2$

$NO + NO_3 \rightarrow 2NO_2$

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

$K_p = K_c(RT)^{\Delta n}$ where Δn = (product mol - reactant mol)

Chapter 8

KEY IDEAS:
- Intermolecular Forces
- Solid state

M[illegible]ting and boiling points:

When the kinetic energy of moving particles exceeds the intermolecular forces keeping it as a solid, it phase changes into a gas

∴ Substances with strong intermolecular forces are hard to turn into gases

STRONGEST intermolecular Force:

ion-dipole: ionic substance dissolved in water

- Cl^- in H_2O
- Na^+ in H_2O

Fluorine: lowest boiling point, Weak intermolecular forces

Dispersion Forces:

Weak, based on e^- cloud shape. ↑ disperison Force when ↑e^- and ↑Polarizability

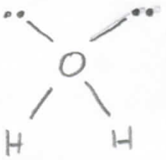
Dipole moment: When two atoms of different electronegativities share a bond

H_2: $\underbrace{H - H}_{\text{(No dipole moment)}}$ (Non-polar)

HCl: $\underbrace{\overset{2.1}{H} \rightarrow \overset{3}{Cl}}_{(0.9)}$ (dipole moment)

C[illegible]: $\underset{\leftarrow\!+ \quad +\!\rightarrow}{O=C=O}$ (Non-polar)

H_2O:



(Non-Polar)

Solids

- **Crystalline:** highly regular
- **Amorphous:** irregular

Sodium chloride structure:

$2(Na^+ radius) + 2(Cl^- radius)$ = Edge Length

Forces in Solids:

① $\boxed{\text{Molecular}}$ Low Melting Point
- dispersion [Ar], [HCl], [H_2O]
- dipole
- hydrogen bonding

② $\boxed{\text{Network}}$ HIGHEST melting point
- Covalent bonding [SiO_2, TiO_2, Al_2O_3]

③ $\boxed{\text{Metallic}}$
- range of melting points [Cu]
- Delocalized bonding

④ $\boxed{\text{Ionic}}$
- electrical bonding

[NaCl, NaOH, $NaClO_3$, KNO_3, $CuSO_4$]

4 Basic Lattice Structures (Unit cell):

① Simple Cubic:

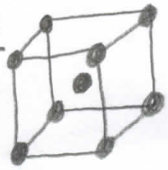
Coordination # = $\boxed{6}$

Spheres/unit cell = $\boxed{1}$

② Body Centered Cubic:

Coordination # = $\boxed{8}$

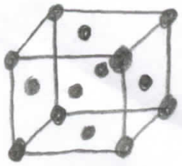
Spheres/unit cell = $\boxed{2}$



③ [illegible]ubic close packed "or" Face centered cubic:

Co[illegible]ation # = $\boxed{12}$

[illegible]it cell = $\boxed{4}$



④ Hexagonal close-packed:

Coordination # = $\boxed{12}$

Spheres/unit cell = $\boxed{6}$

Mass of a unit cell:

$$\frac{(\#spheres)(molar\ mass)}{6.022 \times 10^{23}\ mol}$$

Density of a solid:

Given: EDGE LENGTH

Volume = $(edge\ length)^3$

Density = $\frac{MASS}{Volume}$

Radius of an atom:

Given: DENSITY

Volume = $\frac{MASS}{Density}$

Edge Length = $\sqrt[3]{V}$

$(Length)^2$ = area

∴ $r = \frac{L}{\sqrt{8}}$

*KEY IDEAS

- Quantum
- Periodic table trends
- Lattice Energy

Chapter 5

Paramagnetic: partially filled orbitals

Diamagnetic: All orbitals filled

5.1 (Orbital energies)

- He^+ behaves like hydrogen but more stable because 2protons
- n=1 electrons screen n=2, n=3, n=4 electrons
- electrons in lower orbitals screen better
- The higher the quantum number "l", the more that electron is screened

- VALENCE electron = Outer most orbitals
- CORE electrons = all other electrons below it

Pauli Exclusion Principle (no exceptions)

-each electron in an atom has a unique set of quantum numbers

Aufbau Principle

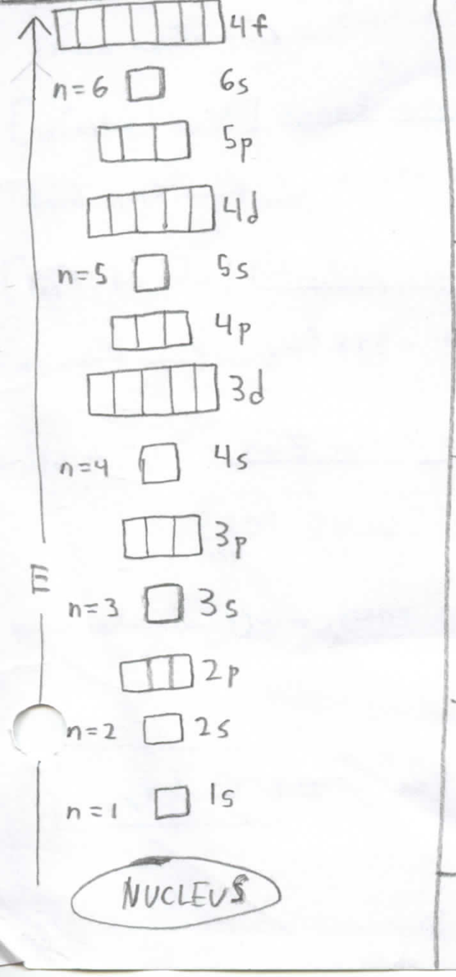
electrons are placed in the most stable orbital without violating the Pauli Exclusion principle

- The higher the value of n, the less stable the orbital
- "n" being equal, the higher the value of l, the less stable the orbital

Hund's Rule most stable orbital configuration is electrons all with same spin orientation

Electron Orbital Filling in order:

1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f



Groups (1-8): Valence e^- = group #

Groups (9-11): NO General Rule

Groups (12-18): Valence e^- = (group # - 10)

Exceptions:

Cr: [Ar] $4s^1 3d^5$

Cu: [Ar] $4s^1 3d^{10}$

Nb: [Kr] $5s^1 4d^4$

Ru: [Kr] $5s^1 4d^7$

(Affected here

41, 42, 44, 45, 46, 47

Au: [Xe] $6s^1 4f^{14} 5d^{10}$

Affected here

57, 58, 64, 78, 79

U: [Rn] $7s^2 5f^3 6d^1$

Neutral transition metals: s first

Va: [Ar] $4s^2 3d^3$

Cation transition metals: d first

Fe^{3+}: [Ar] $3d^5$

Cr^{3+}: [Ar] $3d^3$

Rules: (has exceptions)

Neutral atoms and anions:

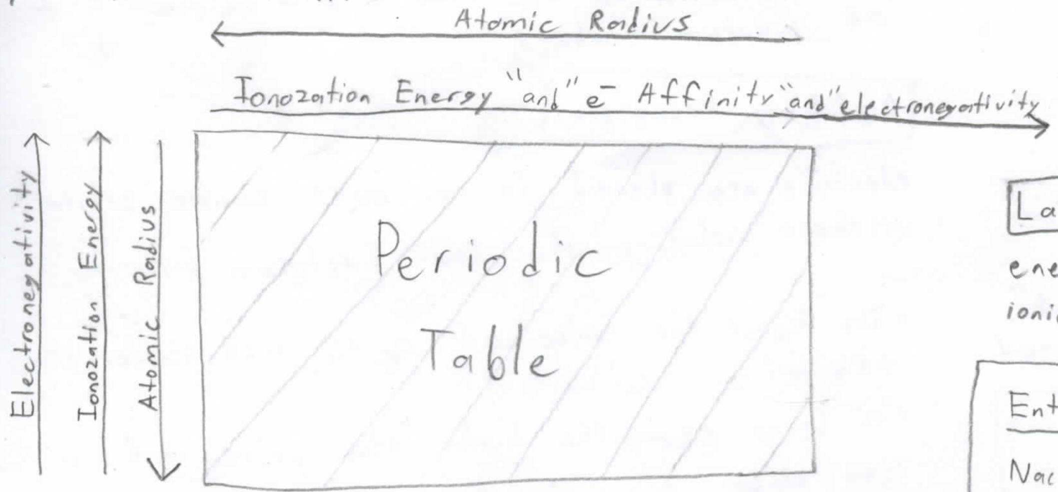
fill s orbital before d

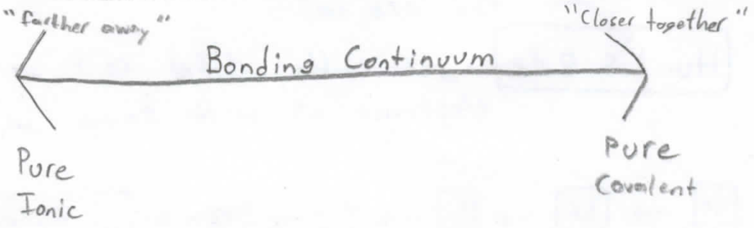
Cations:

fill d orbital before s

Periodic Table Trends:

Atomic Radius	Ionozation Energy	(how bad it wants an electron) Electron Affinity
↓ from left to Right	↑ from left to Right	↑ from left to Right
↑ Top to bottom	↓ Top to bottom	





0 COVALENT 0.5 POLAR 2 IONIC ∞

Lattice energy

energy to separate one mole of ionic solid into gaseous ions

Enthalpy of formation eqn: (don't balance)

$Na_{(s)} + \frac{1}{2}Cl_{2(g)} \xrightarrow{\Delta Hf^\circ} NaCl_{(s)}$

① Sublimation Energy: $[Na_{(s)} \rightarrow Na_{(g)}]$

$\Delta H_1^\circ = +105 kJ/mol$ ⟵ From book

② Bond Energy: $[\frac{1}{2}Cl_{2(g)} \rightarrow Cl_{(g)}]$

$\Delta H_2^\circ = +240 kJ/mol$ ⟵ From book

③ Ionozation Energy: $[Na_{(g)} \rightarrow Na^+_{(g)}]$

$\Delta H_3^\circ = +495.5 kJ/mol$ ⟵ From book

④ Electron Affinity: $[Cl_{(g)} \rightarrow Cl^-_{(g)}]$

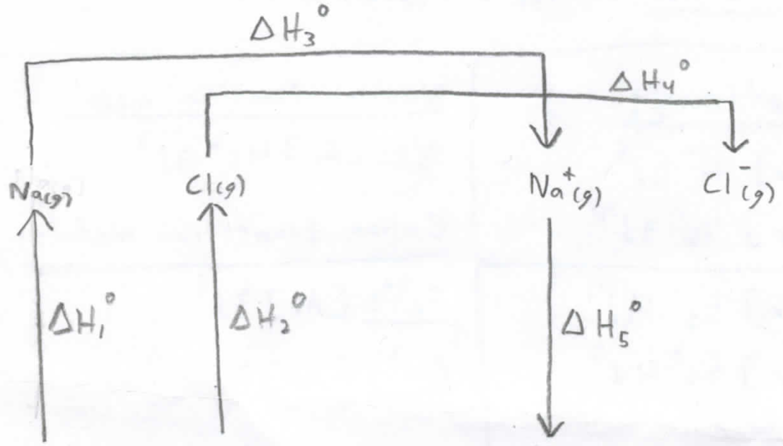
$\Delta H_4^\circ = -348.5 kJ/mol$ ⟵ From book

⑤ Condensation Energy: $[Na^+_{(g)} \rightarrow NaCl_{(s)}]$

$\Delta H_5^\circ =$ SOLVE FOR

*Lattice energy = −(Condensation energy)

Born Haber Cycle [NaCl]



$Na_{(s)} + \frac{1}{2}Cl_{2(g)} \xrightarrow{\Delta HF^\circ} NaCl_{(s)}$

$\Delta Hf^\circ = \Delta H_1^\circ + \Delta H_2^\circ + \Delta H_3^\circ + \Delta H_4^\circ + \Delta H_5^\circ$

$-411 = 105 + \frac{240}{2} + 495.5 - 348.5 + \Delta H_5^\circ$

$\therefore \Delta H_5^\circ = -783 kJ/mol$

Lattice Energy = $\boxed{783\ kJ/mol}$