

AP Chemistry Notes

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Special Thanks to Stephen Bosley (Boser)

Contents

1	FOREWORD/DISCLAIMER	4
2	Solubility Rules	5
2.1	Soluble	5
2.2	Insoluble	5
2.3	Naming Rules	5
3	Periodic Table of Elements	5
4	Poly Atomic Naming	6
5	Common Units, Constants and Charges	6
5.1	Fundamental Constants	6
5.2	Charge	7
5.3	Radius	7
6	Atomic Theory	7
6.1	J.J. Thompson	7
6.2	Robert Millikan	7
6.3	Ernest Rutherford	7
6.4	Chadwick	7
6.5	John Dalton	7
7	Naming	8
7.1	Binary	8
7.1.1	Greek Prefixes	8
7.2	Ionic	8
7.3	Acids	8
7.3.1	Polyatomic	8
7.3.2	Binary	8
8	Cations	9
9	Reaction Type	9
9.1	Combination (Synthesis)	9
9.2	Decomposition	9
9.2.1	Special Binary Salt Splits	10
9.3	Combustion	10
10	Blackbody Radiation	10
11	Bohr Model	11
11.1	Energy Level Formula	11
11.1.1	Energy Change during Level Jumps	11

12 Wavelength	11
12.1 De Broglie Formulas	11
13 Quantum Values	12
13.1 Quantum Value Table	12
13.2 Special cases	12
14 Periodicity	13
14.1 Electron Configuration	13
14.2 Isoelectricity	13
15 Nuclear Chemistry	13
15.1 Isotopes	13
15.2 Radiation	14
15.2.1 Alpha Radiation	14
15.2.2 Beta Radiation	14
15.2.3 Gamma Radiation	14
15.2.4 Positron Radiation	15
15.2.5 Electron Capture	15
15.3 Nuclear Equations	15
15.3.1 Radiation Table	15
15.4 Nuclear Stability	16
15.4.1 Forces Involved	16
15.4.2 Belt of Stability	16
15.4.3 Magic Numbers	16
15.4.4 Half-Life	17
16 Ionization and Affinity	17
16.1 Ionization Energy	17
16.2 Electron Affinity	17
17 Reactions of Metals	17
18 Chemical Bonds	17
18.1 Intramolecular	17
18.1.1 Ionic Bonding	18
18.1.2 Covalent Bonding	18
18.1.3 Metallic Bonding	18
18.2 Intermolecular	18
18.2.1 Ion-Dipole	18
18.2.2 Dipole-Dipole	18
18.2.3 Hydrogen Bond	18
18.2.4 London Dispersion/Van der Waals	18
18.2.5 Intermolecular Flowchart	19
18.3 Rule of Octet	19
19 Lewis Structures	19
19.1 Structures for Atoms	19
19.2 Structures for Ions	19
19.3 Structure for Ions of Molecules	19
19.4 Lewis Structures for Molecular Structures (Covalent)	20
19.5 Resonance Structures	21
19.5.1 Formal Charge	21
20 Lattice Energies of Ionic Solids	21
21 Bond Lengths of Covalent Bonds	22
22 Electronegativity	22
22.1 Dipole	22
22.1.1 Dipole Moment	22

23 Bond Enthalpy	23
24 VSEPR	23
24.1 Bond Shape Table	23
25 Organic Chemistry	23
25.1 Polarity	23
25.2 Alkanes	24
25.3 Alkane Branch Structure Naming	24
25.3.1 Branch Structure Naming Table	25
25.4 Alkenes	26
25.4.1 Alkene Naming	26
25.5 Alkynes	26
25.5.1 Alkyne Naming	27
26 Functional Groups	27
26.1 Alcohol	27
26.2 Aldehyde	27
26.3 Carboxylic Acid	28
26.4 Ester	28
26.5 Ketone	28
26.6 Ether	28
26.7 Amine	29
26.8 Amide	29
26.9 Haloalkane	29
27 Complex Ions	29
27.1 Cations	29
27.2 Anions	30
27.3 Coordination Number	30
27.4 Naming	30
27.4.1 Cations	30
27.4.2 Anions	30
28 Acidic and Basic Redox	30
28.1 Acidic	30
28.2 Basic	31
29 Thermodynamics	31
29.1 Enthalpy	31
29.1.1 Stoichiometry Problems	31
29.1.2 Calorimetry	32
29.1.3 Hess Law	32
29.1.4 Standard Heat of Formation	32
29.2 Entropy	33
29.2.1 State of Matter	33
29.2.2 Number of Moles of Gasses	33
29.2.3 Pressure of Gas	33
29.3 Gibbs Law of Free Energy	33
29.3.1 ΔH , ΔS , ΔG , Relationship Table	34
30 Chemical Kinetics and Rate Laws	34
30.1 Physical State	34
30.2 Concentration	34
30.3 Temperature	34
30.4 Pressure of Gas	34
30.5 Catalysts and Inhibitors	34
30.6 Rate Laws	35
30.6.1 Order Table	35

31 Reaction Mechanisms	36
31.1 Elementary Steps	37
32 Equilibrium	37
32.1 Types of Equilibrium	37
32.2 Equilibrium Constant Expressions	37
32.2.1 Converting Constants	37
33 Gas Laws	37
33.1 Gas Units and Conversions	37
33.2 Ideal Gas Law	37
33.3 Real Gas Law	38
33.4 Combined Gas Law	38
33.5 Daltons Law of Partial Pressures	38
33.6 Gas Collection over a Water Solution	39
34 ICE ICE (Baby)	39
35 Acids and Bases	39
35.1 Definitions of Acids and Bases	39
35.2 pH and pOH	40
35.2.1 Changing Concentrations	40
35.3 Strong Acids and Bases	40
35.3.1 Strong Acids	40
35.3.2 Strong Bases	40
35.4 Weak Acids and Bases	40
35.4.1 K_a Constant	41
35.4.2 K_b Constant	41
35.5 Common Ion Effect	41
35.6 Buffer	41
36 Equilibrium of Saturated, Soluable Salts	42
37 Kinetic Molecular Theory	42
37.1 Postulates:	42
37.2 Root Mean Square Velocity	43
37.3 Effusion and Diffusion	43
37.3.1 Effusion	43
37.3.2 Diffusion	43
37.3.3 Finding the rate	43
38 Electro Chemistry	43
38.1 Identifying Oxidation Numbers	43
38.2 Galvanic/Voltaic Cells	43
38.3 Calculating Cell Potential	44
38.3.1 Nernst Equation to Find E°_{cell}	44
39 Balancing Redox Reactions	44
39.1 Acidic	44
39.2 Basic	45

1 FOREWORD/DISCLAIMER

First and foremost, I am going to say what everyone has on their minds. **No** you really should not just forget about taking notes anymore in AP Chemistry class because of this packet. This packet is meant to be a **review** and should be used as such. However that does not mean you can use this packet as your main notes and write notes in the margins to supplement your learning. *Please take into account that this entire thing was written over the course of 4 days. As such it is inevitable that I made mistakes in spelling and/or formulas.*

If you have any questions/comments/fixes to the text you can email me at theangrybaby@gmail.com
Good Luck

2 Solubility Rules

2.1 Soluble

- Nitrates NO_3^{-1} - All nitrates are soluble
- Chlorates ClO_3^{-1} - All chlorates are soluble
- Alkali metal Cations and Ammonium cation compounds NH_4^{+1} are all soluble
- Chlorides, Bromides, and Iodides are all soluble EXCEPT Ag^{+1} , Pb^{+2} , and Hg^{+2}
- Acetates - All are soluble except Ag^{+}
- Sulfates - All are soluble except Ba^{+2} , Pb^{+2} , Hg^{+2} , Ca^{+2} , Ag^{+1} , and Sr^{+2}

2.2 Insoluble

- Carbonates CO_3^{-2} - all carbonates are insoluble except alkali metals and ammonium compounds
- Chromates CrO_4^{-2} - all chromates are insoluble except alkali metals, ammonium, Ca^{+2} , and Sr^{+2}
- Hydroxides OH^{-1} - all hydroxides are insoluble except alkali metals, ammonium, Ba^{+2} , Sr^{+2} , and Ca^{+2} although the last two (Sr^{+2} and Ca^{+2}) are only slightly soluble so a precipitate can form.
- Phosphates PO_4^{-3} all are insoluble except alkali metals and ammonium
- Sulfites SO_3^{-2} all are insoluble except alkali metals and ammonium
- Sulfides S^{-2} all are insoluble except Alkali metals, alkali earth metals and ammonium

2.3 Naming Rules

- All strong acids and bases are soluble and should be written as the ions when completing net ionic reactions
 - ▷ Sulfuric acid (H_2SO_4) should be written as $H^{+} + HSO_4^{-1}$
- The strong acids are: HCL , HBR , HI , HNO_3 , $HClO_4$, and H_2SO_4
- Strong bases are any alkali metal hydroxides ($LiOH$, $NaOH$, etc) and $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$
- All acids and bases should be left in their molecular form:
 - ▷ Acetic acid $\rightarrow HC_2H_3O_2$

3 Periodic Table of Elements

Periodic Table of the Elements															
Hydrogen 1 1.00794															Helium 2 4.00260
Lithium 3 6.941	Beryllium 4 9.0122													Boron 5 10.811	Neon 10 20.180
Sodium 11 22.990	Magnesium 12 24.305													Aluminum 13 26.982	Argon 18 39.948
Potassium 19 39.098	Calcium 20 40.078	Scandium 21 44.956	Titanium 22 47.867	Vanadium 23 50.942	Chromium 24 51.996	Manganese 25 54.938	Iron 26 55.845	Cobalt 27 58.933	Nickel 28 58.693	Copper 29 63.546	Zinc 30 65.38	Gallium 31 69.723	Germanium 32 72.61	Arsenic 33 74.922	Selenium 34 78.96
Rubidium 37 85.468	Sr 87.62	Yttrium 39 88.906	Zr 91.224	Nb 92.906	Mo 95.94	Tc 98	Ru 101.07	Rh 102.91	Pd 106.42	Ag 107.87	Cd 112.41	In 114.82	Sn 118.71	Sb 121.76	Te 127.60
Cesium 55 132.91	Ba 137.33	* 57-70	Lu 174.97	Hf 178.49	Ta 180.95	W 183.84	Re 186.21	Os 190.23	Ir 192.22	Pt 195.08	Au 196.97	Hg 200.59	Tl 204.38	Pb 207.2	Bi 208.98
Francium 87 [223]	Ra [226]	* *	103 [209]	104 [210]	105 [211]	106 [212]	107 [213]	108 [214]	109 [215]	110 [216]	111 [217]	112 [218]	114 [214]	116 [216]	118 [218]
* Lanthanide series															
57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb		
71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
** Actinide series															
89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No		

4 Poly Atomic Naming

- Zinc Zn^{+2}
- Silver Ag^{+1}
- Ammonium NH_4^{+1}
- Hydroxide OH^{-1}
- Cyanide CN^{-1}
- Nitrate NO_3^{-1}
- Acetate $C_2H_3O_2^{-1}$
- Chlorate ClO_3^{-1}
- Bromate BrO_3^{-1}
- Iodate IO_3^{-1}
- Manganate MnO_3^{-1}
- Sulfate SO_4^{-2}
- Bisulfate (Hydrogen Sulfate) HSO_4^{-1}
- Carbonate CO_3^{-2}
- Bicarbonate (Hydrogen Carbonate) HCO_3^{-1}
- Selenate SeO_4^{-2}
- Biselenate (Hydrogen Selenate) $HSeO_4^{-1}$
- Oxalate $C_2O_4^{-2}$
- Phosphate PO_4^{-3}
- Hydrogen Phosphate HPO_4^{-2}
- Dihydrogen Phosphate $H_2PO_4^{-1}$
- Chromate CrO_4^{-2}

Per _ Ate	Ate	Ite	Hypo _ Ite
Per _ Ic	Ic	Ous	Hypo _ Ous
+1 Oxygen	Most Common Ion	-1 Oxygen	-2 Oxygen

5 Common Units, Constants and Charges

5.1 Fundamental Constants

- Avogadros Number (N)
▷ $6.02214199 * 10^{23} mol^{-1}$
- Plancks Constant (h)
▷ $6.62606876 * 10^{-34} J * s$
- Speed of Light (c)
▷ $2.99792458 * 10^8 m/s$

5.2 Charge

- e^- charge = $-1.602 * 10^{-19}$ coulombs
- p^+ charge = $1.602 * 10^{-19}$ coulombs
- Atomic Mass Unit (amu) = $1.66054 * 10^{-24}$
 - ▷ p^+ = 1.0073 amu
 - ▷ n^0 = 1.0087 amu
 - ▷ e^- = $5.486 * 10^{-4}$ amu

5.3 Radius

Angstroms (\AA) = 10^{-10} meters

6 Atomic Theory

6.1 J.J. Thompson

- Discovered e^- and $\frac{\text{charge}}{\text{mass}}$ ratio
 - ▷ Charge to Mass ratio: $1.76 * 10^8$ Coulombs/Gram (Charge of e^- /mass)
- **Plum Pudding** Model of atom

6.2 Robert Millikan

- Found charge and mass of e^-
- Millikan Oil Drop:
 - ▷ Charge oil drops in a field and adjust field until drops levitate

6.3 Ernest Rutherford

- Discovered 3 types of radiation (Decay Particles)
 - ▷ Alpha particles: He^{2+} size, very damaging, stoppable - α
 - ▷ Beta particles - e^- size, damaging, hard to stop - β
 - ▷ Gamma particles - tiny, not so damaging, unstoppable - γ
- Also discovered proton and new **dense nucleus model**
 - ▷ Rutherford worked with α particles most and discredited Thompsons model of the nucleus

6.4 Chadwick

- Discovers neutron by shooting radiation at light elements and it watching it kick out a neutral particle

6.5 John Dalton

- **Four Postulates**
 - ▷ Everything made of atoms
 - ▷ Atoms of one element differ from those of a different element
 - ▷ Atoms will combine in whole number ratios
 - ▷ Atoms can not be created or destroyed
- Law of Constant Composition
 - ▷ In a compound, atom ratios are constant

7 Naming

7.1 Binary

- Smallest atomic number comes first
- Second element ends with -ide

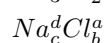
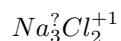
7.1.1 Greek Prefixes

- 1-Mono
- 2-Di
- 3-Tri
- 4-Tetra
- 5-Penta
- 6-Hexa
- 7-Hepta
- 8-Octa
- 9-Nona
- 10-Deca

Example
Cl_2O
Dichlorine Monoxide

7.2 Ionic

- Finding Charge:



$$\frac{(a*b)}{c} = d$$

7.3 Acids

7.3.1 Polyatomic

- Per...ate \rightarrow Per...ic acid
 $\triangleright HNO_4 \rightarrow$ pernitric acid
- ---ate \rightarrow ---ic acid
 $\triangleright H + NO_3 \rightarrow HNO_3$ (Nitric Acid)
- ---ite \rightarrow ---ous acid
 $\triangleright HNO_2 \rightarrow$ nitrous acid
- Hypo...ite \rightarrow hypo...ous acid
 $\triangleright HNO \rightarrow$ hyponitrous acid

7.3.2 Binary

- Hydro + (stem)ic
 $\triangleright H + Br \rightarrow$ Hydrobromic acid
 $\triangleright H + N \rightarrow$ Hydronitric acid
 \triangleright Hydrocarbonic acid $\rightarrow HC$
 \triangleright Carbonic Acid $\rightarrow HCO_3$

8 Cations

- Which cation forms a white precipitate with HCl ?
▷ Ag^+ (reversed proves Cl^-).
- What color is a typical Manganese solution?
▷ Pink/light purple. The precipitate is dark black.
- How would you test for Al^{3+} and what would it look like?
▷ Add Aluminum, it will make a precipitate red and leave the solution clear.
- Which cation forms a gel like precipitate?
▷ Aluminum.
- Which cation turns deep red with $KSCN$?
▷ Iron.
- How do you confirm the presence of zinc and what color is it?
▷ Add acid, then ammonia, which results in a white/bluish precipitate.
- What cation turns a deep blue with ammonia?
▷ Copper.

9 Reaction Type

9.1 Combination (Synthesis)

When two or more chemicals react to form one product

Example
$2Mg + O_2 \rightarrow 2MgO$
$2Na + S \rightarrow Na_2S$

- Metal + Non-Metal \rightarrow Metal Nonmetal (Binary Salt)
- Metal Oxide + Water \rightarrow Metal Hydroxide
▷ $CaO + H_2O \rightarrow Ca(OH)_2$
▷ $K_2O + H_2O \rightarrow 2KOH$
- Metal Oxide + $CO_2 \rightarrow$ Metal Carbonate
▷ $Na_2O + CO_2 \rightarrow Na_2CO_3$
- Nonmetallic Oxides + water \rightarrow Acids (nonmetal oxides retains its oxide number)
▷ $Na_2O + SO_3 \rightarrow Na_2SO_4$

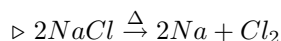
9.2 Decomposition

When one chemical decomposes into 2 or more

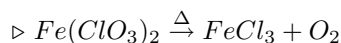
Example
$2Ag_2O \xrightarrow{\Delta} 4Ag + O_2$
$\Delta = \text{Heat}$

- Metal Carbonate $\xrightarrow{\Delta}$ Metal Oxide + CO_2
▷ $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$
- Metal Hydroxide $\xrightarrow{\Delta}$ Metal Oxide + H_2O
▷ $Mg(OH)_2 \xrightarrow{\Delta} MgO + H_2O$

- Metal Nonmetal $\xrightarrow{\Delta}$ Metal + Nonmetal (diatomic in nature)

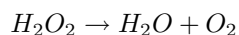
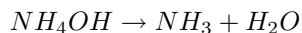
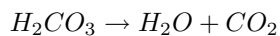
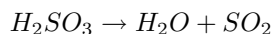


- Metal Chlorates $\xrightarrow{\Delta}$ Metal Chlorides + O_2

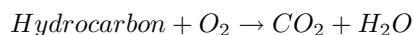


9.2.1 Special Binary Salt Splits

These binary salts split into different elements



9.3 Combustion



.... \downarrow

$C_xH_y \rightarrow$ double x (multiply by 2) then add 2

- C_1 : meth
- C_2 : eth
- C_3 : pro
- C_4 : bu
- C_5 : pent
- C_6 : hex
- C_7 : hept
- C_8 : oct
- C_9 : non
- C_{10} : dec

10 Blackbody Radiation

When an object is heated it will emit radiant energy

$$E = h\nu$$

- E = Energy
- h = Max Planck's constant ($6.626 \times 10^{-34} J \cdot s$)
- ν = frequency

Photoelectric effect: Metal will give off e^- s if light shines on it. Light shining on a clean sheet of metals will release e^- s if ν is strong enough.

11 Bohr Model

Neils Bohr:

1. Only orbits of certain radii, corresponding to certain definite energies are permitted for the electron in a hydrogen atom.
2. An electron in a permitted orbit has a specific energy and is in an allowed energy state. An electron in an allowed state will not radiate energy and therefore will not spiral into the nucleus.
3. Energy is emitted or absorbed by the e^- only as the e^- changes from one allowed energy state to another.
4. **Flawed theory because it only works for hydrogen**

11.1 Energy Level Formula

$$E_n = (-2.18 * 10^{-18} J) \left(\frac{1}{n^2} \right)$$

- $E_1: -2.18 * 10^{-18} J$
- $E_2: -5.45 * 10^{-19} J$
- $E_3: -2.42 * 10^{-19} J$
- $E_4: -1.36 * 10^{-19} J$
- $E_5: -8.72 * 10^{-20} J$
- $E_6: -6.056 * 10^{-20} J$
- $E_\infty: 0$

11.1.1 Energy Change during Level Jumps

$$\Delta E = E_F - E_0$$

- $n = 3 \rightarrow 2 \mid -3.03 * 10^{-19} J$
- $n = 4 \rightarrow 2 \mid -4.09 * 10^{-19} J$
- $n = 5 \rightarrow 2 \mid -4.578 * 10^{-19} J$
- $n = 6 \rightarrow 2 \mid -4.844 * 10^{-19} J$

12 Wavelength

12.1 De Broglie Formulas

$$\lambda = \frac{h}{mv}$$

or

$$\lambda = \frac{h}{p}$$

- λ = Wavelength
- h = Planck's Constant ($6.626 * 10^{-34} J * s$)
- m = Mass of particle in **Kg**
- v = Velocity of particle ($\frac{meters}{second}$)
- p = Momentum

Example
$m = 9.11 * 10^{-28} g$
$v = 5.97 * 10^6 m/s$
$\lambda = \frac{6.626 * 10^{-34} J * s}{(9.11 * 10^{-31} Kg)(5.97 * 10^6 m/s)} = 1.22 * 10^{-10} m$

13 Quantum Values

1. Principle Quantum number - (n)

$$n = 1 \text{ (lowest)}$$

$$n = \infty \text{ (at 8 or 9)}$$

$$\text{Follows Bohrs } E_n = (-2.18 \times 10^{-18} J) \left(\frac{1}{n^2} \right)$$

2. Azimuthal Quantum number - (l)

$$l = n - 1$$

if...

- $l = 0 \rightarrow$ S shape
- $l = 1 \rightarrow$ P shape
- $l = 2 \rightarrow$ D shape
- $l = 3 \rightarrow$ F shape

Example
$n = 3$
$l = 2$
\downarrow
3d

3. Magnetic Quantum number (orbital) - (m_l)

$-l$ and l including zero

$$m_0 = 0$$

$$m_1 = -1, 0, 1$$

$$m_2 = -2, -1, 0, 1, 2$$

4. Spin magnetic quantum number - (m_s)

$$+\frac{1}{2} \text{ or } -\frac{1}{2}$$

13.1 Quantum Value Table

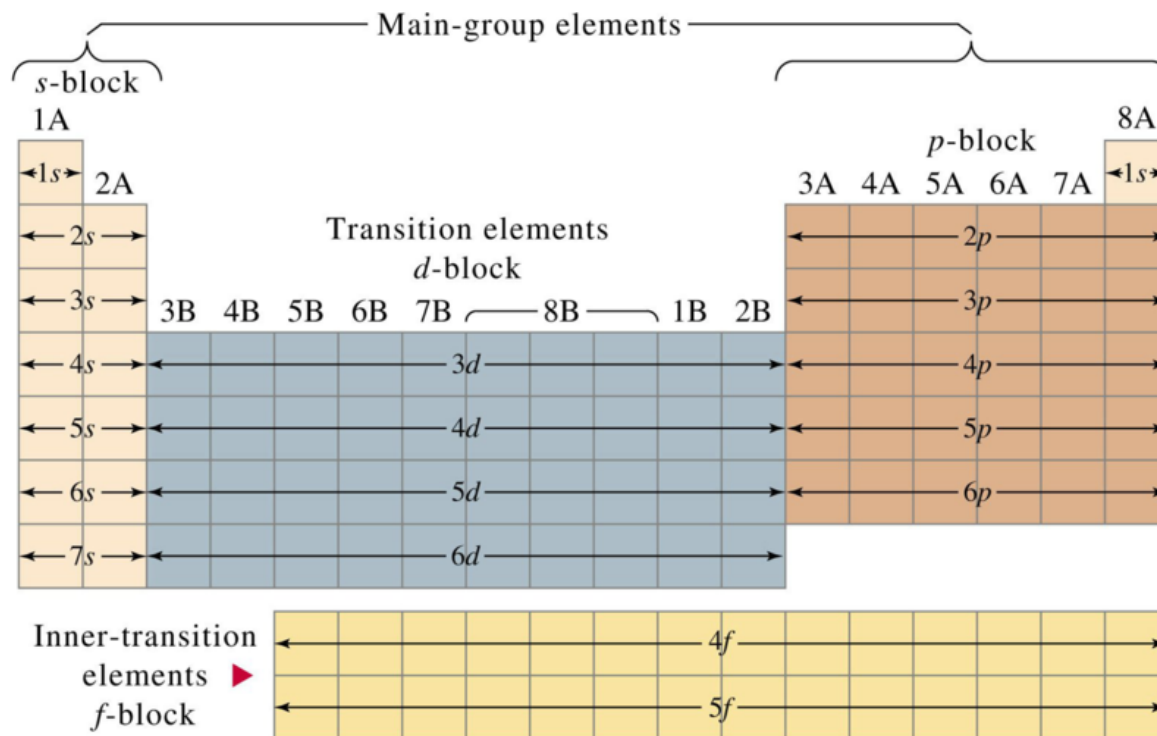
n	Possible l values	Subshell	m_l values	# of orbitals in subshell	total # of orbitals in shell	e^- in shell
1	0	1s	0	1	1	2
2	0	2s	0	1	4	8
	1	2p	-1,0,1	3	-	-
3	0	3s	0	1	9	18
	1	3p	-1,0,1	3	-	-
	2	3d	-2,-1,0,1,2	5	-	-
4	0	4s	0	1	16	32
	1	4p	-1,0,1	3	-	-
	2	4d	-2,-1,0,1,2	5	-	-
	3	4f	-3,-2,-1,0,1,2,3	7	-	-

13.2 Special cases

- Chromium has 6 half-filled orbitals
- Copper has one half-filled orbital and 5 filled orbitals

14 Periodicity

14.1 Electron Configuration



14.2 Isoelectricity

Two atoms are considered isoelectric when they gain or lose electrons to become ions and have the same electron configuration as each other.

Example
Na^{+1} : $1S_2, 2S_2, 2P_6$
Ne : $1S_2, 2S_2, 2P_6$

15 Nuclear Chemistry

Nuclear Chemistry involves changes in the nucleus of an atom.

Normal	Nuclear
Reactions involve electron transfer	Reactions involve decay of nucleus i.e. transforming one element into another
Reaction affected by factors such as pH, temp, pressure, [], etc.	Affected by the type of decay and the half-life of what is decaying
Reactions involve relatively small energy: 400 kJ-1500kJ	Reactions deal with huge amounts of energy

15.1 Isotopes

Isotopes: Atoms of the same element that have a different number of neutrons

$$X - A$$

A_ZX

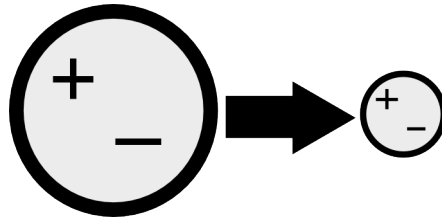
AX

- X = Element Symbol
- A = Atomic Mass
- Z = Atomic Number

15.2 Radiation

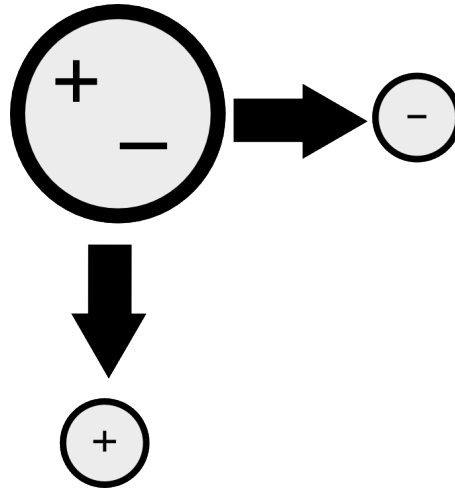
15.2.1 Alpha Radiation

When a big nucleus ejects a He^{+2} size chunk of itself.



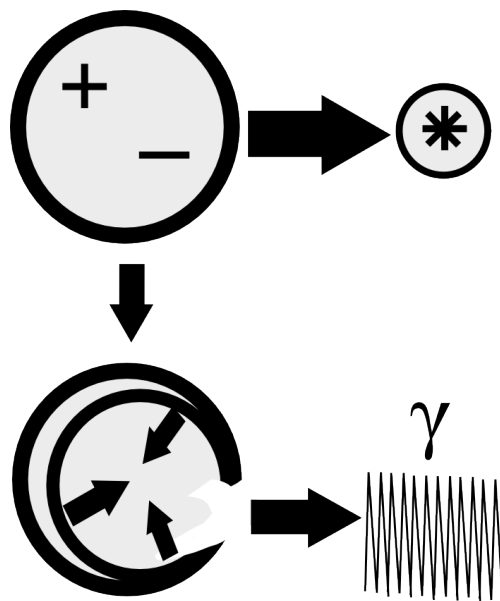
15.2.2 Beta Radiation

When a neutrally charged particle (equal amount of p^+ s and e^- s) ejects its e^- s leaving only the p^+ s.



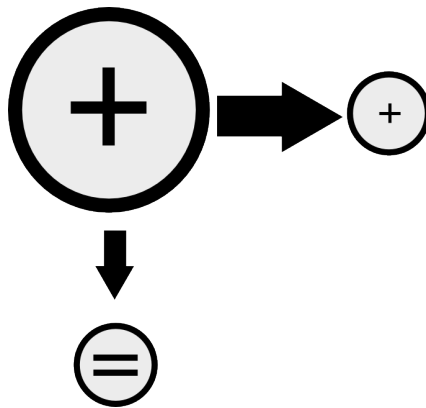
15.2.3 Gamma Radiation

When a particle experiences some type of radiation (called * here) that causes the remaining nucleus to collapse. This causes gamma (γ) rays to be emitted. Gamma radiation is also caused when a positron and an electron smash into each other.



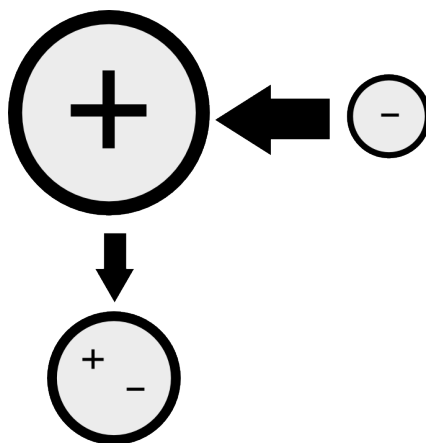
15.2.4 Positron Radiation

When a positively charged nucleus emits its p^+ leaving only the n^0 .



15.2.5 Electron Capture

When an electron in orbit falls into the nucleus (positively charged) and makes it neutrally charged.



15.3 Nuclear Equations

15.3.1 Radiation Table

Neutron:	1_0n
Proton:	${}^1_1p^+$
Electron:	${}^0_{-1}e^-$
Positron:	${}^0_1e^-$
Alpha Particle:	4_2He or ${}^4_2\alpha$
Beta Particle:	${}^0_{-1}e^-$ or ${}^0_{-1}\beta$

Example
Alpha ${}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^4_2He$
Beta ${}^{131}_{53}I \rightarrow {}^{131}_{54}Xe + {}^0_{-1}e^-$ ${}^1_0n \rightarrow {}^1_1p + {}^0_{-1}e^-$
Positron ${}^{11}_6C \rightarrow {}^{11}_5B + {}^0_1e^-$ ${}^1_1p \rightarrow {}^1_0n + {}^0_1e^-$
Electron Capture ${}^{81}_{37}Rb + {}^0_{-1}e^- \rightarrow {}^{81}_{36}Kr$ ${}^1_1p + {}^0_{-1}e^- \rightarrow {}^1_0n$
Positron-Electron Collision (Gamma) ${}^0_1e^- + {}^0_{-1}e^- \rightarrow {}^0_0\gamma$

15.4 Nuclear Stability

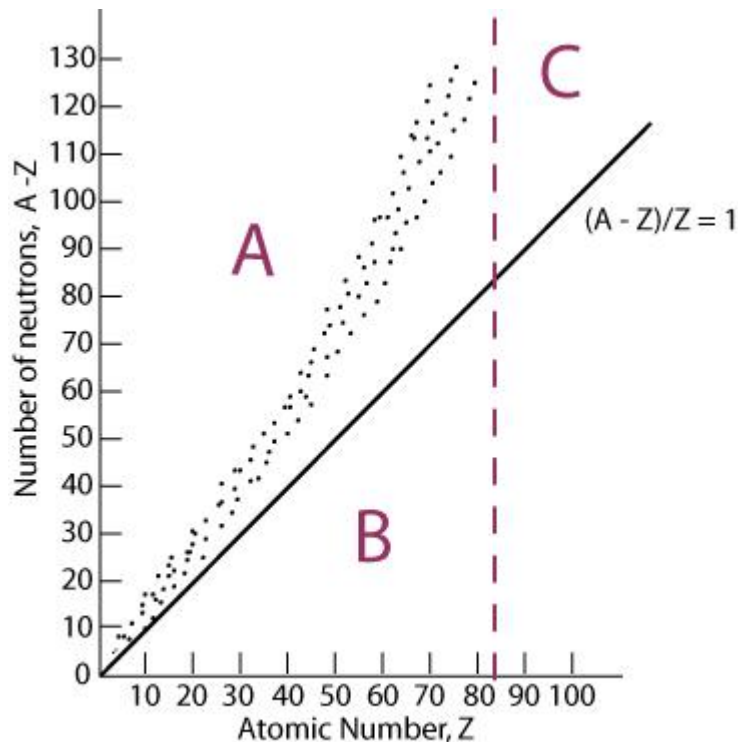
Understanding why are some nuclides are radioactive while others are not.

15.4.1 Forces Involved

- Electrostatic
 - ▷ Try to rip apart the nucleus because of like charges
- Strong Nuclear
 - ▷ Try to pull together the nucleus because subatomic particles naturally stick together
- The Glue
 - ▷ Neutrons act as the glue and more of it is required when the electrostatic force gets really strong

15.4.2 Belt of Stability

- Area A
 - ▷ More neutrons than protons - **Beta decay** → creates protons
- Area B
 - ▷ Less neutrons than protons - **Positron emission** (Smaller B) or **Electron Capture** (Larger B)
- Area C
 - ▷ Every element above 83 p^+ is radioactive and no glue can hold it together - **Alpha decay**



15.4.3 Magic Numbers

The Magic Numbers tend to be stable if you have either a proton or neutron in those numbers. If you have both, they are very stable.

(p^+)	2	8	20	28	50	82	-
(n°)	2	8	20	28	50	82	126

- If (p^+) and (n°) even → likely stable
- If either is odd → could go either way
- If (p^+) and (n°) odd → likely unstable

15.4.4 Half-Life

The time it takes $\frac{1}{2}$ the amount of a substance to decay.

Example	
5g of nuclide	
$\frac{1}{2}$ life of 15 years	
How much of the original nuclide remains after 45 years?	
5	
↓	(15 years)
2.5	
↓	(30 years)
1.25	
↓	(45 years)
0.625g	

16 Ionization and Affinity

16.1 Ionization Energy

The energy needed to remove an e^- (how easy it is to lose an e^-). **Needs** energy (+).

16.2 Electron Affinity

How much a gaseous atom will be attracted to a free e^- (how easy it is to gain an e^-). **Releases** energy (-).

17 Reactions of Metals

Metal Oxides = Basic

- Metal + Water \rightarrow Metal Hydroxide + H_2
- Metal + O_2 (Li or any non-Alkali metal) \rightarrow Metal Oxide
- K + O_2 (Any other Alkali metal) \rightarrow Metal Peroxide (O_2^{-1})
 - ▷ K + $O_2 \rightarrow KO_2$
- Metal Oxide + $H_2O \rightarrow$ Metal Hydroxide
 - ▷ $Na_2O + H_2O \rightarrow NaOH$
- Metal Oxide + Acid \rightarrow Salt + H_2O
 - ▷ $Na_2O + HCl \rightarrow NaCl + H_2O$

Nonmetal Oxides = Acidic

- Nonmetal Oxide + $H_2O \rightarrow$ Acid
 - ▷ $CO_2 + H_2O \rightarrow H_2CO_3$
 - ▷ $SO_2 + H_2O \rightarrow H_2SO_3$
 - ▷ $P_4O_{10} + H_2O \rightarrow H_3PO_4$
- Nonmetal Oxide + Base \rightarrow Salt + H_2O
 - ▷ $CO_2 + NaOH \rightarrow Na_2CO_3 + H_2O$

18 Chemical Bonds

When 2 or more atoms are strongly attached (attracted) to each other.

18.1 Intramolecular

These forces act inside an atom or molecule:

18.1.1 Ionic Bonding

Gain/lose e^- s (strong metal + strong nonmetal)

18.1.2 Covalent Bonding

Share e^- s (weak metal or nonmetal + nonmetal)

18.1.3 Metallic Bonding

There are two models that explain metallic bonding:

- **Electron Sea Model** ¹

▷ Metal atoms are floating in a sea of e^- s. No one e^- belongs to any particular atom.

- **Orbital Bonding Model**

▷ The valence e^- s are overlapped and shared so much you have bonds of delocalized e^- s that are free to move but are still holding the atoms together.

Properties that result from metallic bonding include:

- Conductivity of electricity and heat
- Malleability and ductility
- Ability to form alloys

18.2 Intermolecular

These forces act between molecules:

18.2.1 Ion-Dipole

Ions bonding to molecules with a dipole (polar solvent). **The strongest intermolecular force.**

18.2.2 Dipole-Dipole

Polar near another polar. **Weaker than Ion-Dipole but still strong, based on how strong the polarity is.**

18.2.3 Hydrogen Bond

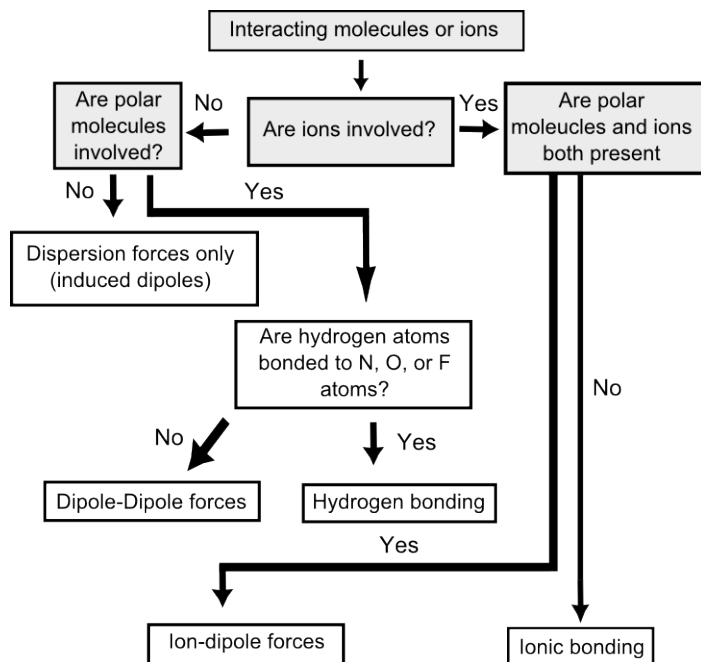
Either ($H - N$), ($H - O$), or ($H - F$). No shielding e^- s on Hydrogen so central atoms e^- pair gets full pull of Hydrogen nucleus.

18.2.4 London Dispersion/Van der Waals

An induced dipole between 2 polar molecules. An increase in pressure or decrease in temperature will cause one side to have a more positive force as the majority of e^- s move to other side.

¹Of the two theories, this is generally the more accepted one

18.2.5 Intermolecular Flowchart

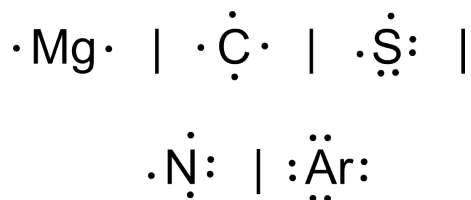


18.3 Rule of Octet

Atoms tend to bond in such a way as to gain, lose, or share e^- s in order to gain a complete valence (outer s and p).

19 Lewis Structures

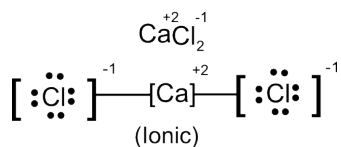
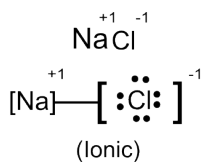
19.1 Structures for Atoms



19.2 Structures for Ions

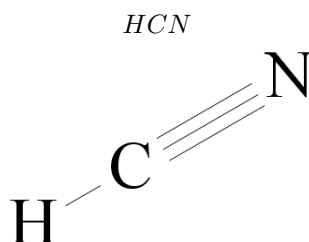
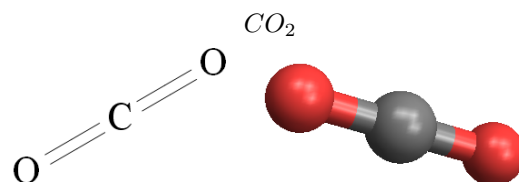
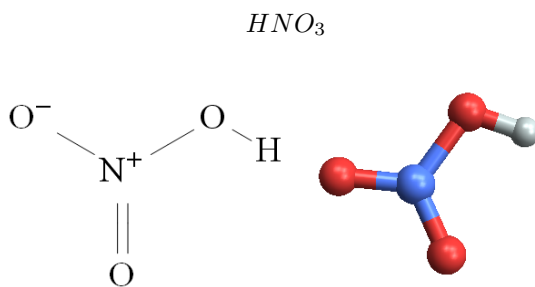
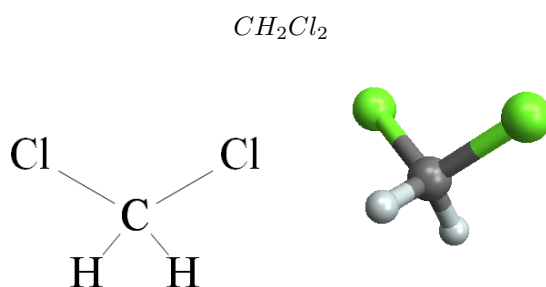
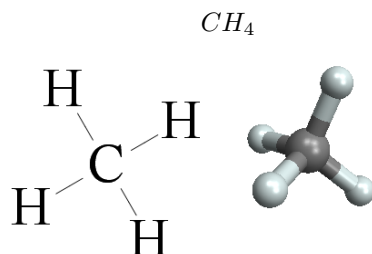


19.3 Structure for Ions of Molecules

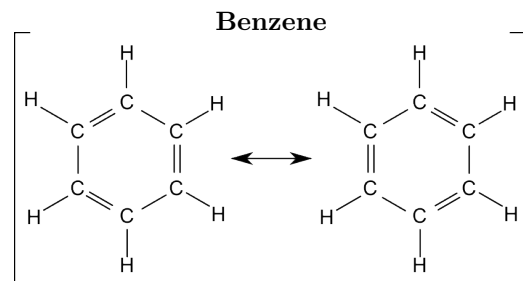


19.4 Lewis Structures for Molecular Structures (Covalent)

1. Add valence e^- s from all the atoms.
2. Write the symbols for the atoms. If there are more than 2 atoms, identify the central atom. Connect them with a single line which represents 2 shared e^- s. Subtract the number of e^- s from total found in step 1.
 - ▷ Central atom will be closest to Si, P or Metaloid staircase.
3. Complete octets around the atoms bonded to the central atom (Hydrogen does not get more than 2).
4. Place the remaining pairs around the central atom even if doing so gives more than an octet to the central atom.
5. If there are not enough pairs to complete an octet in the central atom, then you need to try using double or triple bonds.

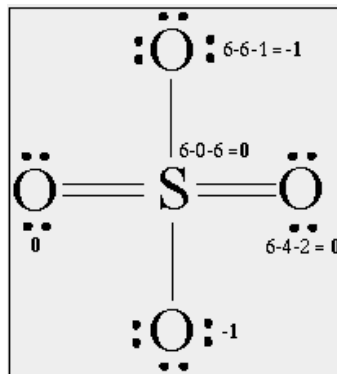


19.5 Resonance Structures



19.5.1 Formal Charge

Valence e^- s of an atom - (total unbonded e^- s + $\frac{1}{2}$ total bonded e^- s)



Molecular structures that tend to be the common one have a formal charge is closest to zero and any negative charge is on the most electronegative element.

20 Lattice Energies of Ionic Solids

Coulombs Law

$$E = \frac{KQ_1Q_2}{d}$$

- Q_1/Q_2 = ion charges
- d = Distance between ions of the final crystalized lattice form.
- † **The greater the charge, the higher the energy.**
- † **The closer the ions, the higher the energy.**

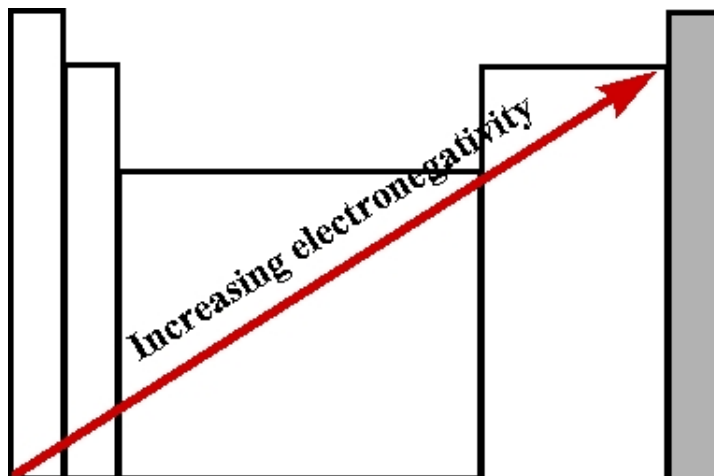
Example
Which has a greater lattice energy?
$\overset{+1}{Na}\overset{-1}{Cl}$ vs $\overset{+2}{Mg}\overset{-1}{Cl}_2$
$\overset{+2}{Mg}\overset{-1}{Cl}_2$ has greater charges thus a higher lattice energy.
$\overset{+1}{Li}\overset{-1}{Cl}$ vs $\overset{+1}{Na}\overset{-1}{Cl}$
$\overset{+1}{Li}$ is smaller than $\overset{+1}{Na}$ so $\overset{+1}{Li}$ will be closer to $\overset{-1}{Cl}$ than $\overset{+1}{Na}$ will so $\overset{+1}{Li}\overset{-1}{Cl}$ will have a higher lattice energy.

21 Bond Lengths of Covalent Bonds

- Single - Longest
- Double - Medium
- Triple - Shortest

Length		
Single	CO_4^{-4}	$1.42 \overset{\circ}{A}$
Double	CO_2	$1.24 \overset{\circ}{A}$
Triple	CO	$1.13 \overset{\circ}{A}$

22 Electronegativity



Difference in electronegativity determines the character of the bond.

- Large difference \rightarrow Ionic Bond
 - ▷ Biggest difference is 3.3
- Medium difference \rightarrow Polar Covalent
 - ▷ HF - 1.8
- Small/No difference \rightarrow Non-Polar Covalent
 - ▷ H_2 - 0

22.1 Dipole



Arrow points towards more electronegative atom.

22.1.1 Dipole Moment

Numeric value that represents how strong the dipole is

Example
Which has the greater dipole moment? <i>OR</i> Which has greater electronegative difference?
<i>HI or HF</i>
Answer: HF

23 Bond Enthalpy

ΔH : Energy given off or taken in during a reaction.

- $\Delta H = -$
▷ Exothermic
 - $\Delta H = +$
▷ Endothermic
- ⊢ **Breaking bonds requires energy**
- ⊢ **Forming bonds releases energy**

Example	
$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$	
Breaking	Forming
$4 * (C - H) = (4 * 413)$	$2 * (C=O) = (2 * 799)$
$2 * (O_2) = (2 * 495)$	$4 * (H - O) = (4 * 463)$
2642	3450
$\Delta H = \text{Broken} - \text{Formed}$	
$\Delta H = 2642 - 3450 = -808 \text{ KJ}$	

24 VSEPR

VSEPR stands for Valence Shell Electron Pair Repulsion. Make sure when counting bonds to treat double and triple bonds like a single bond. Also keep in mind that bonded pairs and lone pairs repel.

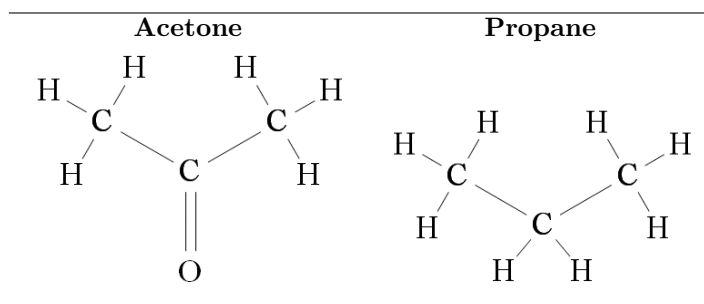
24.1 Bond Shape Table

Shape	Example	Total e^-	Bonded Pairs	Lone e^- Pairs	Hybrid Orbital
Linear	BeH_2	2	2	0	sp
Trigonal Planar	BCl_3	3	3	0	sp_2
Bent	NO_2^-	3	2	1	sp_2
Tetrahedral	CH_4	4	4	0	sp_3
Trigonal Pyramidal	NH_3	4	3	1	sp_3
Bent	H_2O	4	2	2	sp_3
Trigonal Bipyramidal	PCl_5	5	5	0	sp_3d
See-Saw	SF_4	5	4	1	sp_3d
T-Shape	BrF_3	5	3	2	sp_3d
Linear	ICl_2	5	2	3	sp_3d
Octahedral	SF_6	6	6	0	sp_3d_2
Square Pyramidal	BrF_5	6	5	1	sp_3d_2
Square Planer	ICl_4^-	6	4	2	sp_3d_2

25 Organic Chemistry

25.1 Polarity

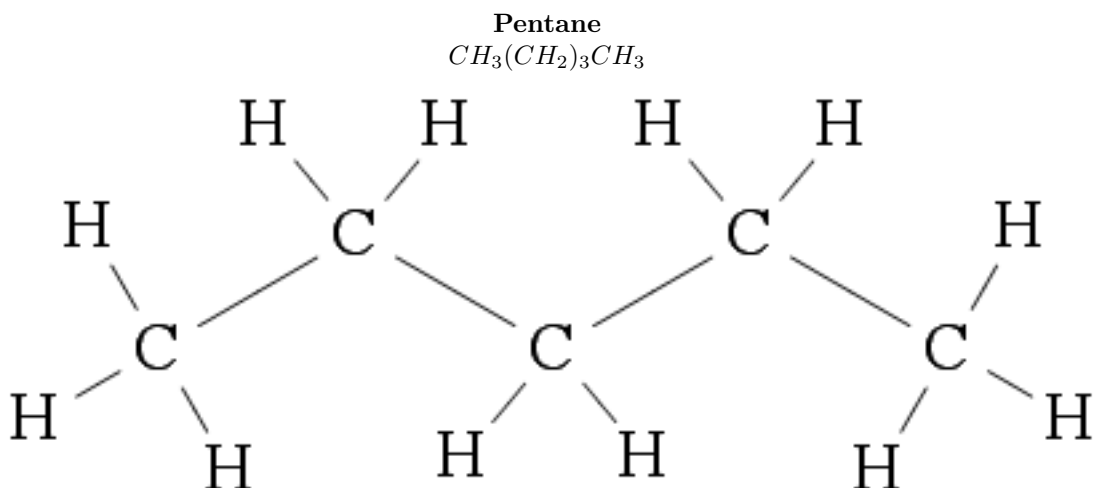
Polarity in regards to organic chemistry relies on an element disrupting the symmetry of a molecule. For example the double bonded oxygen in Acetone allows it to be more polar than Propane.



25.2 Alkanes

† Spotted by seeing a single bond

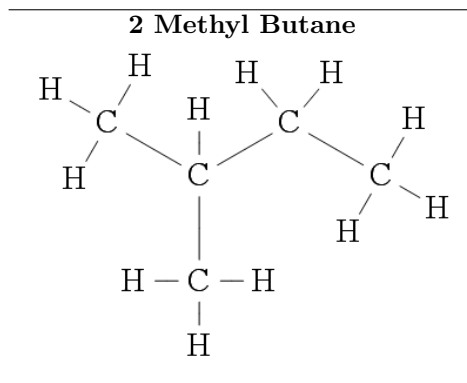
- $CH_4 \rightarrow$ Methane
- $C_2H_6 \rightarrow$ Ethane
- $C_3H_8 \rightarrow$ Propane
- $C_4H_{10} \rightarrow$ Butane
- $C_5H_{12} \rightarrow$ Pentane
- $C_6H_{14} \rightarrow$ Hexane
- $C_7H_{16} \rightarrow$ Heptane
- $C_8H_{18} \rightarrow$ Octane
- $C_9H_{20} \rightarrow$ Nonane
- $C_{10}H_{22} \rightarrow$ Decane



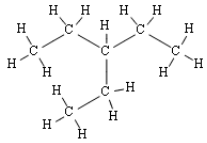
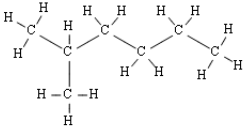
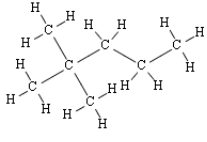
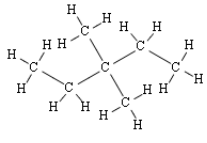
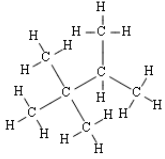
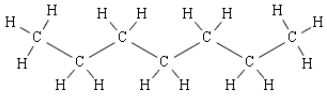
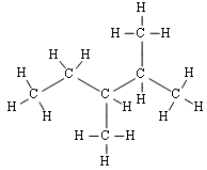
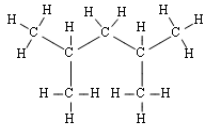
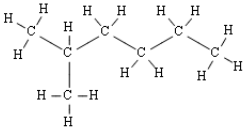
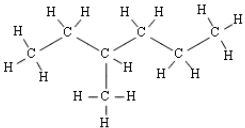
25.3 Alkane Branch Structure Naming

To name all single-bonded Carbon chains, see subsection on Alkanes above.

To name a branch structure first look for the longest unbroken Carbon chain, this is the root name. Then take the root prefix of the alkyl (the branch of the root chain) and add -yl (for instance Methane becomes methyl). Number the Carbon chain giving the side with an alkyl the lowest number. The end result should be something such as *2 Methyl Butane*.



25.3.1 Branch Structure Naming Table

	C_7H_{16}	3-ethylpentane
	C_7H_{16}	isoheptane
	$CH_3CH_2CH_2C(CH_3)_3$	2,2-dimethylpentane
	$CH_3CH_2C(CH_3)_2CH_2CH_3$	3,3-dimethylpentane
	$(CH_3)_2CHC(CH_3)_3$	2,2,3-trimethylbutane
	$CH_3(CH_2)_5CH_3$	<i>N</i> -heptane
	$C_2H_5CH(CH_3)CH(CH_3)_2$	2,3-dimethylpentane
	$(CH_3)_2CHCH_2CH(CH_3)_2$	2,4-dimethylpentane
	$CH_3(CH_2)_3CH(CH_3)_2$	2-methylhexane
	$CH_3CH_2CH_2CH(CH_3)CH_2CH_3$	3-methylhexane

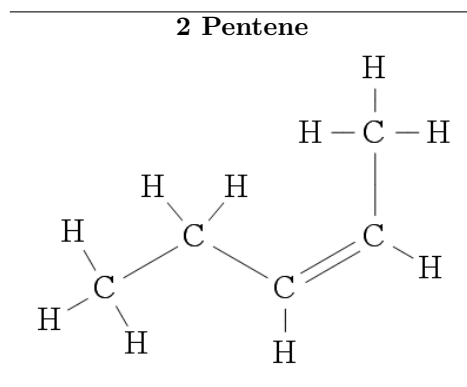
25.4 Alkenes

† Spotted by seeing a double bond

- $CH_2 \rightarrow$ Methene
- $C_2H_4 \rightarrow$ Ethene
- $C_3H_6 \rightarrow$ Propene
- $C_4H_8 \rightarrow$ Butene
- $C_5H_{10} \rightarrow$ Pentene
- $C_6H_{12} \rightarrow$ Hexene
- $C_7H_{14} \rightarrow$ Heptene
- $C_8H_{16} \rightarrow$ Octene
- $C_9H_{18} \rightarrow$ Nonene
- $C_{10}H_{20} \rightarrow$ Decene

25.4.1 Alkene Naming

Naming Alkenes is similar to naming Alkanes save for the naming of the root chain. To name the root chain you must give side where the double bond is the lowest number and name all branches after using this number scheme. You should end up with something like *2 Pentene*



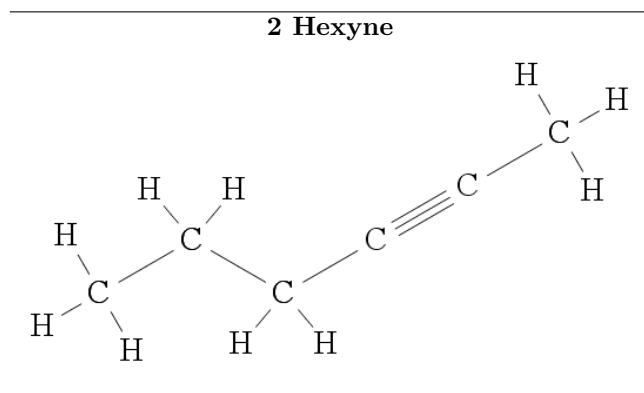
25.5 Alkynes

† Spotted by seeing a triple bond

- $CH \rightarrow$ Methyne
- $C_2H_2 \rightarrow$ Ethyne
- $C_3H_4 \rightarrow$ Propyne
- $C_4H_6 \rightarrow$ Butyne
- $C_5H_8 \rightarrow$ Pentyne
- $C_6H_{10} \rightarrow$ Hexyne
- $C_7H_{12} \rightarrow$ Heptyne
- $C_8H_{14} \rightarrow$ Octyne
- $C_9H_{16} \rightarrow$ Nonyne
- $C_{10}H_{18} \rightarrow$ Decyne

25.5.1 Alkyne Naming

Naming Alkynes is similar to naming Alkenes. Identify the root chain as you would using Alkenes except now you identify the triple bond instead of the double bond.

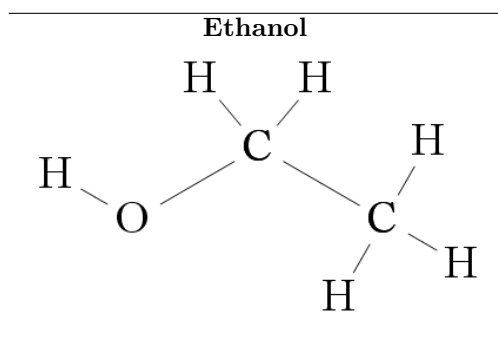


26 Functional Groups

» When discussing functional groups, the letter **R** is used to signify any hydrocarbon or hydrocarbon chain.

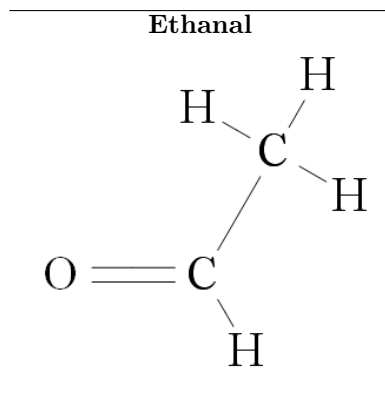
26.1 Alcohol

- **Root Name:** -ol
- **Identification:** R-OH



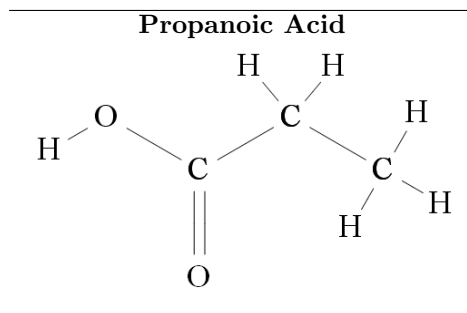
26.2 Aldehyde

- **Root Name:** -al
- **Identification:** R-CHO



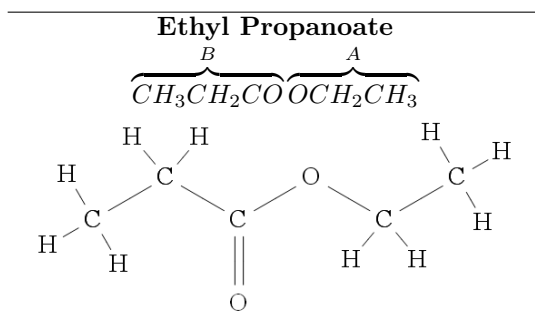
26.3 Carboxylic Acid

- Root Name: -oic
- Identification: R-COOH



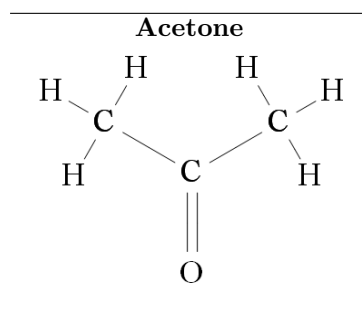
26.4 Ester

- Root Name: A-yl B-oate
- Identification: R-COO-R



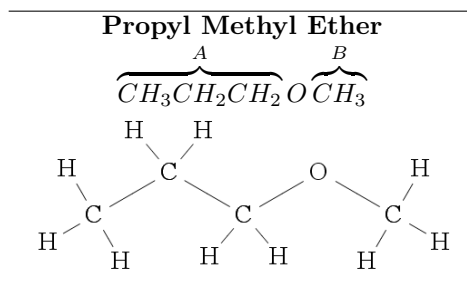
26.5 Ketone

- Root Name: -one
- Identification: R-CO-R



26.6 Ether

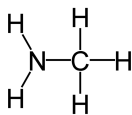
- Root Name: A-yl B-yl Ether
- Identification: R-O-R



26.7 Amine

- **Root name:** -amine
- **Identification:** R-NH₂

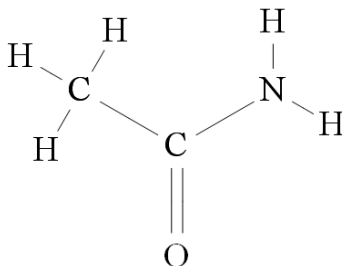
Methylamine



26.8 Amide

- **Root Name:** -amide
- **Identification:** R-CONH₂

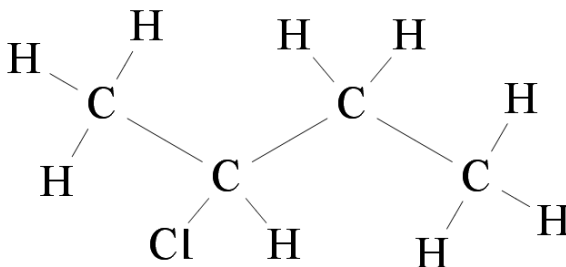
Ethanamide



26.9 Haloalkane

- **Root Name:** None, use standard naming of root chain
- **Identification:** Some Hydrogens in a hydrocarbon are replaced with a halogen (F, Cl, Br, I)

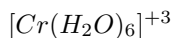
2 Chloro Butane



27 Complex Ions

Complex Ions are usually metal ions with attached ligands (Lewis Bases).

27.1 Cations



→ The charge of a cation is the charge of the transition metal (Cr in this case).

27.2 Anions



⊢ The charge of an anion is determined by the individual charges of the elements.

$$\triangleright Al^{+3} + 4(OH)^{-1}$$

$$\triangleright 3 - 4$$

$$\triangleright -1$$

27.3 Coordination Number

Generally (especially with cations) the coordination number is twice the charge of the transition metal.

Example
$[Cr(H_2O)_6]^{+3}$
$Cr^{+3} \rightarrow 3 * 2 = 6$

27.4 Naming

27.4.1 Cations

- Give the prefix associated with the coordination number
- Give appropriate name for ligand
- Name the transition metal
- Give roman numeral of transition metal

Example
$\left[\overset{\text{Chromium}}{\underbrace{Cr}} \underset{\text{Aqua}}{\underbrace{(H_2O)}_6} \overset{\text{Hexa}}{\underbrace{\quad}_6} \right]^{+3}$
Hexa Aqua Chromium (III)

27.4.2 Anions

- Give prefix associated with the coordination number
- Give appropriate Ligand name
- Name transition metal with -ate ending
- Give roman numeral

Example
$[Al(OH)_4]^{-1}$
Tetra Hydroxo Aluminate
<i>No roman numeral because Al is always +3</i>

28 Acidic and Basic Redox

28.1 Acidic

- Find oxidation number
- Write $\frac{1}{2}$ reaction with e^- s
- Add H_2O , then H^+ and balance accordingly
- Balance for e^- s and everything else
- Add together both balanced $\frac{1}{2}$ reactions and cancel out where possible to simplify

28.2 Basic

- Find oxidation number
- Write $\frac{1}{2}$ reaction with e^- s
- Add H_2O , then H^+ and balance accordingly
- Add OH amounts to both sides equal to the number of H^+
- Cancel out the H^+ with the OH to form H_2O
- Move all H_2O to one side
- Balance for e^- s and everything else
- Add together both balanced $\frac{1}{2}$ reactions and cancel out where possibly to simplify

Example
$\overset{+7}{Mn} \overset{-2}{O}_4 + \overset{+3}{C} \overset{-2}{O}_4 \rightarrow \overset{+4}{Mn} \overset{-2}{O}_2 + \overset{+4}{C} \overset{-2}{O}_3$
$(4OH + C_2O_4 \rightarrow 2CO_3 + 2e^- + 2H_2O) * 3$
$(2H_2O + MnO_4 + 3e^- \rightarrow MnO_2 + 4OH) * 2$
$12OH + 3C_2O_4 \rightarrow 6CO_3 + 6e^- + 6H_2O$
$4H_2O + 2MnO_4 + 6e^- \rightarrow 2MnO_2 + 8OH$
$4OH + 3C_2O_4 + 2MnO_4 \rightarrow 2MnO_2 + 6CO_3^{-2}$

29 Thermodynamics

The study of energy and its transformations

Units of Energy:

- Joules and Calories
- ▷ 1 cal = 4.184 J

The two main driving forces of thermodynamics is **Enthalpy** and **Entropy**:

29.1 Enthalpy

Enthalpy stands for the **Heat of the reaction** and is denoted by ΔH .

If:

- $\Delta H < 0$
 - ▷ Reaction is exothermic
- $\Delta H > 0$
 - ▷ Reaction is endothermic

There are 4 ways to find ΔH .

29.1.1 Stoichiometry Problems

Example
How much heat is released when 3.2 grams of Hydrogen is reacted with excess Oxygen?
$2H_2 + O_2 \rightarrow 2H_2O$
$\Delta H^\circ = -572 \text{ KJ}$
$\frac{3.2 \text{ g } H_2}{1} * \frac{1 \text{ mole } H_2}{2.02 \text{ g } H_2} * \frac{-572 \text{ KJ}}{2 \text{ mole } H_2} = -453.069 \text{ KJ}$
$Ratio = \frac{\text{Energy Released}}{\text{Coefficient of Hydrogen in formula}}$

29.1.2 Calorimetry

Find the ΔH by running a reaction and heating or cooling a substance.

$$q = m * c * \Delta T$$

- q = Heat released or absorbed
- m = Mass of what is being heated (grams)
- c = Specific heat. Unique to every substance ($\frac{J}{g \cdot ^\circ C}$)
 ▷ Specific heat of water is 4.184
- ΔT = Change in temperature

Example
Burn 0.1 grams of CH_4 and it heats 100 grams H_2O from $20^\circ C$ to $33.29^\circ C$.
$q = 100 * 4.184 * 13.29 = 5560 J = 5.560 KJ$
$\frac{0.1 \text{ grams } CH_4}{1} * \frac{1 \text{ mole } CH_4}{16 \text{ g } CH_4} = 0.00625 \text{ moles } CH_4$
$\frac{5.560}{0.00625} = 889.6 \frac{KJ}{Mole}$

29.1.3 Hess Law

Multiple reactions can be added together then ΔH s can be added together.

Example
$Si + 2H_2 \rightarrow SiH_4 \quad \Delta H = +34 \frac{KJ}{Mole}$
$Si + O_2 \rightarrow SiO_2 \quad \Delta H = -911 \frac{KJ}{Mole}$
$H_2 + \frac{1}{2}O_2 \rightarrow H_2O \quad \Delta H = -242 \frac{KJ}{Mole}$
Find ΔH for:
$SiH_4 + 2O_2 \rightarrow SiO_2 + 2H_2O$
$SiH_4 \rightarrow Si + 2H_2 \quad \Delta H = -34 \frac{KJ}{Mole}$
$Si + O_2 \rightarrow SiO_2 \quad \Delta H = -911 \frac{KJ}{Mole}$
$2H_2 + 2O_2 \rightarrow 2H_2O \quad \Delta H = -484 \frac{KJ}{Mole}$
$SiH_4 + 2O_2 \rightarrow SiO_2 + 2H_2O \quad \Delta H = -1429 \frac{KJ}{Mole}$

29.1.4 Standard Heat of Formation

Standard heat (enthalpy) of formation (ΔH_f°)² is the energy involved in forming **one** mole of a chemical from its elements under standard conditions.

† Elemental substances (O_2, H_2 , etc.) always have a ΔH of **zero**.

Example
Find the ΔH for:
$2H_2O_2 \rightarrow 2H_2O + O_2$
$\Delta H_f H_2O_2 = -187$
$\Delta H_f H_2O = -285$
$\underbrace{2*(-187)}_{-374} \quad \underbrace{2*(-285)}_{-570}$
$\Delta H = \sum product - \sum reactant$
$\Delta H = -570 - (-374) = -196 \frac{KJ}{Mole}$

²This symbol may be shortened to ΔH or ΔH_f in this subsection.

29.2 Entropy

Entropy stands for the **Disorder of the reaction** and is denoted by ΔS .

If:

- $\Delta S < 0$
▷ Order is increasing
- $\Delta S > 0$
▷ Disorder is increasing

29.2.1 State of Matter

If:

- Solid \rightarrow Liquid
▷ $\Delta S = +$
- Gas \rightarrow Solid
▷ $\Delta S = -$

Solid	Liquid	Gas
Lowest ΔS	–	Highest ΔS

29.2.2 Number of Moles of Gasses

⊢ Solids and liquids do not apply

Example	
$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$	
4 moles gas \rightarrow 2 moles gas	$\Delta S = -$
2 moles gas \rightarrow 4 moles gas	$\Delta S = +$

29.2.3 Pressure of Gas

- When pressure **increases**, disorder **decreases**.
- When pressure **decreases**, disorder **increases**.

Example
What has more disorder?
N_2 at 1 atm
N_2 at 0.001 atm
Answer: N_2 at 0.001 atm

29.3 Gibbs Law of Free Energy

Gibbs Law determines ΔG which signifies whether a reaction is spontaneous or not.

$$\Delta G = \Delta H - (T * \Delta S)$$

- ΔG = Free Energy in a system
- ΔH = Enthalpy (KJ)
- ΔS = Entropy ($\frac{KJ}{K}$)
▷ ΔS **MUST** be converted from $\frac{J}{K}$ to $\frac{KJ}{K}$.
- T = Temperature in Kelvin
▷ To convert $C^\circ \rightarrow K$ add 273

If:

- $\Delta G < 0$
▷ Spontaneous
- $\Delta G > 0$
▷ Not spontaneous

29.3.1 ΔH , ΔS , ΔG , Relationship Table

$\Delta H = -$	$\Delta S = +$	$\Delta G = -$	Always spontaneous
$\Delta H = -$	$\Delta S = -$	$\Delta G = -$	Spontaneous at low temperatures
$\Delta H = +$	$\Delta S = +$	$\Delta G = +$	Spontaneous at high temperatures
$\Delta H = +$	$\Delta S = -$	$\Delta G = +$	Never spontaneous

Example
$POCl_3 \rightarrow 2PCl_3 + O_2$
$\Delta H = 542 \text{ KJ}$
$\Delta S = 179 \frac{J}{K}$
What temperature is it spontaneous at?
$0 = 542 \text{ KJ} - (T * 0.179 \frac{KJ}{K})$ $-542 = -0.179T$ $T = 3027.93 \text{ K}$

30 Chemical Kinetics and Rate Laws

Factors that affect reaction rates

30.1 Physical State

- Solid
▷ An **increase** in surface area means in an **increase** in the rate.
- Gas - Gas
- Liquid - Gas
- Liquid - Liquid

30.2 Concentration

$$^3\text{Molarity} = \frac{\text{moles}}{\text{liter}}$$

$$[HCL] = 3M$$

An **increase** in concentration is generally an **increase** in rate.

30.3 Temperature

An **increase** in temperature is an **increase** in rate.

30.4 Pressure of Gas

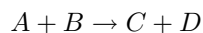
An **increase** in pressure is an *bincrease* in rate.

30.5 Catalysts and Inhibitors

A catalyst lowers the activation energy while an inhibitor increases the activation energy.

³Molarity is signified by []s

30.6 Rate Laws



$$rate = k[A]^m[B]^n$$

- k = Constant
- m = Order of A
- n = Order of B

⊢ Order of 0 → No effect

⊢ Order of 1 → Linear - Double the concentration and you double the rate

⊢ Order of 2 → Squared - Double the concentration and you quadruple the rate

Example:

Trial	[A]	[B]	Rate
1	0.1 M	0.1 M	0.04 M/s
2	0.2 M	0.1 M	0.08 M/s
3	0.1 M	0.2 M	0.04 M/s

Solve for m:

$$\frac{trial\ 2}{trial\ 1} = \left(\frac{[A]}{[A]}\right)^m = \frac{rate}{rate} = \left(\frac{0.2}{0.1}\right)^m = \frac{0.08}{0.04}$$

$$2^m = 2$$

$$m = 1$$

Solve for n:

$$\left(\frac{0.2}{0.1}\right)^n = \frac{0.04}{0.04}$$

$$1^n = 1$$

$$n = 0$$

$$rate = k[A]^1[B]^0$$

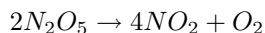
Solve for k:

$$0.04 = k[0.1]^1[0.1]^0$$

$$k = 0.4$$

30.6.1 Order Table

Comments	Zero Order	First Order	Second Order
Rate Law	$rate = k$	$rate = k[A]^1$	$rate = k[A]^2$
Integrated Rate law	$[A] - [A]_0 = -kt$ $[A] = -kt + [A]_0$	$\ln[A] - \ln[A]_0 = -kt$ $\ln[A] = -kt + \ln[A]_0$	$\frac{1}{[A]} - \frac{1}{[A]_0} = kt$ $\frac{1}{[A]} = kt + \frac{1}{[A]_0}$
Graph	[A] vs Time	ln[A] vs time	$\frac{1}{[A]}$ vs time
K = Slope	Slope = $-k$	Slope = $-k$	Slope = k
Half-Life ($t_{\frac{1}{2}}$)	$t_{\frac{1}{2}} = \frac{[A]_0}{2k}$	$t_{\frac{1}{2}} = \frac{0.693}{k}$	$t_{\frac{1}{2}} = \frac{1}{k[A]_0}$

Example:

$[N_2O_5]$	Time (s)
0.1	0
0.0707	50
0.05	100
0.025	200
0.0125	300
0.00625	400

1. What is the order of the reaction?

$$[A] \neq \text{straight}$$

$$\frac{1}{[A]} \neq \text{straight}$$

$$\ln[A] = \text{straight}$$

Order of 1

2. What is the k constant value?

$$\frac{\ln(0.0707) - \ln(0.1)}{50 - 0} = \frac{-0.347}{50} = 0.00693$$

$$k = 0.00693$$

3. What is the concentration of N_2O_5 at $t = 150$?

$$\ln[A] = -(0.00693)(150) + \ln(0.1)$$

$$\ln[A] = -3.34$$

$$[A] = 0.0354 \text{ M}$$

4. What is the rate at 150 seconds?

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

$$\text{rate} = 0.00693 * [0.0354]$$

$$\text{rate} = 2.45 * 10^{-4} \text{ M/s}$$

5. What is the half life?

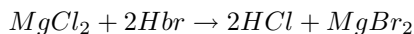
$$t_{\frac{1}{2}} = \frac{0.693}{k}$$

$$t_{\frac{1}{2}} = \frac{0.693}{0.00693}$$

$$t_{\frac{1}{2}} = 100 \text{ s}$$

31 Reaction Mechanisms

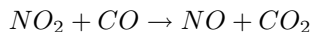
Many/most reactions do not take place in one step. If a reaction were to react in one step, then you could use the balanced reaction to determine the rate law. For example, assume the following occurred in one step.



$$\text{rate} = k[MgCl_2]^1[HBr]^2$$

In reality though, things are not always as easy.

Through experimentation we figure out that the rate law for:



is

$$\text{rate} = k[NO_2]^2$$

Because the rate law does not link up with the equation, it is not a single step reaction.

31.1 Elementary Steps

- Unimolecular - 1 reactant
- Bimolecular - 2 reactants
- Termolecular - 3 reactants

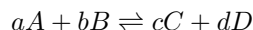
32 Equilibrium

The state where the concentration or partial pressures (if it is a gas) of all the reactants and products remain constant with time. For equilibrium to occur, the forward reaction rate must equal the reverse rate. In other words, the amounts do not have to be equal, but the rates must be.

32.1 Types of Equilibrium

- Static \rightarrow No movement
- Dynamic \rightarrow Movement such as a sealed container of water

32.2 Equilibrium Constant Expressions



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

$$K_p = \frac{(PC^c)(PD^d)}{(PA^a)(PB^b)}$$

- K_c = Concentration constant
- K_p = Partial Pressure constant

32.2.1 Converting Constants

To convert between the two constants K_c and K_p use the formula:

$$K_p = K_c(RT)^{\Delta n}$$

- $\Delta n = \sum \text{Product Coefficients} - \sum \text{Reactant Coefficients}$

33 Gas Laws

33.1 Gas Units and Conversions

$$1 \text{ Atm} = 760 \text{ Torr (mmHg)} = 101.3 \text{ kPa} = 14.7 \text{ PSI}$$

33.2 Ideal Gas Law

$$Pv = nRT$$

- P = Pressure (Atm)
- v = Volume (L)
- n = Number of moles
- R = 0.0821 (constant)
- T = Temperature (Kelvin)

Example	
3 grams of HCl at $26^\circ C$ in a 3 Liter container. What is the pressure?	
$\frac{P(3)}{3} = \frac{(\frac{3 \text{ grams}}{36.5 \text{ g/mole}})(0.081)(26+273)}{3}$ $P = 0.0664 \text{ Atm}$	

33.3 Real Gas Law

For use when the ideal gas law fails. The ideal gas law fails when these two postulates fail:

- Molecules do have volume
- Molecules are attracted

This law is also used when there are conditions with **high pressure** and **low temperature**.

$$Pv = nRT$$

↓

$$\left(P + \frac{n^2 a}{v^2}\right) * (v - nb) = nRT$$

- a = constant that fixes the **intermolecular force issue**
- b = constant that fixes the **volume issue**

⊢ a and b are unique to each type of gas

⊢ All other variables are the same as the ideal gas law

33.4 Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

33.5 Daltons Law of Partial Pressures

For a mixture of gases in a container, the total pressure (P_{tot}) is equal to the sum of the pressures each gas exerts as if it were alone.

$$P_{tot} = P_1 + P_2 + P_3 + \cdots + P_n$$

Example:

A mixture of 1g H and 1g He in a 1 L container is at $27^\circ C$. Calculate the mole fraction of each gas, partial pressures of each and total pressure.

$$H_2 \ 1 \ g * \frac{1}{2} = 0.5 \ moles$$

↓

$$x = \frac{(H \ mole)}{(H + He \ moles)} = \frac{0.5}{0.75} = 0.667$$

$$He \ 1 \ g * \frac{1}{4} = 0.25 \ moles$$

↓

$$x = \frac{0.25}{0.75} = 0.333$$

H_2	He
$Pv = nRT$	$Pv = nRT$
$P(1) = (0.5)(0.0821)(300)$	$P(1) = (0.25)(0.0821)(300)$
P = 12.3 Atm	P = 6.15 Atm

$$P_{tot} = 12.3 + 6.15 = 18.45 \ Atm$$

33.6 Gas Collection over a Water Solution

Example:

A 0.986 g sample has Zinc and some impurities. Excess HCl is added and reacts with the Zinc but not the impurities. Find the percent Zinc in the sample if 240 mL of H_2 are collected over H_2O at $30^\circ C$ and 1.032 Atm (*HINT: This is the P_{tot}*).

$$\begin{aligned}
 P_{tot} &= P_1 + P_{H_2O} \\
 1.032 &= P_1 + 0.042 \\
 P_1 &= 0.99 \text{ Atm } H_2 \\
 &\Downarrow \\
 Pv &= nRT \\
 (0.99)(0.240) &= n(0.0821)(303) \\
 n &= 0.0096 \text{ mole } H_2 \\
 &\Downarrow \\
 0.0096 \text{ mole } H_2 &= 0.0096 \text{ mole Zn} \\
 \frac{0.0096 \text{ mole Zn}}{1} * \frac{65.4 \text{ g Zn}}{\text{Mole}} &= 0.628 \text{ g Zn} \\
 \frac{0.628 \text{ g Zn}}{0.986 \text{ g total}} * 100 &= 63.7\% \text{ Zn}
 \end{aligned}$$

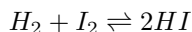
34 ICE ICE (Baby)

Given initial values for a system at equilibrium and one of the equilibrium values, you should find:

- a - The other equilibrium values
- b - The equilibrium constant

Example:

A closed system initially containing $1 * 10^{-3} M H_2$ and $2 * 10^{-3} M I_2$ at $448^\circ C$ is allowed to reach equilibrium. Analysis of the equilibrium mixture shows the $[HI] = 1.7 * 10^{-3} M$. Find the equilibrium concentration for H_2 and I_2 as well as the K_c value.



Initial	$1 * 10^{-3} M$	$2 * 10^{-3} M$	$0 M$
Change	$-0.935 * 10^{-3}$	$0.935 * 10^{-3}$	$1.87 * 10^{-3}$
Equilibrium	$0.065 * 10^{-3}$	$1.065 * 10^{-3}$	$1.87 * 10^{-3}$

$$\begin{aligned}
 [H_2] &= 0.065 M \\
 [I_2] &= 1.065 * 10^{-3} M \\
 K_c &= \frac{[1.87 * 10^{-3}]^2}{[1.065 * 10^{-3}][0.065 * 10^{-3}]}
 \end{aligned}$$

35 Acids and Bases

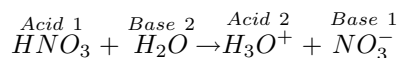
35.1 Definitions of Acids and Bases

1. Arrhenius

▷ An acid dissociates in water to form H^+ ions and a base dissociates to form OH^- ions.

2. Bronsted-Lowry

- ▷ Acids are proton donors (H^+) and a base is a proton acceptor.
 ▷ Conjugate acid base pair



3. Lewis Acid

▷ Acid is an e^- pair acceptor while a base is an e^- pair donor.

35.2 pH and pOH

pH and pOH are measures of the amount of ions in a solution that either cause the solution to be acidic or basic.

pH Scale	
Basic \Rightarrow 0 \leftrightarrow 14 \Leftarrow Acidic	
Important Formulas	
$pH = -\log[H^+]$	
$pOH = -\log[OH^-]$	
$pH + pOH = 14$	
$[H^+] = 1 * 10^{-pH}$	
$[OH^-] = 1 * 10^{-pOH}$	
Example	
What is the concentration of HCl with a pH of 3?	
$[HCl] = 0.001 \text{ M}$	

35.2.1 Changing Concentrations

$$M_1 V_1 = M_2 V_2$$

$$(0.25 \text{ M})(5 \text{ mL}) = M_2(50 \text{ mL})$$

$$M_2 = 0.025 \text{ M}$$

35.3 Strong Acids and Bases

Strong acids and bases completely dissociate in water.

35.3.1 Strong Acids

- HCl
- H_2SO_4
- HBr
- HI
- HNO_3
- $HClO_4$

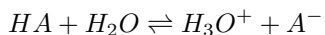
35.3.2 Strong Bases

- Group 1 - Hydroxides
 - ▷ $NaOH$
 - ▷ KOH
- Group 2 - Heavier Hydroxides
 - ▷ $Ca(OH)_2$
 - ▷ $Sr(OH)_2$
 - ▷ $Ra(OH)_2$

35.4 Weak Acids and Bases

Weak acids and bases do not completely dissociate in water.

35.4.1 K_a Constant



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

Example
Benzoic acid dissociates as follows:
$HC_7H_5O_2 \rightleftharpoons \underbrace{H^+ + C_7H_5O_2^-}_{x^2}$
$[HC_7H_5O_2] = 0.4 \text{ M}$ $K_a = 6.3 * 10^{-5}$
What is the pH?
$K_a = \frac{[H^+][C_7H_5O_2^-]}{[HC_7H_5O_2]}$ $6.3 * 10^{-5} = \frac{x^2}{0.4}$

35.4.2 K_b Constant

The K_b constant is used when bases are involved in a reaction (as opposed to K_a which is used in reactions with acids). To convert between K_b and K_a use the following formula:

$$K_a * K_b = K_w$$

- $K_w = 1 * 10^{-14}$

Example
$F^- + H_2O \rightleftharpoons HF + OH^-$
$K_a = 7.2 * 10^{-4}$
What is the K_b constant?
$K_b = \frac{1 * 10^{-14}}{7.2 * 10^{-4}} = 1.39 * 10^{-11}$
Find the pH and pOH.
$K_b = \frac{[HF][OH^-]}{[F^-]}$ $1.39 * 10^{-11} = \frac{x^2}{(\frac{0.002}{20+13.3})}$ $x = 9.13 * 10^{-7}$ $pOH = 6.04$ $pH = 7.96$

35.5 Common Ion Effect

The effect of ionization of a weak electrolyte (acid/base) is **decreased** by adding a strong electrolyte that has an ion in common with the weak electrolyte.

35.6 Buffer

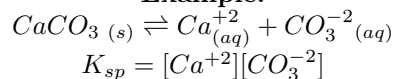
Made of 2 components:

1. Weak acid
2. The salt of that acid

36 Equilibrium of Saturated, Soluble Salts

Solubility is how well a solute dissolves in a solvent⁴.

Example:



- K_{sp} is the solubility product
 - ▷ A large K_{sp} means the solution is very soluble (meaning lots of products)
 - ▷ A small K_{sp} means the solution is not very soluble.

1. Given K_{sp} , find the ion concentration.

$$K_{sp} = [Ca^{+2}][CO_3^{-2}] = 4.5 * 10^{-9}$$

$$[Ca^{+2}] = [CO_3^{-2}] = \sqrt{4.5 * 10^{-9}} = 6.7 * 10^{-5} M$$

2. Given K_{sp} , find the solubility (g/L).

$$6.7 * 10^{-5} M = \frac{6.7 * 10^{-5}}{1} * \frac{100.1}{1 mole} = 6.37 * 10^{-3} g/L$$

3. Given solubility, find ion concentration.

$$\text{Solubility of Silver Chloride at } 25^\circ C \text{ is } 1.3 * 10^{-7} \frac{g}{100 mL}$$

$$1.3 * 10^{-7} \frac{g}{100 mL} \rightarrow \frac{g}{L} * \frac{10}{10} = 1.3 * 10^{-6} \frac{g}{L}$$

$$\frac{1.3 * 10^{-6}}{L} * \frac{1 mole}{143.35 g} = 9.11 * 10^{-9} \frac{m}{L}$$

4. Given solubility, find K_{sp}

$$K_{sp} = [Ag^+][Cl^-] = (9.11 * 10^{-9})^2 = 8.3 * 10^{-17}$$

37 Kinetic Molecular Theory

37.1 Postulates:

- The volume of the individual particles of a gas can be assumed to be negligible.
 - ▷ So volume is determined by the space between molecules
- The gas particles are in constant motion. The pressure exerted by a gas is due to collisions of the gas with the walls of the container.
- Gas particles are not attracted to one another.
- The average kinetic energy of a gas is directly proportional to the Kelvin temperature.

$$K_{energy} = \frac{3}{2}(0.0821)T$$

OR

$$K_{energy} = \frac{1}{2}(Molar Mass)(Velocity)^2$$

A) CO at 760 torr and 0°C

B) N₂ at 760 torr and 0°C

C) H₂ at 760 torr and 0°C

Q. Which will have the highest kinetic energy?

A. All will have the same kinetic energy

Q. Which will have a higher velocity?

A. H₂ will because if all kinetic energies are constant according to the formula $k = \frac{1}{2}mv^2$ the smallest mass will yield the highest velocity to keep k constant.

⁴Virtually every salt is soluble to some degree.

37.2 Root Mean Square Velocity

$$U_{rms} = \sqrt{\frac{3RT}{M}}$$

- $R = 8.314 \frac{J}{K \cdot Mole}$
- M = molar mass (Kg/mole)
▷ **NOT** g/mole

37.3 Effusion and Diffusion

37.3.1 Effusion

When you pass a gas through a small opening into an evacuated chamber.

37.3.2 Diffusion

When you mix gases

37.3.3 Finding the rate

The formula for finding the rate is as follows:

$$\frac{\text{Rate of Effusion of Gas 1}}{\text{Rate of Effusion of Gas 2}} = \frac{\sqrt{(\text{Molar Mass 2})}}{\sqrt{(\text{Molar Mass 1})}}$$

- Which effuses faster, He or NO_2 ?
▷ He - it moves faster because it is smaller
- For the reaction: $H_2 + N_2$ at $20^\circ C$ and having a rate of effusion for H_2 being 10 mL/min what is the rate for N_2 ?
▷ $\frac{10}{x} = \frac{\sqrt{28}}{\sqrt{2}}$
 $x = \frac{10\sqrt{2}}{\sqrt{28}} = 2.67 \text{ mL/min}$
- The rate for the gas is 24 mL/min, at the same temperature methane has a rate of 47.8 mL/min. What is the molar mass of the unknown gas?
▷ $\frac{24}{47.8} = \frac{\sqrt{16}}{\sqrt{x}}$
 $x = 63.7 \text{ g/mole}$

38 Electro Chemistry

38.1 Identifying Oxidation Numbers

H_2O	H_2SO_4	Cl_2
H +1	H +1	Cl 0
O -2	S +6	
	O -2	

The chemical that has been **oxidized** is the **reducing agent**. The chemical that has been **reduced** is the **oxidizing agent**.

38.2 Galvanic/Voltaic Cells

There are two beakers with salt and e^- s in each solution. A salt bridge between the two solutions allows passage of ions. One side is identified as the cathode and the other the anode. The cations go to the cathode and the anions go to the anode. The e^- s go to the cathode.

38.3 Calculating Cell Potential

$$E^{\circ} cell = E^{\circ} reduction + E^{\circ} Oxidation$$

Example
$Cu^{+2} + Zn \rightarrow Zn^{+2} + Cu$
Oxidation $Zn \rightarrow Zn^{+2} + 2e^{-}$ $e^{\circ} = 0.76$
Reduction $Cu^{+2} + 2e^{-} \rightarrow Cu$ $e^{\circ} = 0.34$
$E^{\circ} cell = 0.34 + 0.76 = 1.1 \text{ volts}$

38.3.1 Nernst Equation to Find $E^{\circ} cell$

The Nernst equation to be used under standard conditions is:

$$E^{\circ} cell = E^{\circ} - \frac{0.0592}{n} \log(Q)$$

- E° = Normal standard potential
- n = Number of moles of e^{-} s changing
- Q = Reaction Quotient = $\frac{[Product]}{[Reactant]}$

Example
$Cu_{(s)} + \underbrace{Cu^{+2}}_{1 \text{ } M} \rightarrow \underbrace{Cu^{+2}}_{0.1 \text{ } M} + Cu_{(s)}$
$Q = \frac{0.1}{1} = 0.1$
n = 2 (2 e^{-} being transfered)
$E^{\circ} = 0(1M - 1M = \text{No voltage})$
$E^{\circ} cell = 0 - \frac{0.0592}{2} \log(0.1)$

This Nernst equation is to be used when **the temperature is not standard** and **the concentrations are not equal**.

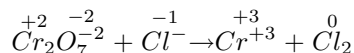
$$E^{\circ} cell = E^{\circ} - \frac{RT}{nF} \ln(Q)$$

- R = 9.31 $\frac{\text{volt coulomb}}{\text{mole Kelvin}}$
- F = 96,500 per mole e^{-}
- T = Temperature in Kelvin

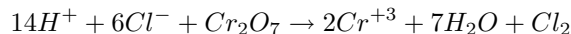
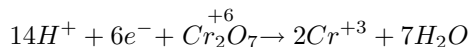
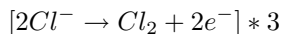
Example
$Zn + \underbrace{Cu^{+2}}_{2 \text{ } M} \rightarrow Cu + \underbrace{Zn^{+2}}_{0.5 \text{ } M}$
$E^{\circ} cell = 1.1 - \frac{0.0592}{2} \log(\frac{0.5}{2})$
$E^{\circ} cell = 1.1178 \text{ volts}$

39 Balancing Redox Reactions

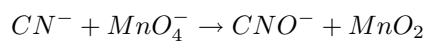
39.1 Acidic



Half-Reactions



39.2 Basic



Half-Reactions

