Periodic Table reference sheet (handwritten annotations)

Constants / Formulas

Avagadro # - $6.022e23$

Speed of Light - $3.00e8$

Plancks Constant - $6.626e-34$

$FC = (Ve^-) - (nonbond\ e^-) - (\tfrac{1}{2}\ bond\ e^-)$

$BO = \tfrac{1}{2}(Bond\ e^- - antibond\ e^-)$

Ionization Energy
$X\ gas \rightarrow e^- + x^+\ gas$

Electron Affinity
$X\ gas + e^- \rightarrow x^-\ gas$

Top left key

A = mass # = pro + neu

A - Z = neutrons

Z = protons = electrons

1A = group #

$\pm x$ = predictable charge

s-block
p-block
d-block
f-block

* p-block will lose electrons from the s + p orbitals to form cations

inner transition elements / rare earth
Lanthanide Series
Actinide Series

Main group elements — s-block
Main group elements — p-block
transition metals — d-block

Atoms or Molecules ⇄ Moles ⇄ Grams
(6.02×10^{23} ; Atomic/Molar Mass)

Legend

Alkali Metal
Alkaline Earth
Transition Metal
Basic Metal
Metalloid
Nonmetal +H
Halogen
Noble Gas
Lanthanide
Actinide

Electron Domain

$AX_2 - AX_6$

Molecular Domain

$AX_{2-6}E$

A = central atom
X = bonded e⁻ group
E = lone pair e⁻ group

Polyatomic Ions

Ion	Name
NH_4^+	Ammonium
CH_3COO^-	Acetate
$C_2H_3O_2^-$	Acetate
BrO_3^-	Bromate
CO_3^{2-}	Carbonate
ClO_2^-	Chlorite
CrO_4^{2-}	Chromate
PO_3^{3-}	Phosphite
SiO_4^{4-}	Silicate
CN^-	Cyanide
$Cr_2O_7^{2-}$	Dichromate
HCO_3^-	Hydrogen Carbonate / Bicarbonate
HSO_4^-	H Sulfate Bisulfate
HPO_4^{2-}	H P Bisphosphate
OH^-	Hydroxide
ClO^-	Hypochlorite
IO_3^-	Iodate
NO_3^-	Nitrate
NO_2^-	Nitrite
$C_2O_4^{2-}$	Oxalate
ClO_4^-	Perchlorate
IO_4^-	Periodate
MnO_4^-	permanganate
O_2^{2-}	peroxide
PO_4^{3-}	Phosphate
SO_4^{2-}	Sulfate
SO_3^{2-}	Sulfite
SCN^-	thiocyanate
$S_2O_3^{2-}$	thiosulfate

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Element	Electrons	Electronic Configuration
Hydrogen (H)	1	$1s^1$
Helium (He)	2	$1s^2$
Lithium (Li)	3	$1s^2\,2s^1$
Beryllium (Be)	4	$1s^2\,2s^2$
Boron (B)	5	$1s^2\,2s^2\,2p^1$
Carbon (C)	6	$1s^2\,2s^2\,2p^2$
Nitrogen (N)	7	$1s^2\,2s^2\,2p^3$
Oxygen (O)	8	$1s^2\,2s^2\,2p^4$
Fluorine (F)	9	$1s^2\,2s^2\,2p^5$
Neon (Ne)	10	$1s^2\,2s^2\,2p^6$
Sodium (Na)	11	$1s^2\,2s^2\,2p^6\,3s^1$
Magnesium (Mg)	12	$1s^2\,2s^2\,2p^6\,3s^2$
Aluminum (Al)	13	$1s^2\,2s^2\,2p^6\,3s^2\,3p^1$
Silicon (Si)	14	$1s^2\,2s^2\,2p^6\,3s^2\,3p^2$
Phosphorous (P)	15	$1s^2\,2s^2\,2p^6\,3s^2\,3p^3$
Sulfur (S)	16	$1s^2\,2s^2\,2p^6\,3s^2\,3p^4$
Chlorine (Cl)	17	$1s^2\,2s^2\,2p^6\,3s^2\,3p^5$
Argon (Ar)	18	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$
Potassium (K)	19	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1$
Calcium (Ca)	20	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2$
Scandium (Sc)	21	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^1$
Titanium (Ti)	22	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^2$
Vanadium (V)	23	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^3$
Chromium (Cr)	24	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^5$
Manganese (Mn)	25	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^5$
Iron (Fe)	26	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^6$
Cobalt (Co)	27	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^7$
Nickel (Ni)	28	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^8$
Copper (Cu)	29	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^1\,3d^{10}$
Zinc (Zn)	30	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}$
Gallium (Ga)	31	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^1$
Germanium (Ge)	32	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^2$
Arsenic (As)	33	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^3$
Selenium (Se)	34	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^4$
Bromine (Br)	35	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^5$
Krypton (Kr)	36	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6$
Rubidium (Rb)	37	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1$
Strontium (Sr)	38	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2$
Yttrium (Y)	39	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^1$
Zirconium (Zr)	40	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^2$
Niobium (Nb)	41	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1\,4d^4$
Molybdenum (Mb)	42	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1\,4d^5$
Technetium (Tc)	43	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^5$
Ruthenium (Ru)	44	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1\,4d^7$
Rhodium (Rh)	45	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1\,4d^8$
Palladium (Pd)	46	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,4d^{10}$
Silver (Ag)	47	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^1\,4d^{10}$
Cadmium (Cd)	48	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}$
Indium (In)	49	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^1$
Tin (Sn)	50	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^2$
Antimony (Sb)	51	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^3$
Tellurium (Te)	52	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^4$
Iodine (I)	53	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^5$
Xenon (Xe)	54	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6$
Cesium (Cs)	55	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^1$
Barium (Ba)	56	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2$
Lanthanum (La)	57	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,5d^1$
Cerium (Ce)	58	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,5d^1\,4f^1$
Osmium (Os)	76	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^6$
Iridium (Ir)	77	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^7$
Platinum (Pt)	78	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^1\,4f^{14}\,5d^9$
Gold (Au)	79	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^1\,4f^{14}\,5d^{10}$
Mercury (Hg)	80	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}$
Thallium (Tl)	81	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^1$
Lead (Pb)	82	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^2$
Bismuth (Bi)	83	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^3$
Polonium (Po)	84	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^4$
Astatine (At)	85	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^5$
Radon (Rn)	86	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6$
Francium (Fr)	87	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^1$
Radium (Ra)	88	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^1$
Actinium (Ac)	89	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^2\,6d^1$
Thorium (Th)	90	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}\,4p^6\,5s^2\,4d^{10}\,5p^6\,6s^2\,4f^{14}\,5d^{10}\,6p^6\,7s^2\,6d^2$

$\lambda = $ wavelength $\qquad c = 3e8\ m/s \qquad \Delta E\,1pho = \dfrac{\text{Total } E}{\#\ \text{Photons}}$ - Photons $\qquad KE = h\nu - \phi$ - KE

$\nu = $ frequency (Hz or s⁻¹) $\qquad E = $ total energy $\qquad \lambda = {}^{h}/mv$ w/ $\lambda = m,\ m = kg,\ v = m/s$ - DeBrogli

$h = 6.626e{-}34 \qquad \phi = $ binding energy $\qquad \Delta x \Delta p = {}^{h}/4\pi$ - Heisenbergs Uncertianty

$\nu = {}^{c}/\lambda \qquad c = \nu\lambda \qquad \lambda = {}^{c}/\nu$ - Frequency, wavelength, light

$E = h\nu \qquad E = {}^{hc}/\lambda \qquad \nu = {}^{E}/h \qquad \lambda = {}^{hc}/E$ - Photo electric

$\#\text{Photons} = \dfrac{\text{Total Energy}}{\text{Energy of 1 photon}}$ - Photons

$\Delta E = -2.178e{-}18\ J \left(\dfrac{1}{n_f^2} - \dfrac{1}{n_i^2} \right)$ - Diffrence in ΔE levels hydrogen

434 nm Blue-violet — 486 nm Blue-green — 657 nm Red; absorption / emission; Bohr model energy levels $n=\infty,\ n=5,\ n=4,\ n=3,\ n=2,\ n=1$

Lyman series; 122 nm, 103 nm, 97 nm, 95 nm, 94 nm; Balmer series 656 nm, 486 nm, 434 nm, 410 nm; Paschen series 1875 nm, 1282 nm, 1094 nm; $n=1, n=2, n=3, n=4, n=5, n=6$

Electromagnetic spectrum: Low frequency / High frequency; Frequency, ν (Hz) $10^4,\ 10^6,\ 10^8,\ 10^{10},\ 10^{14},\ 10^{16},\ 10^{18},\ 10^{20},\ 10^{22},\ 10^{24}$; Radio (AM TV FM Cell), Microwave, Infrared, Visible light, Ultraviolet, X-ray, Gamma ray; Low energy / High energy; Long wavelength / Short wavelength; Wavelength, λ (m) $10^5,\ 10^3,\ 10,\ 10^{-1},\ 10^{-3},\ 10^{-5},\ 10^{-7},\ 10^{-9},\ 10^{-11},\ 10^{-13},\ 10^{-15}$; Wavelength, λ (nm) 750 Red, 700, 650, 600, 550, 500, 450, 400 Violet

Ionization Energy	Name chemical element		
3,8939	Cesium	6,7589	Niobium
4,0727	Francium	6,7665	Chromium
4,1771	Rubidium	6,8251	Hafnium
4,3407	Potassium	6,8281	Titanium
4,9	Lawrencium	7,0924	Molybdenum
5,1391	Sodium	7,28	Technetium
5,17	Actinium	7,2856	Bismuth
5,2117	Barium	7,3439	Tin
5,2784	Radium	7,3605	Ruthenium
5,3917	Lithium	7,4167	Lead
5,4259	Lutetium	7,434	Manganese
5,473	Praseodymium	7,4589	Rhodium
5,525	Neodymium	7,5496	Tantalum
5,5387	Cerium	7,5762	Silver
5,5769	Lanthanum	7,6398	Nickel
5,582	Promethium	7,6462	Magnesium
5,6437	Samarium	7,7264	Copper
5,6704	Europium	7,8335	Rhenium
5,6949	Strontium	7,864	Tungsten
5,7864	Indium	7,881	Cobalt
5,8638	Terbium	7,8994	Germanium
5,89	Protactinium	7,9024	Iron
5,9389	Dysprosium	8,1517	Silicon
5,9738	Americium	8,298	Boron
5,9858	Aluminum	8,3369	Palladium
5,9915	Curium	8,417	Polonium
5,9993	Gallium	8,4382	Osmium
6,0215	Holmium	8,6084	Antimony
6,0262	Plutonium	8,9587	Platinum
6,1077	Erbium	8,967	Iridium
6,1082	Thallium	8,9938	Cadmium
6,1132	Calcium	9,0096	Tellurium
6,1501	Gadolinium	9,2255	Gold
6,1843	Thulium	9,3	Astatine
6,1941	Uranium	9,3227	Beryllium
6,1979	Berkelium	9,3942	Zinc
6,2173	Yttrium	9,7524	Selenium
6,2542	Ytterbium	9,7886	Arsenic
6,2657	Neptunium	10,36	Sulfur
6,2817	Californium	10,4375	Mercury
6,3067	Thorium	10,4513	Iodine
6,42	Einsteinium	10,4867	Phosphorus
6,5	Fermium	10,7485	Radon
6,5615	Scandium	11,2603	Carbon
6,58	Mendelevium	11,8138	Bromine
6,6339	Zirconium	12,1298	Xenon
6,65	Nobelium	12,9676	Chlorine
6,7462	Vanadium	13,5984	Hydrogen
		13,6181	Oxygen
		13,9996	Krypton
		14,5341	Nitrogen
		15,7596	Argon
		17,4228	Fluorine
		21,5645	Neon
		24,5874	Helium

Ionization Energy → **Effective Nuclear Charge**

Electronegativity →

more neg (Electron Affinity) ↑   big (Atomic Radius) ↓

- Metals
- Semimetals
- Nonmetals

Metallic Character ←   big

Nonmetallic Character →

Electronegativity ↑   Ionization Energy ↑

Atomic Radius ← big

Electron Affinity → more neg
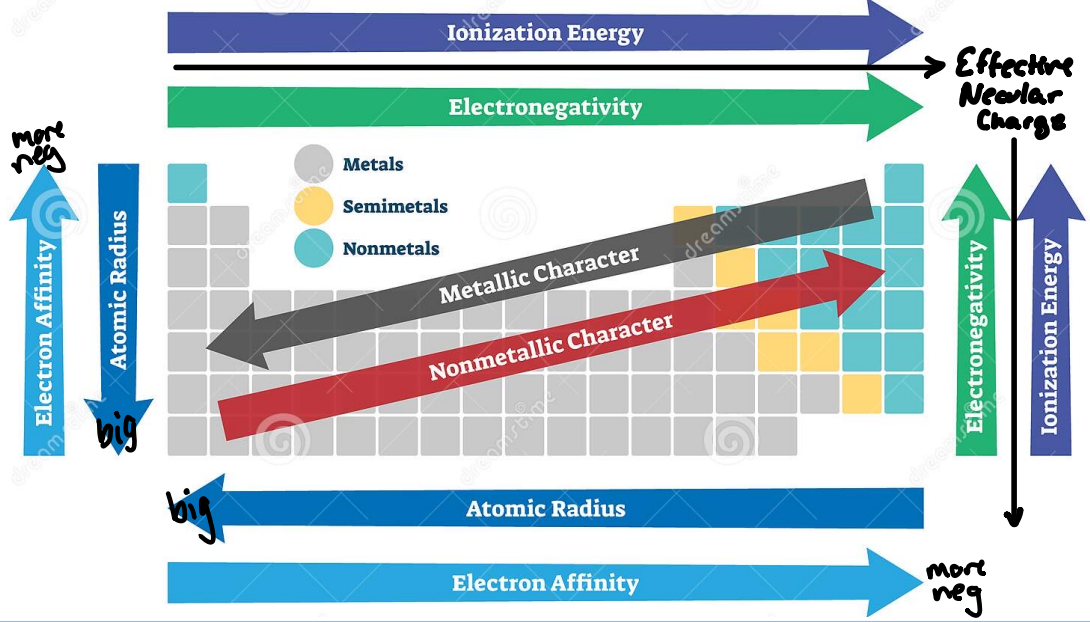
Principal Quantum Number, *n*: The Energy Level

It characterizes the energy of the electron in a particular orbital.
– It is Bohr's energy level.

Values of *n* can be any whole number integer > 1.

It determines the size (overall) and energy of an orbital.
– The larger the value of *n*, the more energy the orbital has.
– The larger the value of *n*, the larger the orbital.

Energies are defined as being "negative."
– An electron's energy is lowered (made more negative) as a result of its interaction with the nucleus of the atom.
  • An electron would have $E = 0$ when it escapes the atom.

– As *n* gets larger, the following occurs:
  • The amount of energy between orbitals gets smaller.
  • The energy of the orbital becomes greater (less negative).

$$E_n = -2.18 \times 10^{-18} \text{ J } (1/n^2) \quad \text{where } n = 1, 2, 3, \ldots.$$

Angular Momentum Quantum Number, *l*: The Orbital Quantum Number

- The angular momentum quantum number (*l*) determines the shape of the orbital.
- *l* can have integer values from 0 to $(n-1)$. $n = 2 \rightarrow l = 0, 1$
- Each *l* value is designated by a particular letter that designates the shape of the orbital.
  – *s* orbitals are spherical.
    • Have *l* value of zero (0)
  – *p* orbitals are like two balloons tied at the knots (dumbbell shape).
    • Have *l* value of one (1)
  – *d* orbitals are mainly like a flat four-leaf clover.
    • Have *l* value of two (2)
  – *f* orbitals are mainly like eight balloons tied at the knots.
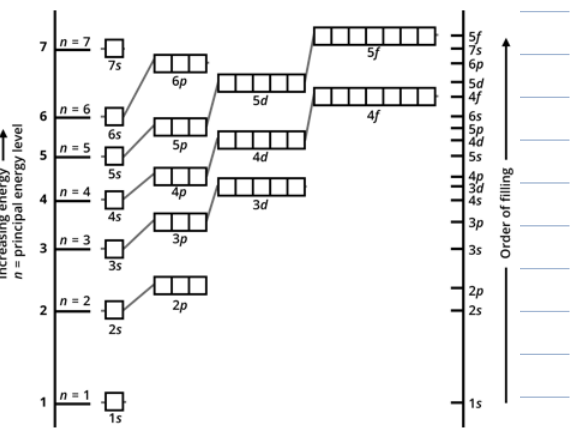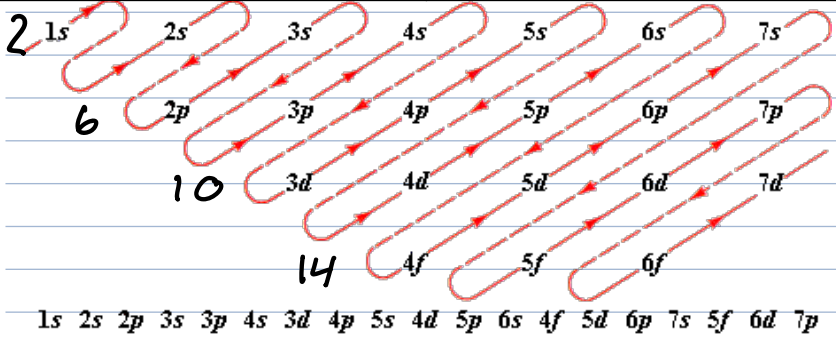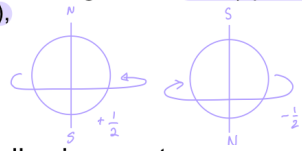   • Have *l* value of three (3)

-g - no known elements use it

Magnetic Quantum Number, *m*ₗ: The Position or Orientation Quantum Number
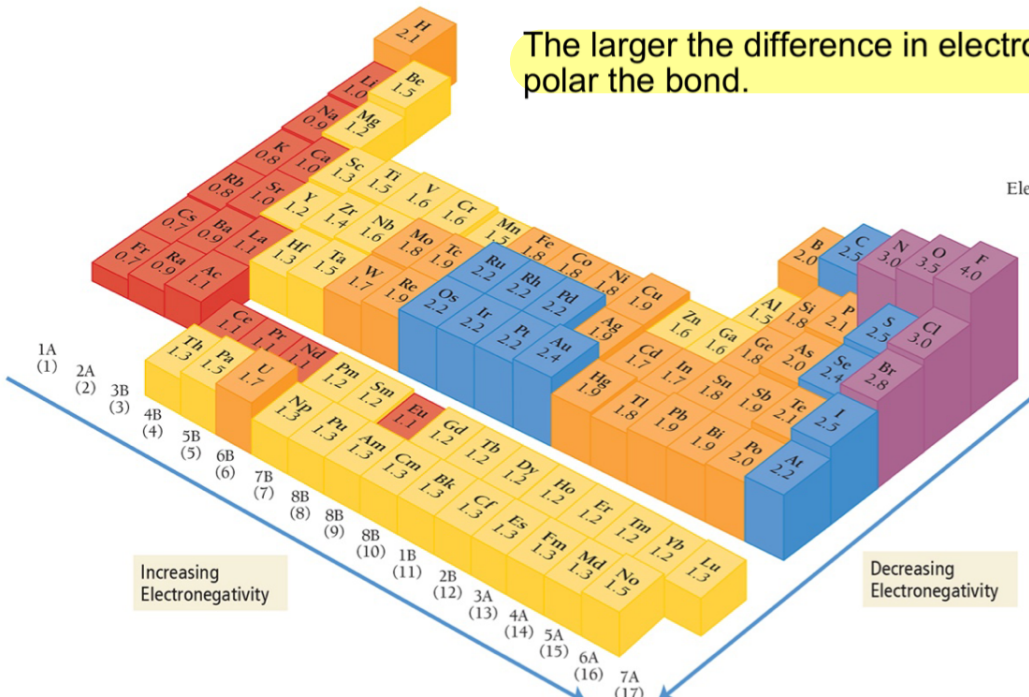
- The magnetic quantum number is an integer that specifies the orientation of the orbital.
  – The direction in space the orbital is aligned relative to the other orbitals.

- Values are integers from $-l$ to $+l$.
  – Including zero
  – Gives the number of orbitals of a particular shape
    • When $l = 2$, the values of m_l are $-2, -1, 0, +1, +2$, which means there are five orbitals with $l = 2$.
    sub shell d

Spin Quantum Number, *m*ₛ

- The spin quantum number describes the spin behavior (i.e., angular momentum) of an electron in an atomic orbital.
  – It specifies the orientation of the electron in the orbital.
    • All electron spins are equal in magnitude; differ only in orientation

- The numerical values assigned are $+ \frac{1}{2}$ (spin up) or $- \frac{1}{2}$ (spin down),

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- In general, the following are true:
  – The number of sublevels within a level = *n*.
  – The number of orbitals within a sublevel = $2l + 1$.
  – The number of orbitals in a level = n^2.

2   1s   2s   3s   4s   5s   6s   7s
6   2p   3p   4p   5p   6p   7p
10 3d   4d   5d   6d   7d
14 4f   5f   6f

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p

Increasing energy / n = principal energy level

Order of filling

The larger the difference in electronegativity, the more polar the bond.

Electronegativity
- 0.7–1.1
- 1.2–1.6
- 1.7–2.1
- 2.2–2.6
- 2.7–4.0

Electronegativity Difference (ΔEN)	Bond Type
Small (0–0.4)	Covalent
Intermediate (0.4–2.0)	Polar covalent
Large (2.0+)	Ionic

Increasing Electronegativity

Decreasing Electronegativity

- polar molecules are unequal
- nonpolar molecules are equal

Electron Groups	Bonding Groups	Lone Pairs	Electron Geometry	Molecular Geometry	Approximate Bond Angles
2	2	0	Linear	Linear	180°
3	3	0	Trigonal planar	Trigonal planar	120°
3	2	1	Trigonal planar	Bent	<120°
4	4	0	Tetrahedral	Tetrahedral	109.5°
4	3	1	Tetrahedral	Trigonal pyramidal	<109.5°
4	2	2	Tetrahedral	Bent	<<109.5°
5	5	0	Trigonal bipyramidal	Trigonal bipyramidal	120° (equatorial) 90° (axial)
5	4	1	Trigonal bipyramidal	Seesaw	<120° (equatorial) <90° (axial)
5	3	2	Trigonal bipyramidal	T-shaped	<90°
5	2	3	Trigonal bipyramidal	Linear	180°
6	6	0	Octahedral	Octahedral	90°
6	5	1	Octahedral	Square pyramidal	<90°
6	4	2	Octahedral	Square planar	90°

Hybridization

# of e⁻ groups	hybridization
2	sp
3	sp^2
4	sp^3
5	sp^3d
6	sp^3d^2

Bond Strength

- multiple bonds are stronger single < double < tripple
- shorter bonds are stronger
- higher BO → shorter bond → stronger bond

Pi & Sigma

- Double bond is 1pi
- Tripple bond is 2pi
- All bonds are 1 sigma

Molecular Orbital Diagrams

Top left (H₂ molecule): Atomic orbital | Molecular orbitals | Atomic orbital — Energy, $1s$ H atom, σ_{1s}^* Antibonding, Bonding σ_{1s}, $1s$ H atom, H₂ molecule

Top middle (He₂⁺ ion): Atomic orbital | Molecular orbitals | Atomic orbital — Energy, $1s$ He atom, σ_{1s}^* Antibonding, Bonding σ_{1s}, $1s$ He⁺ ion, He₂⁺ ion

Top right (H₂⁻ molecule): Energy, $1s$ H atom, σ_{1s}^*, σ_{1s}, $1s$ H⁻ ion, H₂⁻ molecule

$$\text{H}_2^- \text{ bond order} = \frac{2-1}{2} = +\frac{1}{2}$$

Left (Li₂ molecule): Atomic orbital | Molecular orbitals | Atomic orbital — Energy, $2s$ Li atom, σ_{2s}^* Antibonding, Bonding σ_{2s}, $2s$ Li atom, Li₂ molecule

Li₂ molecule

$$BO = \tfrac{1}{2}\left(\text{Bond } e^- - \text{antibond } e^-\right)$$

Percent Yeild $= \dfrac{\text{actual}}{\text{theor}} \cdot 100$

Molarity $= \dfrac{\text{solute in mL}}{\text{volume of solution L}}$

n = moles
M = n/L
n = ML
L = n/M

	B₂	C₂	N₂		O₂	F₂	Ne₂
σ_{2p}^*				σ_{2p}^*			↑↓
π_{2p}^*				π_{2p}^*	↑ ↑	↑↓ ↑↓	↑↓ ↑↓
σ_{2p}			↑↓	π_{2p}	↑↓ ↑↓	↑↓ ↑↓	↑↓ ↑↓
π_{2p}	↑ ↑	↑↓ ↑↓	↑↓ ↑↓	σ_{2p}	↑↓	↑↓	↑↓
σ_{2s}^*	↑↓	↑↓	↑↓	σ_{2s}^*	↑↓	↑↓	↑↓
σ_{2s}	↑↓	↑↓	↑↓	σ_{2s}	↑↓	↑↓	↑↓
BO	1	2	3		2	1	0
BE	290	620	946		498	159	—
BL	159	131	110		121	143	—

Solution Dilution $= M_c V_c = M_d V_d$ * might not be total volumes

c = concentrate   M = molarity
d = dilute   V = volume

Compounds Containing the Following Ions Are Generally Soluble	Exceptions
Li⁺, Na⁺, K⁺, and NH₄⁺	None
NO₃⁻ and C₂H₃O₂⁻	None
Cl⁻, Br⁻, and I⁻	When these ions pair with Ag⁺, Hg₂²⁺, or Pb²⁺, the resulting compounds are insoluble.
SO₄²⁻	When SO₄²⁻ pairs with Sr²⁺, Ba²⁺, Pb²⁺, Ag⁺, or Ca²⁺, the resulting compound is insoluble.

Compounds Containing the Following Ions Are Generally Insoluble	Exceptions
OH⁻ and S²⁻	When these ions pair with Li⁺, Na⁺, K⁺, or NH₄⁺, the resulting compounds are soluble.
	When S²⁻ pairs with Ca²⁺, Sr²⁺, or Ba²⁺, the resulting compound is soluble.
	When OH⁻ pairs with Ca²⁺, Sr²⁺, or Ba²⁺, the resulting compound is slightly soluble.
CO₃²⁻ and PO₄³⁻	When these ions pair with Li⁺, Na⁺, K⁺, or NH₄⁺, the resulting compounds are soluble.

$NaCl(s) \rightarrow (Na^+(aq) + Cl^-(aq)$ Break down of salts.
$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$
$Al_2(SO_4)_3(s) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$
$Na_3PO_4(s) \rightarrow 3Na^+(aq) + PO_4^{3-}(aq)$
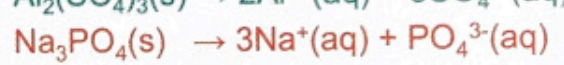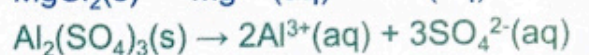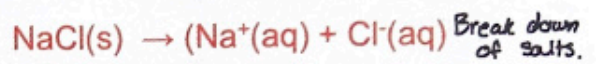
Steps in writing a net ionic equation

- Write the balanced molecular equation.
  - Predict products by exchanging cations and anions in reactants.
- Separate strong electrolytes into ions.
- Cancel spectator ions.
- Use the remaining species to write the net ionic equation.

8.9 Oxidation–Reduction Reactions

- The reactions in which electrons are transferred from one reactant to the other are called **oxidation–reduction** reactions.

   Atoms that **lose electrons** are being **oxidized**.
   - The reaction is referred to as **oxidation**.

   Atoms that **gain electrons** are being **reduced**.
   - The reaction is referred to as **reduction**.

   The following rules are in order of priority:
   1. Free elements have an oxidation state = 0.
      - $Na = 0$ and $Cl_2 = 0$ in $2\,Na(s) + Cl_2(g)$
   2. Monatomic ions have an oxidation state equal to their charge.
      - $Na = +1$ and $Cl = -1$ in $NaCl$
   3. (a) The sum of the oxidation states of all the atoms in a compound is 0.
      - $Na = +1$ and $Cl = -1$ in $NaCl$, $(+1) + (-1) = 0$

      (b) The sum of the oxidation states of all the atoms in a polyatomic ion equals the charge on the ion.
      - $N = +5$ and $O = -2$ in NO_3^-, $(+5) + 3(-2) = -1$

4. Oxidation state of oxygen is -2, except in the O_2^{2-} ion, it is -1
5. Oxidation state of hydrogen is +1, except in metal hydrides it is -1.
6. Group I metals have an oxidation state of +1 in all their compounds.
   - $Na = +1$ in $NaCl$
7. Group II metals have an oxidation state of +2 in all their compounds.
   - $Mg = +2$ in $MgCl_2$
8. In their compounds, nonmetals have oxidation states according to the table on the right.

Nonmetal	Oxidation State	Example
Fluorine	−1	MgF_2 -1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	−2	CO_2 -2 ox state
Group 7A	−1	CCl_4 -1 ox state
Group 6A	−2	H_2S -2 ox state
Group 5A	−3	NH_3 -3 ox state

- The reactant that reduces an element in another reactant is called the **reducing agent**.
   - The reducing agent contains the element that is oxidized.

- The reactant that oxidizes an element in another reactant is called the **oxidizing agent**.
   - The oxidizing agent contains the element that is reduced.

Redox Reactions

- All the reactions which involve a change in oxidation number of the species involved are redox reactions.
- Double replacement reactions are not redox reactions.

Decomposition reactions
- These reactions are just the opposite of combination reactions.
   a. Hydroxide decompose to give metal oxide and water:
   $$Mg(OH)_2\,(aq) \rightarrow MgO(s) + H_2O(l)$$
   b. Metal carbonates decompose into metal oxide and carbon dioxide
   $$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$$
   c. All metal chlorates when heated decompose into metal chlorides and oxygen gas :
   $$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

Acid–Base Reactions
- When an acid and a base react, the reaction is called a **neutralization reaction**.
   - Moles of acid = Moles of base

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$
| Acid | Base | Water | Salt |

Combination Reactions
a. Metals react with oxygen to give metallic oxides:
   $$Mg(s) + O_2(g) \rightarrow MgO(s)$$
b. Metallic oxides react with water to give hydroxides:
   $$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2$$
c. Metallic oxides react with carbon dioxide to give carbonates:
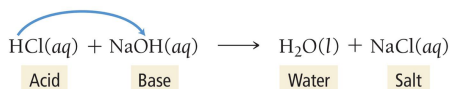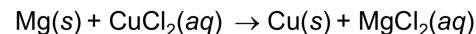 $$MgO(s) + CO_2(g) \rightarrow MgCO_3(s)$$

Combustion reactions
- Common example, hydrocarbon fuel reacts with oxygen to produce carbon dioxide and water
- Consider:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

Single replacement reactions
a. A common reaction: active metal replaces (displaces) a metal ion from a solution.

$$Mg(s) + CuCl_2(aq) \rightarrow Cu(s) + MgCl_2(aq)$$

- The activity series of metals is useful in order to predict the outcome of the reaction.

TABLE 4.6	Activity Series
Element	**Oxidation Half-Reaction**
Lithium	$Li \longrightarrow Li^+ + e^-$
Potassium	$K \longrightarrow K^+ + e^-$
Barium	$Ba \longrightarrow Ba^{2+} + 2e^-$
Calcium	$Ca \longrightarrow Ca^{2+} + 2e^-$
Sodium	$Na \longrightarrow Na^+ + e^-$
Magnesium	$Mg \longrightarrow Mg^{2+} + 2e^-$
Aluminum	$Al \longrightarrow Al^{3+} + 3e^-$
Manganese	$Mn \longrightarrow Mn^{2+} + 2e^-$
Zinc	$Zn \longrightarrow Zn^{2+} + 2e^-$
Chromium	$Cr \longrightarrow Cr^{3+} + 3e^-$
Iron	$Fe \longrightarrow Fe^{2+} + 2e^-$
Cadmium	$Cd \longrightarrow Cd^{2+} + 2e^-$
Cobalt	$Co \longrightarrow Co^{2+} + 2e^-$
Nickel	$Ni \longrightarrow Ni^{2+} + 2e^-$
Tin	$Sn \longrightarrow Sn^{2+} + 2e^-$
Lead	$Pb \longrightarrow Pb^{2+} + 2e^-$
Hydrogen	$H_2 \longrightarrow 2H^+ + 2e^-$
Copper	$Cu \longrightarrow Cu^{2+} + 2e^-$
Silver	$Ag \longrightarrow Ag^+ + e^-$
Mercury	$Hg \longrightarrow Hg^{2+} + 2e^-$
Platinum	$Pt \longrightarrow Pt^{2+} + 2e^-$
Gold	$Au \longrightarrow Au^{3+} + 3e^-$

Increasing ease of oxidation

8.8 Gas-Evolution Chemical Reactions

- In a **gas-evolution reaction**, a gas forms, resulting in bubbling.

TABLE 8.3 Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
Sulfides	None	H_2S	$2\,HCl(aq) + K_2S(aq) \longrightarrow H_2S(g) + 2\,KCl(aq)$
Carbonates and bicarbonates	H_2CO_3	CO_2	$2\,HCl(aq) + K_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g) + 2\,KCl(aq)$
Sulfites and bisulfites	H_2SO_3	SO_2	$2\,HCl(aq) + K_2SO_3(aq) \longrightarrow H_2O(l) + SO_2(g) + 2\,KCl(aq)$
Ammonium	NH_4OH	NH_3	$NH_4Cl(aq) + KOH(aq) \longrightarrow H_2O(l) + NH_3(g) + KCl(aq)$

Chemical Reactions
Chapter 8

Chapter 9-Thermochem

Internal Energy = (ΔE)
(state function)

$\Delta U = Q + W \leftarrow$ heat (J) / work

ΔE Change = $\Delta E = E_{prod/final} - E_{reactants/initial}$

A summary of the important sign conventions

$$\Delta E = q + W = q - P\Delta V$$

ΔE	−	internal energy of a system **decreases**
	+	internal energy of a system **increases**
q	−	the system **releases** heat into surroundings (lowers E)
	+	the system **absorbs** heat from the surroundings (raises E)
W	−	the system **does work** on its surroundings (lowers E)
	+	the **surroundings does work** on the system (raises E)
ΔV	−	the system is **compressed** by the surroundings (W is (+))
	+	the system **expands** against the surroundings (W is (−))

Quantifying Heat -
- when q+, the temp goes up,
- when q-, the temp goes down
- q & Δt are directly proportional

Heat Capacity -

$\frac{J}{°C} \nwarrow C = \frac{q}{\Delta T} \rightarrow °C$

Specific Heat Capicity -

$$q = M \cdot C_s \cdot \Delta t$$

heat(J) mass(g) specific heat($\frac{J}{g°c}$) temp change (°C)

TABLE 9.2 Specific Heat Capacities of Some Common Substances

Lead	0.128
Gold	0.128
Silver	0.235
Copper	0.385
Iron	0.449
Aluminum	0.903
Compounds	
Ethanol	2.42
Water	4.18
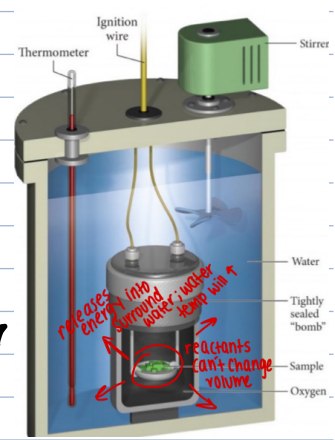
Work -
$W = -P_{external} \Delta V$ $\frac{P_{external}(atm)}{\Delta V = volume (L)} * \frac{1 L \cdot atm}{101.3 J} *$

Constant · Volume Calorimetry = $\Delta E_{rxn} = q_v$ (heat of volume)
Bomb Calorimetry = $\Delta E_{rxn} = -q_{cal}$ $q_{cal} = \Delta T \cdot C_{cal}$ heat capacity of entire assembly

Thermometer, Ignition wire, Stirrer, Water, Tightly sealed "bomb", Sample, Oxygen
releases energy into surround water- water temp will ↑
reactants can't change volume

Enthalpy (H) · $\Delta H = \Delta E + P\Delta V$ →
ΔH = enthalpy change of reaction
ΔH = heat under constant pressure
$-\Delta H \rightarrow$ heat is released, exothermic
$+\Delta H \rightarrow$ heat is absorbed, endothermic constant pressure

Constant Pressure Calorimetry-
measures $\Delta H_{rxn} = q_{rxn} = -q_{solution}$
(system)   (surroundings)

experimental conditions under which the reaction is taking place	property of reaction being directly measured by calorimeter.	apparatus used to perform heat measurements	property of reaction being indirectly established
Constant volume	q_v (heat of the reaction under constant volume)	bomb calorimeter	ΔE_{rxn}
Constant pressure	q_p (heat of the reaction under constant pressure)	coffee-cup calorimeter	ΔH_{rxn}

Hesses Law - computational way to determine ΔH
1) use helper reactions
2) Bond energies

$\Delta H_{rxn} = \Sigma (\Delta h \text{ bonds broken})^{(+)} + \Sigma (\Delta h \text{ bons formed})^{(-)}$

3) Standard enthalpies
$\Delta H_{rxn} = \Sigma n_p \Delta H_f - \Sigma n_r \Delta H_f$
(products)   (reactants)

Example: Estimate the enthalpy change for

$H_{2(g)} + F_{2(g)} \rightarrow 2HF_{(g)}$ $\Delta H_{rxn} = ?$

H-H F-F   H-F
          H-F

Bonds broken: Bonds made:
1 mol H-H: +436 kJ   2 mols H-F: 2(-565 kJ)
1 mol F-F: +159 kJ   _____
+595 kJ   -1130 kJ

$\Delta H_{rxn} = 595 kJ + (-1130 kJ) = -535 kJ$

Lattice Energy - Look at charges, if it is greater, it has a greater mag. If they are the same, look at radai. Larger the radai, smaller the lattice magnitude.

Example: rank the following ionic compounds in order of increasing magnitude of lattice energy:

  6   1   1   2   2
 +3 -2 +1 -1 +1 -1 +1 -2 +2 -1
Al_2O_3, KBr, LiF, Li_2O, $SrCl_2$

$KBr < LiF < SrCl_2 < Li_2O < Al_2O_3$

Bon Haber Cycle - $\Delta H_f° NaCl_{(s)} = \Delta H°_{sub} + IE + \frac{1}{2}BE(Cl_2) + EA_1 + \Delta H°_{lattice}$

- calculate the lattice energy for $NaCl_{(s)}$ reaction of intrest
$Na^+_{(g)} + Cl^-_{(g)} \rightarrow NaCl_{(s)}$ $\Delta H \frac{kJ}{mol}$

a) $\Delta H_f NaCl_{(s)} = -411.2 \frac{kJ}{mol}$: $Na_{(s)} + \frac{1}{2}Cl_{2(g)} \rightarrow NaCl_{(s)}$ leave a) $Na_{(s)} + \frac{1}{2}Cl_{2(g)} \rightarrow \boxed{NaCl_{(s)}}$ -411.2

b) $\Delta H_{sub} Na_{(s)} = 107.5 \frac{kJ}{mol}$: $Na_{(s)} \rightarrow Na_{(g)}$ reverse c) $Na^+_{(g)} + 1e^- \rightarrow Na_{(g)}$ -496

c) $IE_1 Na = 496 \frac{kJ}{mol}$: $Na_{(g)} \rightarrow Na^+_{(g)} + 1e^-$ reverse d) $Cl^-_{(g)} \rightarrow Cl_{(g)} + 1e^-$ 349

d) $EA_1 Cl = -349 \frac{kJ}{mol}$: $Cl_{(g)} + 1e^- \rightarrow Cl^-_{(g)}$ reverse e) x $\frac{1}{2}$ $Cl_{(g)} \rightarrow \frac{1}{2}Cl_{2(g)}$ $\frac{-243}{2} = -121.5$

e) Bond Energy $Cl_2 = 243 \frac{kJ}{mol}$: $Cl_{2(g)} \rightarrow 2Cl_{(g)}$ reverse b) $Na_{(g)} \rightarrow Na_{(s)}$ -107.5
                                                                          _____
$Na^+_{(g)} + Cl^-_{(g)} \rightarrow NaCl_{(s)}$ -787.2 kJ

Gas Pressure

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{F}{A}$$

$Pa = \text{height} \cdot \text{density} \cdot 9.8\, m/s^2$ $h = m$ $d = kg/m^3$

1 atm* = 101,325 Pa

= 760 mmHg*

= 760 torr*

= 1.01325 bar

= 14.7 psi

Boyle's Law

P × V = constant

$$P_1V_1 = P_2V_2$$

inversely proportional

Avogadro's Law

V = constant × n (moles)

V/n = constant $\dfrac{V_1}{n_1} = \dfrac{V_2}{n_2}$

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{\mathcal{M}_B}{\mathcal{M}_A}}$$ **Grahams Law**

$$\frac{time_A}{time_B} = \frac{\sqrt{M_A}}{\sqrt{M_B}}$$ **of Effusion**

Charles's Law

V = constant × T

V/T = constant

$$(V_1/T_1) = (V_2/T_2)$$

direct relationship

combined gas law

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

The Ideal Gas Equation

$$PV = nRT$$

$P =$ atm
$V =$ Liters
$n =$ moles
$T =$ temp k

$R = .08206$

Numerical value	Unit
0.08206	L · atm/K · mol
62.36	L · torr/K · mol
0.08314	L · bar/K · mol
8.314	m³ · Pa/K · mol
8.314	J/K · mol
1.987	cal/K · mol

Combined Gas

$P_1 = \dfrac{P_2T_1V_2}{T_2V_1}$ $P_2 = \dfrac{P_1T_2V_1}{T_1V_2}$

$T_1 = \dfrac{P_1T_2V_1}{P_2V_2}$ $T_2 = \dfrac{P_2T_1V_2}{P_1V_1}$

$V_1 = \dfrac{P_2T_1V_2}{T_2P_1}$ $V_2 = \dfrac{P_1T_2V_1}{P_2T_1}$

$n = \dfrac{PV}{RT}$ $M = \dfrac{nRT}{PV}$

Gas Density

$$d = \frac{P\mathcal{M}}{RT}$$

$M =$ molar mass
$P =$ pressure atm
$R = .08206$
$T =$ Temp kelvin

Vapor Pressure

TABLE 10.3 Vapor Pressure of Water versus Temperature

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0	4.58	55	118.2
5	6.54	60	149.6
10	9.21	65	187.5
15	12.79	70	233.7
20	17.55	75	289.1
25	23.78	80	355.1
30	31.86	85	433.6
35	42.23	90	525.8
40	55.40	95	633.9
45	71.97	100	760.0
50	92.6		

Partial Pressures

$P_{Total} = P_A + P_B + P_C$

$P_{Total} = n_A\frac{RT}{V} + n_B\frac{RT}{V} + n_C\frac{RT}{V}$

$P_{Total} = n_{total}\frac{RT}{V}$

n_A n_b n_C

$\chi_A = \dfrac{n_a}{n_{total}}$ $\chi_B = \dfrac{n_b}{n_{total}}$ $\chi_C = \dfrac{n_c}{n_{total}}$

$P_A = \chi_A \cdot P_{Total}$

Van der Waals

$[P + a(\frac{n}{V})^2] \cdot$

const ant

$[V - nb]$

Constant $=$

nRT

STP

Calc the V of a mole of ideal gas at room tem (25°C) and 1 atm

$T = 273$ $P = 1\,atm$

Root Mean V

$$\mathbf{v}_{rms} = \sqrt{\frac{3RT}{N_A m}}$$

$$v_{rms} = \sqrt{\frac{3RT \leftarrow k \,(8.314)}{M \to kg}}$$
molecular mass

Temperature and Molecular Velocities

$KE = \frac{1}{2}mv^2$ $m = kg$

$KE_{avg} = \frac{1}{2}$ avagadros $m\overline{v^2}$

$KE = \frac{3}{2}RT$

Real Gases

Ideal gas laws assume
– no attractions between gas molecules.
– gas molecules do not take up space.
 • Based on the kinetic molecular theory

But at lower temperatures, the pressure of gases is *less than* that of an ideal gas.
– At the lower temperatures, the gas atoms spend more time interacting with each other and less time colliding with the walls, making the actual pressure less than that predicted by the ideal gas law.

Gases behave ideally at low pressures and high temperatures

At low temperatures and high pressures, however, these assumptions are not valid.

Reason:
– Gas molecules occupy significant volume (at high pressures)
– Gas molecules experience intermolecular forces of attraction and repulsion (at low temperatures)

Intermolecular Forces

gas ⟶ weak
Liquid ⟶ moderate
Solid ⟶ strong

The higher the normal boiling point of the liquid, the stronger the intermolecular attractive forces.
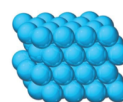
The strength of the attractions between the particles of a substance determines its state (e.g., gas, liquid, solid).
- At room temperature, moderate to strong attractive forces result in materials being solids or liquids.
  • The stronger the attractive forces are, the higher the boiling point of the liquid and the melting point of the solid will be.

Solids

Regular ordered structure	No long-range order
Crystalline solid	Amorphous solid
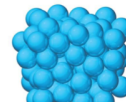
- **Dispersion forces** have the following characteristics: – Excess e⁻ density has (–) charge and depleted has partial (+) charge
  - They are the weakest of the intermolecular attractions.
  - They are present in all molecules and atoms.
  - Their magnitude increases with molar mass.
  -

- **Polar molecules**
  - have dipole–dipole attractive forces.
  - The permanent dipole adds to the attractive forces raising Boiling Point and melting points (similar shape/size)

- **Hydrogen bonds** are
  - a type of dipole–dipole interaction.
  - the strongest of the intermolecular attractive forces a pure substance can have.
  - present when a molecule has H directly bonded to either O, N, or F atoms.
    • The only example of H bonded to F is HF.
  - Hydrogen bonds mean higher boiling points

- **Ion–dipole attractions** are
  - present in mixtures of ionic compounds with polar molecules.
  - the strongest of the intermolecular attractions.
  - especially important in aqueous solutions of ionic compounds. (effects solubility)
  - Like dissolves like

– In a mixture, ions from an ionic compound are attracted to the dipole of polar molecules.

• Immiscible Liquids – only similar compounds mix.
• If non-polar look at atomic mass

## Surface Tension, Viscosity and Capillary Action

ST
- Surface tension is a property of liquids that results from the tendency of liquids to minimize their surface area.
- The surface tension of a liquid is the energy required to increase the surface area a given amount.
- Stronger the intermolecular attractive forces, higher the surface tension

Vis
• **Viscosity** is the resistance of a liquid to flow.
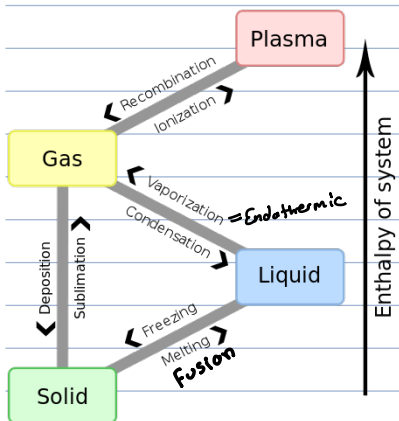 • Larger intermolecular attractions = larger viscosity

Cap
- **Capillary action** is the ability of a liquid to flow up a thin tube against the influence of gravity.
- uses adhesive forces (pull up the tube) and cohesive forces (holds molecules to forces)
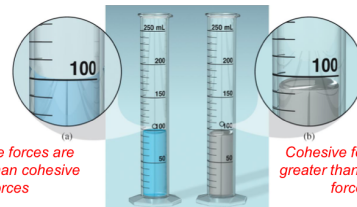
Adhesive forces are greater than cohesive forces *Cohesive forces are greater than adhesive forces*

Phase Changes

Plasma — Recombination / Ionization — Gas — Vaporization = Endothermic / Condensation — Liquid — Freezing / Melting — Fusion — Solid — Deposition / Sublimation

Enthalpy of system

Effect of Intermolecular Attraction on Vaporization and Condensation

- The weaker the attractive forces between molecules, the less energy the molecules will need to vaporize.
- Also, weaker attractive forces means that more energy will need to be removed from the vapor molecules before they can condense.
- The net result will be more molecules in the vapor phase and a liquid that evaporates faster; the weaker the attractive forces, the faster the rate of evaporation.
- Liquids that evaporate easily are said to be **volatile**.
  – Examples: gasoline and fingernail polish remover
- Liquids that do not evaporate easily are called **nonvolatile**.
  – Example: motor oil

heat of vaporization, ΔH_{vap} (always +)

Vapor Pressure

- The pressure exerted by the vapor when it is in dynamic equilibrium with its liquid is called the **vapor pressure**.
- The weaker the attractive forces between the molecules, the more molecules will be in the vapor. Therefore,
  – the weaker the attractive forces, the higher the vapor pressure.
  – the higher the vapor pressure, the more volatile the liquid.

• As temp increases, the vapor pressure increases

Boiling Point
• when vapor pressure is equal to atm.
• its an exponential relation
• Depends on altitude

Clausius–Clapeyron
• vapor pressure and temp

$$\ln\frac{P_2}{P_1} = \frac{-\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

R = 8.314

• Can predict vapor pressure if you know ΔH and boiling point. VP at norm BP = 1atm
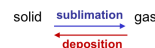
Sublimation and Fusion

Molecules in the solid state have thermal energy that allows them to vibrate.

Surface molecules with sufficient energy may break free from the surface and become a gas. This process is called **sublimation**.
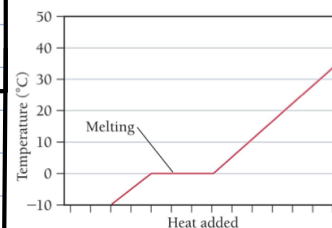
The capturing of vapor molecules into a solid is called **deposition**.

The solid and vapor phases exist in dynamic equilibrium in a closed container at temperatures below the melting point.
  – Therefore, **molecular solids have a vapor pressure.**

solid ⇄ (sublimation / deposition) gas

Heating Curve of a Solid

Heating Curve for Water