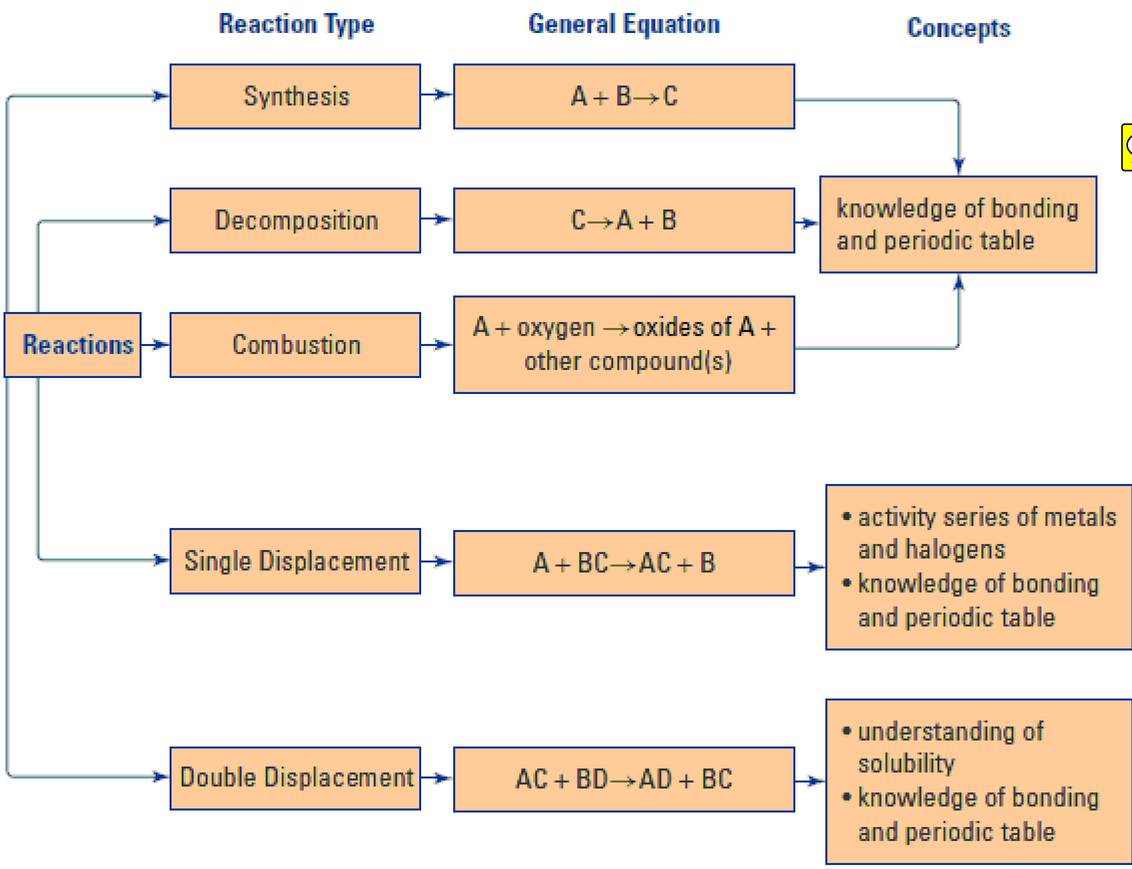
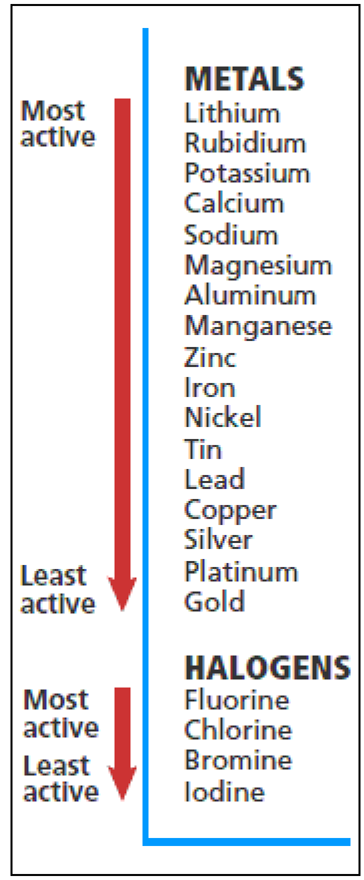
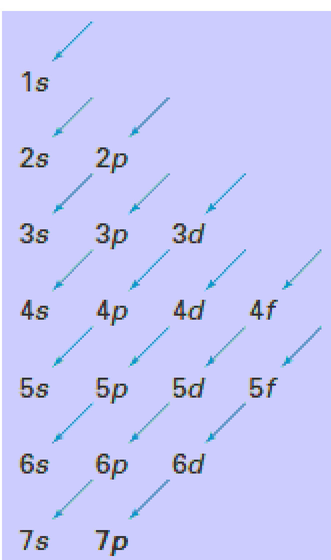
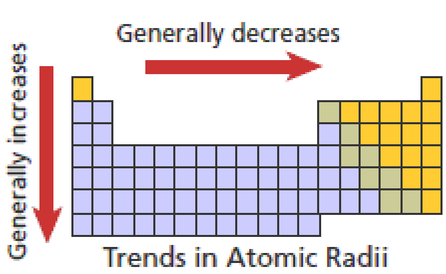
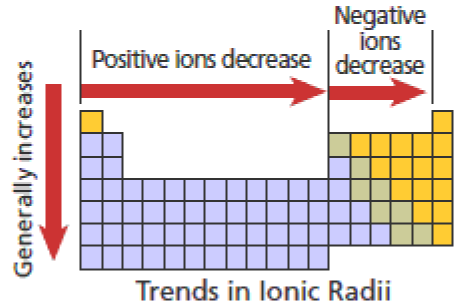


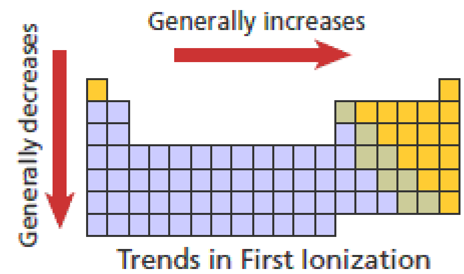


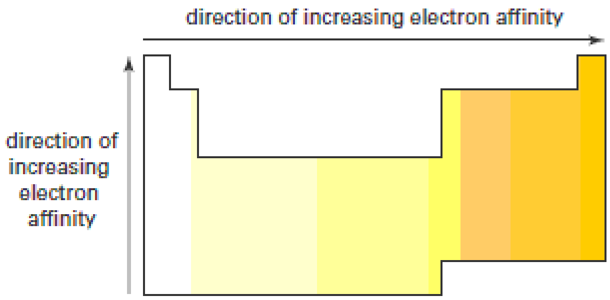


Sublevels (types of orbitals) present	Number of orbitals related to sublevel
s	1
p	3
d	5
f	7

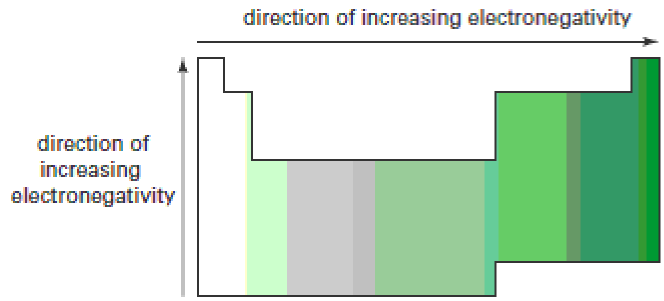








C Trends in electron affinity



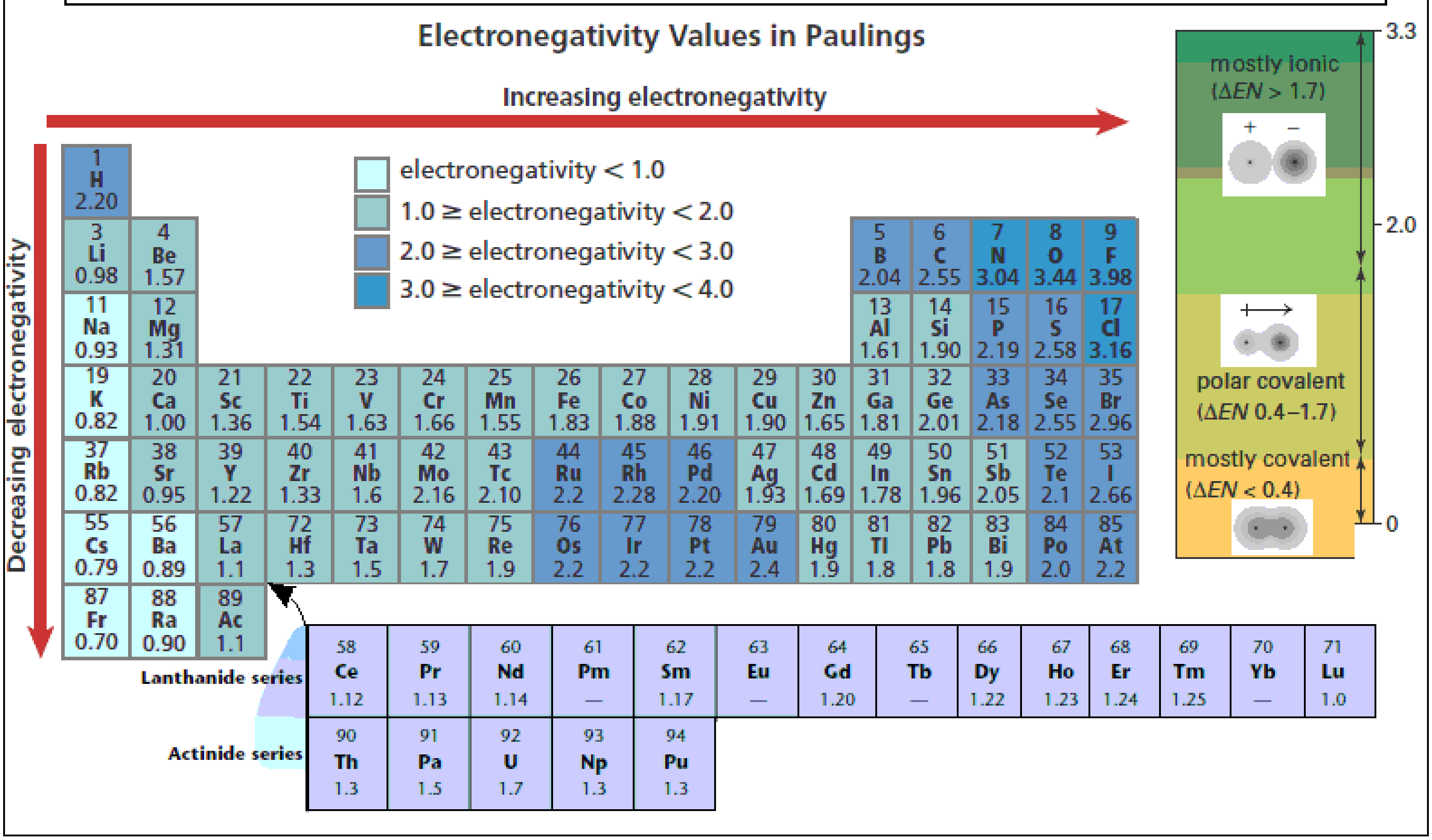
D Trends in electronegativity

Keywords: Atomic Radius & IE : ENC, Electron Affinity & EN : Coulomb's law

Strength of Bonds : 1) Ionic Bond Strength given by Lattice Energy (Colomb's Law) F = [(q1 * q2) / r^2]

2) Covalent bond (bond energy)

3) Metallic Bond (Coulomb's Law)



Guideline	Cations	Anions	Result	Exceptions
1	Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+	NO_3^-, CH_3COO^-, ClO_3^-	soluble	$Ca(ClO_3)_2$ is insoluble

Valence = –1			
Ion	**Name**	**Ion**	**Name**
CN^-	cyanide	$H_2PO_3^-$	dihydrogen phosphite
CH_3COO^-	acetate	$H_2PO_4^-$	dihydrogen phosphate
ClO^-	hypochlorite	MnO_4^-	permanganate
ClO_2^-	chlorite	NO_2^-	nitrite
ClO_3^-	chlorate	NO_3^-	nitrate
ClO_4^-	perchlorate	OCN^-	cyanate
HCO_3^-	hydrogen carbonate	HS^-	hydrogen sulfide
HSO_3^-	hydrogen sulfite	OH^-	hydroxide
HSO_4^-	hydrogen sulfate	SCN^-	thiocyanate

Valence = –2			
Ion	**Name**	**Ion**	**Name**
CO_3^{2-}	carbonate	O_2^{2-}	peroxide
$C_2O_4^{2-}$	oxalate	SiO_3^{2-}	silicate
CrO_4^{2-}	chromate	SO_3^{2-}	sulfite
$Cr_2O_7^{2-}$	dichromate	SO_4^{2-}	sulfate
HPO_3^{2-}	hydrogen phosphite	$S_2O_3^{2-}$	thiosulfate
HPO_4^{2-}	hydrogen phosphate		

Prefix and suffix			Number of oxygen atoms
hypo		ite	$x - 2$ oxygen atoms
		ite	$x - 1$ oxygen atoms
		ate	x oxygen atoms
per		ate	$x + 1$ oxygen atoms

Unit of concentration	Mathematical expression
molar concentration (mol/L), C Molarity	$\frac{\text{mol of solute, } n}{\text{volume of solution, } V}$, also written as $C = \frac{n}{V}$

Gas Constant, R

Units of R	Numerical value of R
$\frac{L \cdot atm}{mol \cdot K}$	0.0821
$\frac{L \cdot kPa}{mol \cdot K}$	8.314
$\frac{L \cdot mm\ Hg}{mol \cdot K}$	62.4

Relationships Among Energy Units

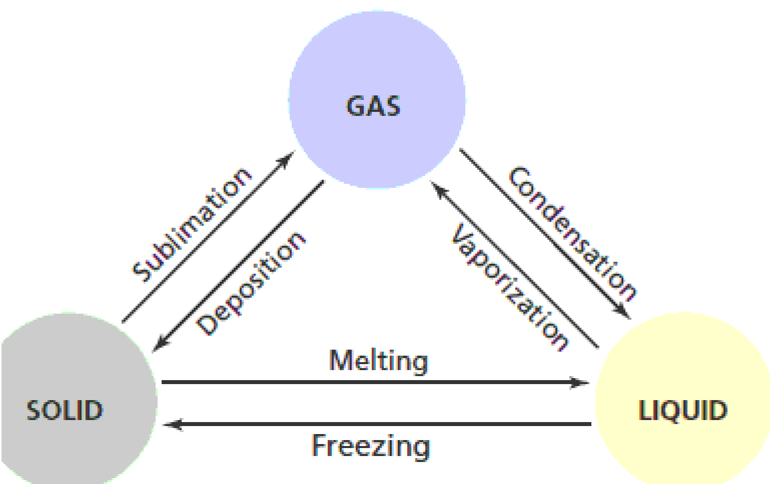
Relationship	Conversion factors
1 J = 0.2390 cal	$\frac{1\ J}{0.2390\ cal}$ $\frac{0.2390\ cal}{1\ J}$
1 cal = 4.184 J	$\frac{1\ cal}{4.184\ J}$ $\frac{4.184\ J}{1\ cal}$
1 kJ = 1000 J	$\frac{1\ kJ}{1000\ J}$ $\frac{1000\ J}{1\ kJ}$
1 Calorie = 1 kcal 1 kcal = 1000 cal	$\frac{1\ Calorie}{1000\ cal}$ $\frac{1000\ cal}{1\ kcal}$

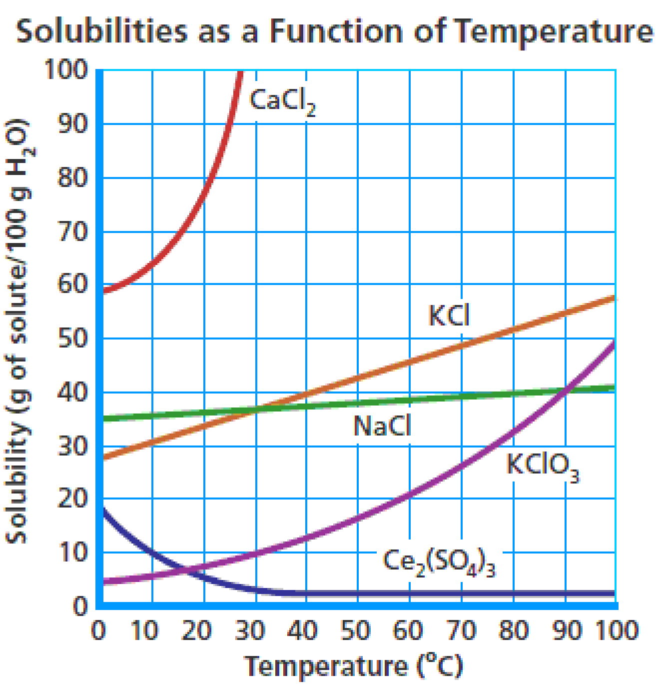
Enthalpy Changes for Exothermic and Endothermic Reactions

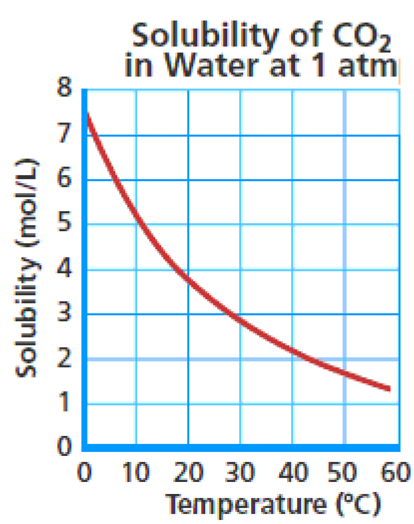
Type of reaction	Sign of ΔH_{rxn}
Exothermic	Negative
Endothermic	Positive

strong forces ← weak forces

Force	Ionic	Polar (dipole–dipole)	Dispersion
Type of force	between ions (intramolecular)	between molecules (intermolecular)	between molecules (intermolecular)
State	usually solid	liquid or gas (can also be solid)	liquid or gas
Example	$NaCl_{(s)}$	$CH_3CH_2OH_{(\ell)}$, $HCl_{(g)}$	$C_5H_{12(\ell)}$, $CH_{4(g)}$, $CO_{2(g)}$

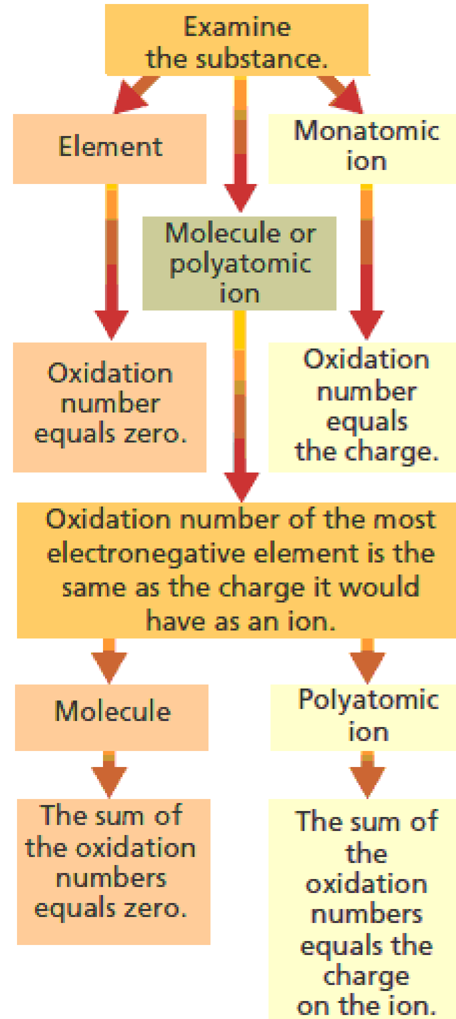






Rules for determining oxidation numbers

Rules	Examples
1. A pure element has an oxidation number of 0.	Na in $Na_{(s)}$, Br in $Br_{2(\ell)}$, and P in $P_{4(s)}$ all have an oxidation number of 0.
2. The oxidation number of an element in a monatomic ion equals the charge of the ion.	The oxidation number of Al in Al^{3+} is +3. The oxidation number of Se in Se^{2-} is –2.
3. The oxidation number of hydrogen in its compounds is +1, except in metal hydrides, where the oxidation number of hydrogen is –1.	The oxidation number of H in H_2S or CH_4 is +1. The oxidation number of H in NaH or in CaH_2 is –1.
4. The oxidation number of oxygen in its compounds is usually –2, but there are exceptions. These include peroxides, such as H_2O_2, and the compound OF_2.	The oxidation number of O in Li_2O or in KNO_3 is –2.
5. In covalent compounds that do not contain hydrogen or oxygen, the more electronegative element is assigned an oxidation number that equals the negative charge it usually has in its ionic compounds.	The oxidation number of Cl in PCl_3 is –1. The oxidation number of S in CS_2 is –2.
6. The sum of the oxidation numbers of all the elements in a compound is 0.	In CF_4, the oxidation number of F is –1, and the oxidation number of C is +4. $(+4) + 4(-1) = 0$
7. The sum of the oxidation numbers of all the elements in a polyatomic ion equals the charge on the ion.	In NO_2^-, the oxidation number of O is –2, and the oxidation number of N is +3. $(+3) + 2(-2) = -1$



Name	carbons	molecular formula
methane	1	CH_4
ethane	2	CH_3CH_3
propane	3	$CH_3CH_2CH_3$
butane	4	$CH_3(CH_2)_2CH_3$
pentane	5	$CH_3(CH_2)_3CH_3$
hexane	6	$CH_3(CH_2)_4CH_3$
heptane	7	$CH_3(CH_2)_5CH_3$
octane	8	$CH_3(CH_2)_6CH_3$
nonane	9	$CH_3(CH_2)_7CH_3$
decane	10	$CH_3(CH_2)_8CH_3$

Organic Compounds and Their Functional Groups

Compound type	General formula	Functional group
alkane	none	propane
alkene	$>C=C<$	propene
alkyne	$-C \equiv C-$	propyne
Halocarbon	R—X (X = F, Cl, Br, I)	Halogen

Key Equations and Relationships

- density: $\text{density} = \frac{\text{mass}}{\text{volume}}$
- conversion between temperature scales: $°C + 273 = K$; $K - 273 = °C$
- law of conservation of mass
  $\text{Mass}_{\text{reactants}} = \text{Mass}_{\text{products}}$
- Determining the number of protons and electrons
  Atomic number = number of protons = number of electrons
- EM Wave relationship: $c = \lambda\nu$
- Energy of a quantum: $E_{\text{quantum}} = h\nu$
- Energy of a photon: $E_{\text{photon}} = h\nu$
- percent error:
  $\text{percent error} = \frac{\text{error}}{\text{accepted value}} \times 100$
- slope of graph: $\text{slope} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{\Delta y}{\Delta x}$
- Determining the number of neutrons
  Number of neutrons = mass number − atomic number
- Energy change of an electron:
  $\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}}$
  $\Delta E = E_{\text{photon}} = h\nu$
- de Broglie's equation: $\lambda = \frac{h}{m\nu}$
- $\text{number of representative particles} = \text{number of moles} \times \frac{6.02 \times 10^{23}\text{ representative particles}}{1\text{ mole}}$
- $\text{number of moles} = \text{number of representative particles} \times \frac{1\text{ mole}}{6.02 \times 10^{23}\text{ representative particles}}$
- $\text{mass} = \text{number of moles} \times \frac{\text{number of grams}}{1\text{ mole}}$
- $\text{number of moles} = \text{mass} \times \frac{1\text{ mole}}{\text{number of grams}}$
- $\text{percent by mass} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$
- $\text{molecular formula} = (\text{empirical formula})n$
- $\text{moles of known} \times \frac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$
- $\frac{\text{actual yield (from experiment)}}{\text{theoretical yield (from stoichiometric calculations)}} \times 100 = \text{percent yield}$
- Kinetic energy: $KE = 1/2mv^2$
- Graham's effusion: $\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$
- Dalton's law of partial pressures: $P_{\text{total}} = P_1 + P_2 + P_3 + \ldots P_n$

- Boyle's law: $P_1V_1 = P_2V_2$, constant temperature
- Charles's law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, constant pressure
- Gay-Lussac's law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, constant volume
- Henry's law: $\frac{S_1}{P_1} = \frac{S_2}{P_2}$

- Combined gas law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
- Ideal gas law: $PV = nRT$
- Finding molar mass: $M = \frac{mRT}{PV}$
- Finding density: $D = \frac{MP}{RT}$
- Molarity-volume: $M_1V_1 = M_2V_2$

- Molarity $(M) = \frac{\text{moles of solute}}{\text{liters of solution}}$

- $q = c \times m \times \Delta T$

- $\Delta H^{\circ}_{rxn} = \Sigma\Delta H^{\circ}_f(\text{products}) - \Sigma\Delta H^{\circ}_f(\text{reactants})$