Unit 2 - Atomic Concepts

Element

Mass Number 14

(protons + neutrons)

N

Atomic Number 7

(# of protons)

Protons = Electrons (since atoms are neutral)

Atom Theories

Dalton	Elements are made of indivisible particles called atoms
Thomson	Discovered the electron using "cathode rays" that are deflected by electric and magnetic fields ("plum pudding" model)
Rutherford	Gold Foil Experiment – atom is mostly empty space with a small, dense, positive nucleus
Bohr	Electrons must exist in fixed, stable orbits (energy levels), which are at a specific distance from the nucleus
Quantum	Electrons are in orbitals (regions of probable location)

Bohr Diagram

- **nucleus & electron configuration**

Ex: Potassium-40 (Potassium=K)

$p^+ = 19$

$n^0 = 21$

2   8   8   1

Electron configuration = 2-8-8-1
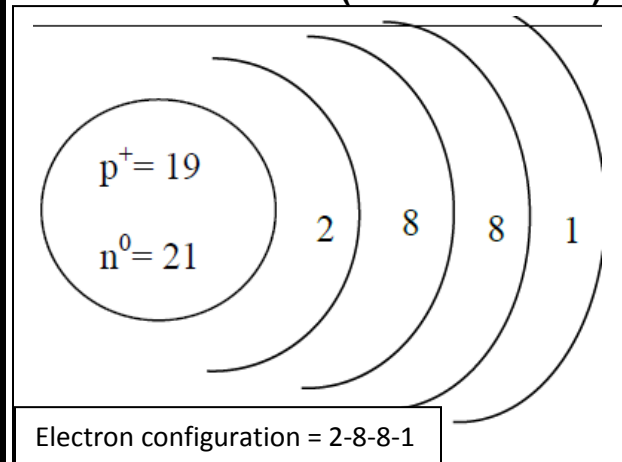
Isotopes *("iso-"= same)*

- **Same # of Protons**
- **Different # of Neutrons**

16

O

8

17

O

8

Atoms

Particle	Charge	Location in the Atom	Mass
Proton	+1	Inside nucleus	1u
Neutron	0	Inside nucleus	1u
Electron	−1	Outside nucleus	0u

Lewis Dot Diagram

- **dots for # of valence electrons (last # in electron configuration)**

K = Potassium

K
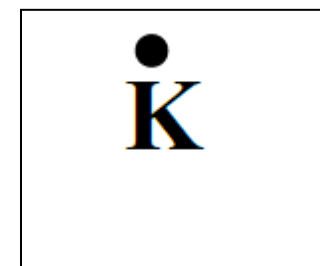
= metalloids

= non-metals

= metals

= alkali metals

= noble gases

Periodic Table of the Elements

= halogens

= alkaline earth metals

Key

6	← Atomic number
C	← Element's symbol
Carbon	← Element's name
12.011	← Atomic mass

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18

1
1 H Hydrogen

2 He Helium 4.003

2
3 Li Lithium 6.941
4 Be Beryllium 9.0122
5 B Boron 10.81
6 C Carbon 12.011
7 N Nitrogen 14.007
8 O Oxygen 15.999
9 F Fluorine 18.998
10 Ne Neon 20.180

3
11 Na Sodium 22.990
12 Mg Magnesium 24.305
13 Al Aluminum 26.98
14 Si Silicon 28.086
15 P Phosphorus 30.974
16 S Sulfur 32.06
17 Cl Chlorine 35.453
18 Ar Argon 39.948

4
19 K Potassium 39.098
20 Ca Calcium 63.078
21 Sc Scandium 44.956
22 Ti Titanium 47.88
23 V Vanadium 50.94
24 Cr Chromium 51.996
25 Mn Manganese 54.938
26 Fe Iron 55.847
27 Co Cobalt 58.9332
28 Ni Nickel 58.693
29 Cu Copper 63.546
30 Zn Zinc 65.39
31 Ga Gallium 69.72
32 Ge Germanium 72.61
33 As Arsenic 74.922
34 Se Selenium 78.96
35 Br Bromine 79.904
36 Kr Krypton 83.80

5
37 Rb Rubidium 85.468
38 Sr Strontium 87.62
39 Y Yttrium 88.9059
40 Zr Zirconium 91.224
41 Nb Niobium 92.91
42 Mo Molybdenum 95.94
43 Tc Technetium 97.907
44 Ru Ruthenium 101.07
45 Rh Rhodium 102.906
46 Pd Palladium 106.42
47 Ag Silver 107.868
48 Cd Cadmium 112.41
49 In Indium 114.82
50 Sn Tin 118.710
51 Sb Antimony 121.757
52 Te Tellurium 127.60
53 I Iodine 126.904
54 Xe Xenon 131.29

6
55 Cs Cesium 132.905
56 Ba Barium 137.327
57 La Lanthanum 138.906
72 Hf Hafnium 178.49
73 Ta Tantalum 180.95
74 W Tungsten 183.85
75 Re Rhenium 186.207
76 Os Osmium 190.2
77 Ir Iridium 192.22
78 Pt Platinum 195.08
79 Au Gold 196.967
80 Hg Mercury 200.59
81 Tl Thallium 204.383
82 Pb Lead 207.2
83 Bi Bismuth 208.98
84 Po Polonium 208.982
85 At Astatine 209.987
86 Rn Radon 222.018

7
87 Fr Francium 223.020
88 Ra Radium 226.005
89 Ac Actinium 227.028
104 Rf Rutherfordium (261)
105 Db Dubnium (262)
106 Sg Seaborgium (263)
107 Bh Bohrium (262)
108 Hs Hassium (265)
109 Mt Meitnerium (266)
110 Uun Ununnilium (269)
111 Uuu Unununium (272)
112 Uub Ununbium (277)
114 Uuq Ununquadium (285)
116 Uuh Ununhexium (289)
118 Uuo Ununoctium (293)

Rare-Earth Elements

Lanthanide Series
58 Ce Cerium 104.115
59 Pr Praseodymium 140.908
60 Nd Neodymium 144.24
61 Pm Promethium 144.913
62 Sm Samarium 150.36
63 Eu Europium 151.965
64 Gd Gadolinium 157.25
65 Tb Terbium 158.925
66 Dy Dysprosium 162.50
67 Ho Holmium 164.93
68 Er Erbium 167.26
69 Tm Thulium 168.934
70 Yb Ytterbium 173.04
71 Lu Lutetium 174.967

Actinide Series
90 Th Thorium 232.038
91 Pa Protactinium 231.036
92 U Uranium 238.029
93 Np Neptunium 237.048
94 Pu Plutonium 244.064
95 Am Americium 243.061
96 Cm Curium 247.070
97 Bk Berkelium 247.070
98 Cf Californium 251.080
99 Es Einsteinium 252.083
100 Fm Fermium 257.095
101 Md Mendelevium 258.099
102 No Nobelium 259.101
103 Lr Lawrencium 260.105

UNIT 3: COMPOUND FORMATION SUMMARY

Atoms to Ions

To form an ion, an atom needs to gain or lose electrons. This changes the atom's size!

Cation: lose electrons to get + charge

$$Mg \xrightarrow{lose\ 2e^-} [Mg]^{2+}$$

Anion: gain electrons to get - charge

$$F \xrightarrow{gain\ 1e^-} [F]^-$$

Ionic bonds

Ionic bonds hold cations and anions together

$$[Na]^+ [Cl]^- \qquad [F]^- [Mg]^{2+} [F]^-$$

NaCl $\qquad\qquad$ MgF$_2$

1 cation; 1 anion $\quad$ 1 cation; 2 anions

- Between a metal and a non-metal OR a cation/anion from Table E
- Electrons are transferred (from a metal to a nonmetal)

- Conductive *only* when dissolved in water or melted (broken up into ions)
- High melting points: always solid at STP

Covalent bonds

Covalent (molecular) bonds hold electron-greedy nonmetals together

- Between nonmetals
- Electrons are shared between nonmetals
- Hold together the nonmetals of the polyatomic ions on Table E

- Not conductive as a solid or liquid
- Relatively low melting points

Naming Ionic Compounds

_____ (___) _____
Name of metal/cation $\quad$ Cation charge $\quad$ Name of nonmetal/anion

LiBr $\quad$ = **lithium** *bromide*
$\qquad$ *charge of Li always known; no Roman numeral needed

Li$_2$(CO$_3$) = **lithium** *carbonate*
$\qquad$ *charge of Li always known
$\qquad$ * Anion name comes straight from Table E

Cu$_2$O$_3$ $\quad$ = **copper (III)** *oxide*
$\qquad$ *charge of Cu NOT always known; needs Roman numeral

Cu$_2$O$_3$
+③ -2 $\quad$ Charge of cation (Cu) here
+③ -2 $\quad$ must be +3.

Balancing Equations

The smallest, whole-number coefficients need to be used to make sure there are the same number of atoms of each element on both sides of the equation.

$$Fe(OH)_2 + 2Ag(NO_3) \rightarrow Fe(NO_3)_2 + 2Ag(OH)$$

Fe: 1	Fe: 1
Ag: ✗ 2	Ag: ✗ 2
(OH): 2	(OH): ✗ 2
(NO$_3$): ✗ 2	(NO$_3$): 2

Mole Calculations

Use gram-formula (molar) mass to convert between grams and moles.

$$\frac{\#\ of\ moles}{1} = \frac{given\ mass}{gram - formula\ mass}$$

Mole = set number of particles

Given mass = mass of a compound you are given or want to produce

Gram-formula (molar) mass = add up mass of elements to find mass of ONE MOLE

$$Ca_3(PO_4)_2 = 310.\ g/mol$$
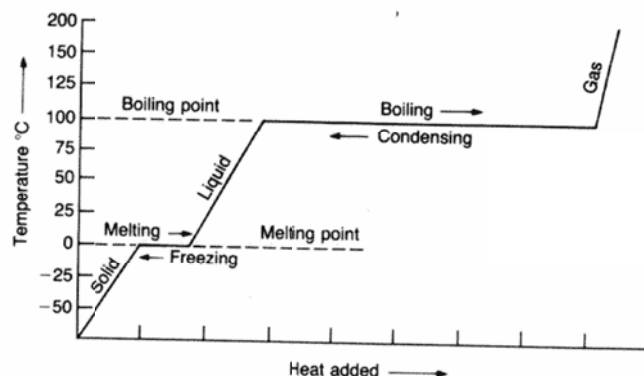$$3(40.0) + 2(31.0) + 8(16.0)$$

UNIT 4: PARTICLE MOTION SUMMARY

Heat, Temperature, Phases

*Applying or removing heat changes a substance's particle arrangement, which changes its **physical** properties.*

Temperature: measurement of *average kinetic energy*

Heat: form of thermal energy that moves from high temperatures to low temperatures
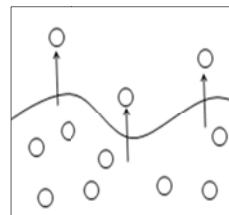

Vapor Pressure

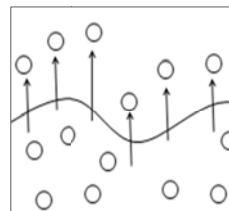*Vapor pressure results from collisions from "escaped" gas particles.*

Substances have low vapor pressure when…
- Temp is decreased
- Stronger IMFs
- Higher boiling point

Substances have high vapor pressure when…
- Temp is increased
- Weaker IMFs
- Lower boiling point

Use **Table H** to find **boiling points** of 4 liquids
A liquid boils when…
vapor pressure = atmospheric pressure

Gas Laws/Relationships

Changes in gas particle movement result in measureable change.

Kinetic Molecular Theory says that **ideal** gases…

Gases are most ideal (least "real") at high temperatures and low pressures (spread apart)
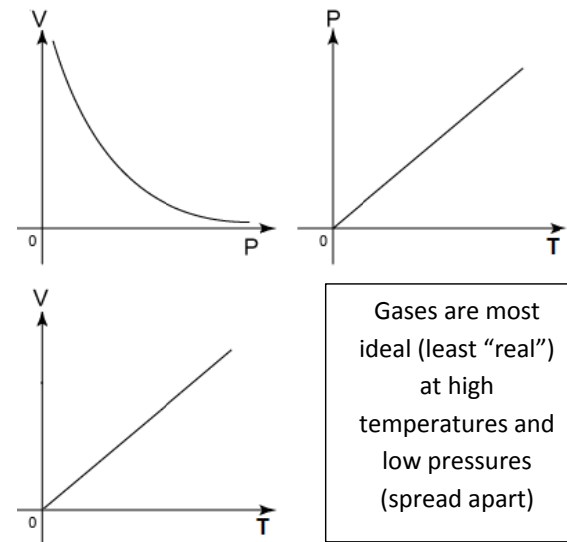
Kinetic Molecular Theory

- particles in all matter are in motion
- all particles have some attraction for each other *(IMFs or InterMolecular Forces)*
- add heat & particle motion (KE) increases
- more KE allows for overcoming of IMFs
- as you go from S → L → G; KE increases and IMFs decrease

- substances with higher MP & BP must have stronger IMFs (strong IMFs require lots of nrg to overcome)

Math: Conversions

Use Table T (temperature) or Table A (pressure) to convert units.

Temperature: $K = °C + 273$

- Temperature reflects how quickly particles are moving

Pressure: $1\ atm = 101.3\ kPa$

- Pressure reflects how often particles are colliding with container walls

Math: Combined Gas Law

Changes in P, V, and T are proportional.

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

Temperature MUST be in Kelvin!

If a variable is not mentioned or is said to remain constant, you do not have to include it in your calculation.

Temperature and pressure at **STP** are recorded on Table A.

Chemistry Unit 4 Fact/Concept/Skill List *(Particle Behavior of Matter - Phases)*

1. **Old phase change terms to know (plus new variations):** melting *(fusion)* freezing *(solidification)*, boiling *(vaporization)*, condensing, sublimating, deposition
2. **Solids:** particles fixed, vibrating, geometric structure, low nrg, strong IMFs, lower avg KE
3. **Liquids:** particles sliding around, medium nrg, medium IMFs, medium avg KE
4. **Gases:** fast, bouncing & rotating, high nrg, weak/no IMFs, highest avg KE
5. Gases *take shape of and fill container,* liquids just take shape but don't fill. Solids do neither.
6. **endothermic** = heat is added, heat goes in *(melting & boiling & sublimating!)*
7. **exothermic** = heat is removed, heat exits, heat goes out *(freezing & condensing & deposition!)*
8. **Temperature** = average kinetic energy of the particles (also know **heat flows** from hot to cold)
9. **Kelvin temp scale:** know conversion to Celcius (formula in Ref Table)
10. **Table S skill:** given a temperature of an element, determine if it is S or L or G
11. Substances with **higher MP or BP have stronger IMFs** *(particles hold tightly, so you must add more heat energy to get particles to move apart & change phase from S to L or L to G)*
12. Substances with lower MP or BP have weaker IMFs
13. **Pressure:** measured in atm or kPa (# of collisions in a gas), converting atm to kPa
14. **Kinetic Molecular Theory:** behaviors of the 3 phases are due to differences in the particle motion *(be able to compare the motion, spacing, energy, etc. of particles in all 3 phases)*
15. **PTV** (as P ↗ the V ↘ , as T ↗ the V ↗ , as P ↗ the V ↗)
16. Know the graphs for P T V
17. **atmospheric pressure** = air pressure (know STP from Table A)
18. **combined gas law** is over there → *(make a shopping list!)*
19. **ideal gases:** fast & straight motion, no IMFs, particles are tiny (compared to the space), can only transfer nrg (in collisions) can't lose or change nrg type
20. **real gases:** slower, have some particle attractions, take up more space (compared to vol of container), can lose nrg (or change KE to heat, etc.)
21. real gases will **behave like ideal** at: *HIGH* temperature & *LOW* pressure (reduces interactions)
22. **vapor pressure** = pressure above a liquid due to evap (boiling & evap are both *vaporization*)
23. **boiling** = when VP is equal to atmospheric pressure
24. **Table H skill:** read VP diagram in Table H, normal BP, high VP vs. low VP
25. **Heating Curves:** find MP, FP, BP, where sol, liq, gas exist, plus where s + l, l + g exist, plus find where KE is steady but PE increasing, where KE increasing but PE constant, exo/endo
26. **Heat of Fusion** & **Heat of Vaporization** (Table B) (FP same as MP) - *know that vaporizing takes more nrg than melting (fusion) b/c it's easier to weaken IMFs than to totally overcome them*
27. **strong IMFs** lead to… high MP, high BP, **low VP**
28. **weak IMFs** lead to… low MP, low BP, **high VP**
29. Given 2 containers of gas: If P T V are the same then # particles (moles) = same (or if # particles & V and T are same, P must be the same, etc) *If 3 are the same, the 4th must be.*

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

UNIT 5: REACTION DYNAMICS SUMMARY

Conservation of Mass

Atoms (and their mass) cannot be created or destroyed over the course of a change.

Balanced equations have an equal number of each element on their reactants and products sides.
- The total mass of the reactants will be equal to the total mass of the products.

Many reactions can be generalized by how the atoms move and rearrange themselves:
- **Synthesis**: A + B → C
- **Decomposition**: C → B + A
- **Single Replacement**: A + BC → B + AC
- **Double Replacement**: AB + CD → AD + CB

Conservation of Energy

To conserve overall energy, changes release heat (exothermic) or absorb heat (endothermic). Changes prefer to decrease potential energy and increase entropy.

Endothermic changes require heat to be put in:
- $reactants + heat \rightarrow products$
- ΔH will be positive for these reactions

Exothermic changes release heat as a product:
- $reactants \rightarrow products + heat$
- ΔH will be negative for these reactions

Use **Table I** to find a list of balanced equations with tabulated values for ΔH.

Entropy measures chemical disorder!

Conservation of Charge

Net charge remains the same throughout a reaction. In a redox reaction, this is true because the number of electrons lost is equal to the number of electrons gained.

Redox reactions are changes in which one species is *oxidized* and another is *reduced*.

OIL RIG | **O**xidation **i**s **l**oss (of e⁻)
Reduction **i**s **g**ain (of e⁻)

LEO says **GER** | **L**oss of **e**⁻ is **o**xidation
Gain of **e**⁻ is **r**eduction

- ALL single replacement reactions are redox reactions.
- Double replacement reactions are NEVER redox reactions.

Math: Mole Ratios

Coefficients in balanced equations represent the number of moles of an element/compound present.

Mole ratios are *ratios of coefficients* in a balanced chemical equation. They can be used to predict the effects of changes to a chemical system.

$$2H_2 + O_2 \rightarrow 2H_2O$$

How many moles of **oxygen** would be used if a car uses 11.4 moles of **hydrogen**?

$$\frac{2\ mol\ H_2}{1\ mol\ O_2} = \frac{11.4\ mol\ H_2}{x\ mol\ O_2}$$

Known mole ratio between compounds mentioned in problem	Change mentioned in problem and unknown quantity (x) to solve for

Math: PE Diagrams (Graphs)

Potential energy (PE) diagrams show changes in PE over the course of a reaction.

- **Catalysts** lower the **activation energy** of a reaction.

Math: Changes in Oxidation States

The "transfer" of electrons results in changes in oxidation state (charge) of an element.

Assigning oxidation states allows you to see which, if any, elements get…
- **Oxidized** (charge UP) or
- **Reduced** (charge DOWN)

Pure elements ALWAYS have oxidation states of 0!

Example:

Balanced rxn: $2CO + O_2 \rightarrow 2CO_2$

Oxidation states: $2\overset{+2\ -2}{CO} + \overset{0}{O_2} \rightarrow 2\overset{+4\ -2}{CO_2}$

Charge of C^{2+} goes up to C^{4+}: **C^{2+} is oxidized**
Charge of O^0 goes down to O^{2-}: **O^0 is reduced**

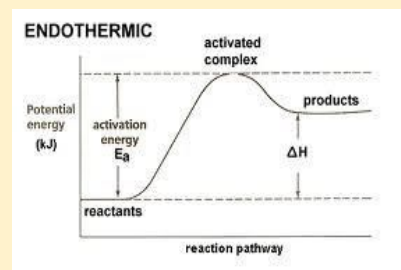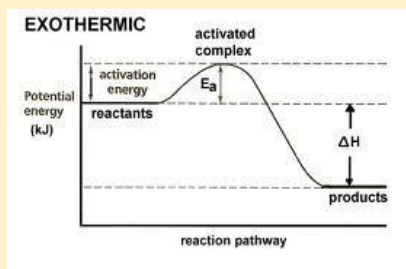# Chemistry Unit 5 Fact/Concept/Skill List *(Changes in Rxns)*

1. Be able to identify chemical vs. physical **changes** and identify chemical vs. physical **properties**.
   (***commonly missed = a substance dissolving in water is a PHYSICAL change)
2. For a reaction to occur you need lots of **effective collisions** *(correct angle, fast speed, lots of force),* to break bonds and allow new bonds to form. *(***KEY: ⇑ = increase)*
3. A **faster reaction rate** means a reaction occurs in a **smaller** amount of **time**.
4. To **increase rxn rate** you need *more frequent* & *more energetic (effective)* collisions: **(1)** ⇑ **concentration** or number of particles **(2)** ⇑ **temp.** which increases KE/speed of particles **(3)** ⇑ **surface area** by crushing which increases # particles exposed **(4)** add a **catalyst** to lower the activation energy
5. **Oxidation state** =*"charge"* of species.
6. **Species** = an atom, ion or molecule that is undergoing change in a chemical rxn
7. **Rules for assigning oxidation #s:** free elements = 0; group 1 = +1; group 2 = +2; O almost always -2; group 17 almost always -1; neutral compounds add to 0, polyatomics add to their charge; may need an equation to calculate elements with various possible charges
8. **REDOX** means reduction-oxidation reactions, which have a CHANGE in oxidation state; one species will increase and one will decrease. **# of e- lost = # of e- gained** (in any REDOX rxn)
9. **Reduction** means a **decrease in charge** (by gaining negative electrons) G.E.R.
10. **Oxidation** means an increase in charge (by losing negative electrons)   L.E.O.
11. To state what's oxidized or reduced, MUST give the species & the CHARGE (like Ca^0 or Cl^-)
12. **How to know "Is it a REDOX rxn?":** (1) assign all charges, (2) if species change charge from reactant side to product side, it's REDOX! *(Tip: look for free elements!)*
13. **exothermic** rxns: feel hot, give off heat, heat/nrg is product, **-△H** in **Table I** *(hi to low PE diagram)*
14. **endothermic** rxns: feel cold, take in heat, heat/nrg is reactant, **+△H** in **Table I** *(low to hi PE diagram)*
15. **PE diagrams** (*both endo* and *exo*): be able to locate **PEP** and **PER, activation nrg, Heat of Reaction (△H),** **PE of activated complex** (know **△H = PEP - PER** and know the **SHAPE** of endo/exo)
16. **Activation Energy** = from start (PEP) up to top of hill *(top of hill = "activated complex")*
17. **low energy = more stable** *(exo rxns produce more stable products b/c heat nrg was given off)*
18. **Entropy** = **disorder**; systems move toward higher entropy;   *lower* sol→liq→aq→gas *higher*
19. OVERALL: systems tend to move towards **low energy** and **high entropy** *(think teen bedroom)*
20. Conservation of… **CHARGE and MASS and ENERGY** *(can't be destroyed or made from nothing, just moved around - in all reactions they are "conserved", meaning total before = total after)*
21. **EASY Conservation of Mass problems:** total grams of reactants **=** total grams of products
22. **HARDER Conservation of Mass problems:** coefficients give **mole ratios** for calculating moles
23. **5 Reaction Types**

Synthesis: A + B → C (creating one new)	**Single Replacement**: A + BC → B + AC (one swap, homewrecker)
Decomposition: C → B + A (breaking one apart)	**Double Replacement**: AB + CD → AD + CB (two swaps, wife swap) ****** these are NEVER redox******
*Combustion: CxHy + O_2 → CO_2 + H_2O (a type of decomp., always produces water & carbon dioxide, has oxy & CxHy as reactants)	

24. **Old Stuff to Know:**
    - **balancing** rxns
    - calculate **gram-formula mass**
    - names of phase changes (sublimating? boiling? condensing? melting? freezing?, etc.)

EXOTHERMIC

activated complex

activation energy

E_a

Potential energy (kJ)

reactants

△H

products

reaction pathway

ENDOTHERMIC

activated complex

products

Potential energy (kJ)

activation energy

E_a

△H

reactants

reaction pathway

UNIT 6 SUMMARY: <u>SOLUTIONS</u>

Electrolytes

Electrolytes can <u>conduct electricity</u> when <u>in their aqueous form</u> by releasing <u>mobile ions</u>.

<u>**Acids**</u> release the H^+ (a.k.a. H_3O^+) ion in water
- Common acids are listed on **Table K**

<u>**Bases**</u> release the OH^- ion in water
- Common bases are listed on **Table L**

<u>**Salts**</u> (soluble ionic compounds) release ions into a solution
- Use **Table F** to see if it's soluble!
- More ions = better electrolyte!

Solids = no *mobile* ions = <u>no conductivity</u>

Electrolyte <u>solutions</u> (aq) are **homogeneous <u>mixtures</u>**, requiring *distillation* or *evaporation* to be separated (NO FILTERING ☹)

Acids and Bases

Acids release the H^+/H_3O^+ ion in solution while bases release the OH^- ion in solution. Their relative strength is indicated using the <u>pH scale</u>.

<u>**Indicators**</u> give us info about relative pH
- **Table M** describes the color of different indicators as pH changes from 1 to 14.
- <u>**Acids**</u> have pH values below 7
- <u>**Bases**</u> have pH values above 7
- <u>**Neutral**</u> solutions have a pH of 7.

One "jump" on the pH scale means a 10x change in H^+ concentration
(10x if going down, 1/10 if going up in pH)

Neutralization/D.R. Reactions

Acids and bases react with one another in a type of <u>double-replacement reaction</u> to make H_2O and a salt.

__ (H)Acid + __ Base(OH) → __ H_2O + __ Salt

Neutralized solutions have *equal* concentrations of H^+ and OH^- ions.

Like all double replacement reactions, the neutralization equation must show...
- **Conservation of charge** by making sure all compounds are neutral (*charges/ oxidation numbers* cancel)
- **Conservation of mass** by making sure *coefficients* balance the equation

Double replacement reactions between conductive salts might create a **precipitate**: an insoluble solid (check **Table F**)

Math: IDing and Counting Ions

Electrolytes dissociate when in water, generating a particular number of <u>cations</u> and <u>anions</u>.

<u>**Cations**</u> (+ charged) might be a...
- Metal from the Periodic Table
- Positive ion from Table E

<u>**Anions**</u> (- charged) might be a...
- Nonmetal from the Periodic Table
- Negative ion from Table E

Example:
$$Ca(SO_3) \rightarrow 1Ca^{2+} + 1SO_3^{2-}$$

Producing a total of 1 + 1 = 2 ions

Math: Concentration (M)

Molarity (M) is a ratio that describes the "strength" of a solution: how much of a <u>solute</u> is dissolved into a particular amount of <u>solution</u>.

$$Molarity\ (M) = \frac{moles\ of\ solute}{Liter\ of\ solution}$$

- Moles may need to be converted from grams using $mol = \frac{g}{gfm}$
- Milliliters (mL) may need to be converted into liters (L) using the ratio

$$1\ L = 1000\ mL$$

Math: Titrations

<u>Titrations</u> allow us to calculate the <u>concentration</u> of an acid or base using the concept of <u>neutralization</u>.

$$M_A V_A = M_B V_B$$

- M represents the **molarity** of the acid (A) or base (B)
- V represents the **volume** of the acid (A) or base (B) in *either* mL or L

The titration equation takes into account that some acids/bases are stronger than others and are "better" at neutralizing their "opposites."

UNIT 7: ATOMS AND THE PERIODIC TABLE, LEVEL 2

Models of the Atom

Many scientists contributed over a long period of time to our __modern__ understanding of atomic structure

Dalton	Discovery: **Indivisible** atoms
Thomson	Discovery: **Electron** Experiment: "**Cathode rays**" were deflected by electric and magnetic fields Model: "plum pudding"
Rutherford	Discovery: **Nucleus**, proton, **empty space**! Experiment: Alpha particles shot at a layer of **gold foil** mostly passed through, but some were deflected by a nucleus
Bohr	Discovery: Fixed **orbits** (energy levels) for electrons
Schrödinger	Discovery: **orbitals** (regions of probable location of electrons) Model: wave-mechanical, quantum, electron cloud

Subatomic Particles and Atoms

An atom's structure and properties come from its number and location of __protons__, __neutrons__, and __electrons__.

Knowing about an atom's subatomic particles allows you to calculate other information:

- **Mass number** = protons + neutrons
- **Atomic number** = protons
- **Net charge** = protons – electrons
- **Nuclear charge** = protons

Periodic Trends

When atoms are arranged on the Periodic Table in order of increasing atomic number, repeating patterns appear within groups and across periods.

Atomic radius (Table S): measure in *m* or *pm*
Electronegativity (Table S): how strongly an atom of an element __attracts electrons in a__ __chemical bond__

Trends related to electronegativity and atomic radius:

First ionization energy: energy required to __remove__ one electron from the valence shell
Metallic character: "willingness" to __lose__ an electron
Chemical reactivity: Atoms with the same number of valence electrons react in chemically similar ways

Average Atomic Mass

An element's average atomic mass is a __weighted__ __average__ of all of its __naturally occurring isotopes__.

$$AAM = (mass)\left(\frac{\%}{100}\right) + (mass)\left(\frac{\%}{100}\right) + \cdots$$

- **Isotopes** have the same number of protons and electrons, but different numbers of neutrons
- The average atomic mass will be closest in number to the __most__ __abundant isotope__ of that element

Excited Electrons and Spectra

We can use Bohr's model of the atom to explain the unique color spectra released by each element.

- An atom's **ground state** is the lowest-energy electron configuration (shown on the P.T.)
- An **excited state** adds energy to ONE electron and promotes it to a higher energy level (unstable)

Allotropes

Elements can exist in multiple forms in the same phase; each "form" is an allotrope with different properties as a result of a different molecular structure.

- O_2 and O_3 (ozone) have different molecular structures, giving them different chemical and physical properties
- Solid carbon can take the form of coal, graphite, diamond, or large fullerenes depending on the connectivity of its covalent bonds
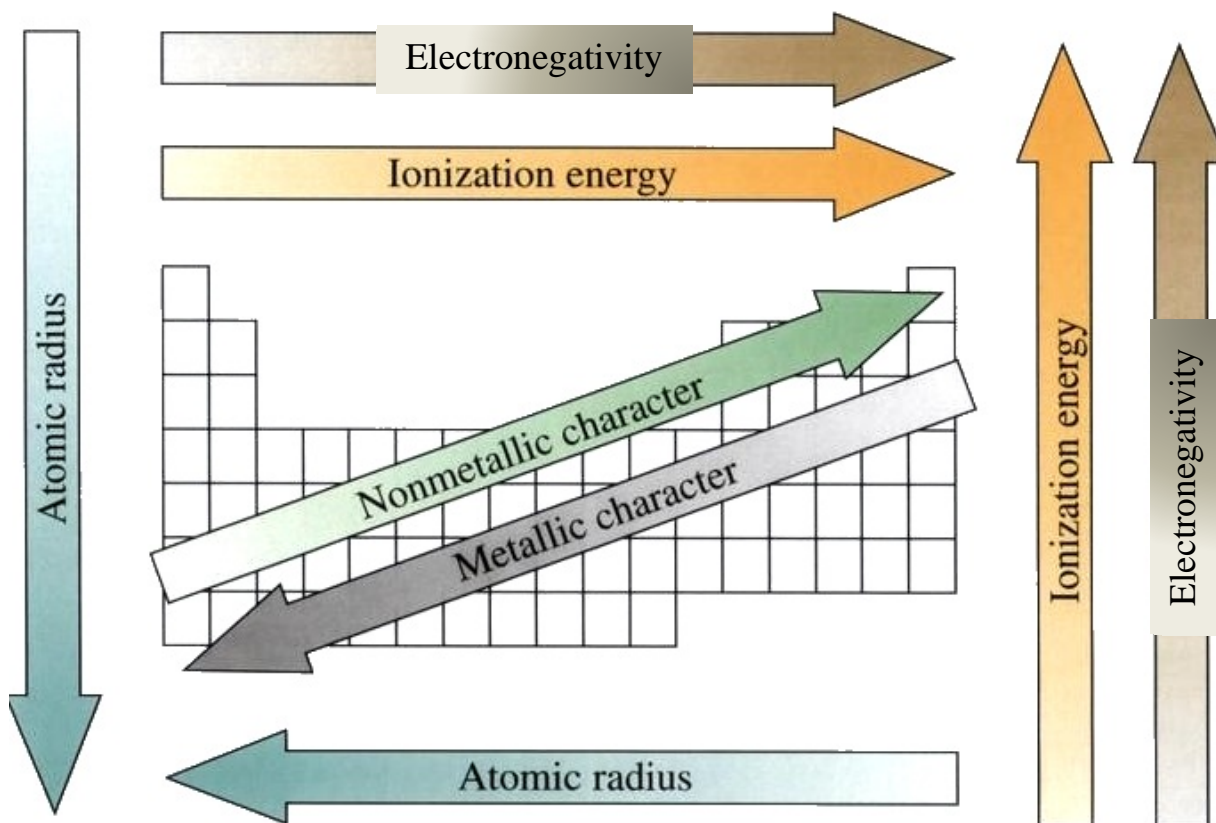
PERIODIC TABLE TRENDS SUMMARY

IMPORTANT TERMS TO DESCRIBE WHY THESE TRENDS EXIST:

Moving across a period:

- **Nuclear charge**: number of protons in the nucleus (increases from left to right)
  - o Positive charge attracts negatively charged electrons
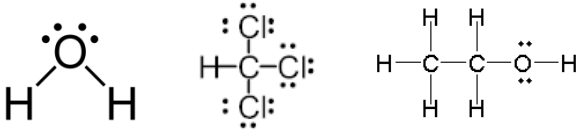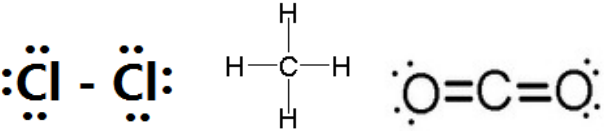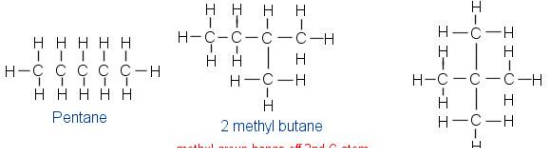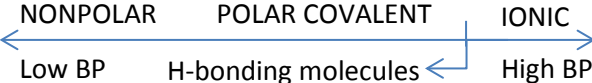
Moving down a group:

- **Electron shells:** energy levels where electrons hang out
  - o Adding electron shells (moving down a group) makes an atom bigger and moves valence electrons farther away from the positively charged nucleus

Electronegativity

Ionization energy

Atomic radius

Nonmetallic character

Metallic character

Atomic radius

Ionization energy

Electronegativity

- Positive ions (cations) are always smaller than their neutral atoms.
  - o Electrons are taken away
- Negative ions (anions) are always bigger than their neutral atoms.
  - o Electrons are added

UNIT 8: BONDING AND HYDROCARBONS, LEVEL 2

Types of Bonds

Bonds can be classified as <u>ionic</u>, <u>polar covalent</u>, <u>nonpolar covalent</u>, or <u>metallic</u> based on the types of elements involved in the bond.

Ionic BONDS: M-NM
- High difference in <u>electronegativity</u> (EN)
- <u>MOST POLAR</u> type of bond
- <u>Cation</u> and <u>anion</u> formed

Covalent BONDS: NM-NM
- **Polar covalent**: $\Delta EN > 0$
  - Electrons move towards the element with the higher EN
- **Nonpolar covalent**: $\Delta EN = 0$
  - Electrons are shared equally

Metallic "BONDS": M only
- "Sea" of mobile electrons
- Makes metals conductive and malleable

Types of Molecules

Within a molecule, electrons can be distributed <u>symmetrically</u> or <u>asymmetrically</u> in chemical bonds.

Polar MOLECULES: asymmetrical (electrons favor one side of the molecule over the other)
- 0 or 1 line of symmetry in molecule

Nonpolar MOLECULES: symmetrical (electrons are distributed evenly throughout the molecule)
- 0 or 1 line of symmetry
- Bonds within the molecule could be polar or nonpolar

Naming Hydrocarbons

Tables P and Q provide prefixes and suffixes for naming saturated and unsaturated hydrocarbons.

Saturated hydrocarbons contain only single bonds
- **Alkanes**: C_nH_{2n+2}

Unsaturated hydrocarbons contain at least one double or triple bond
- **Alkenes**: C_nH_{2n} (one double bond)
- **Alkynes**: C_nH_{2n+2} (one triple bond)

Isomers of a molecule have the same **molecular formula**, but a different structural formula (and a different IUPAC name)

Pentane

2 methyl butane
- methyl group hangs off 2nd C atom
- longest chain is 4 C's long = butane

2,2 dimethyl propane
- 2 methyl groups both hang off 2nd C atom
- longest unbroken chain is 3 C's = propane

Intermolecular Forces of Attraction

A molecule's polarity and size determines the strength of the IMFs that act between molecules.

NONPOLAR   POLAR COVALENT   IONIC
Low BP   H-bonding molecules   High BP

- **Hydrogen bonds** are a super-strong IMF that act between <u>polar covalent</u> molecules that contain **H** and **N, O, or F**
- **Nonpolar hydrocarbons** have weak IMFs affected by mass/size
  - More carbons = higher BP
  - Less branching = higher BP
  - **Distillation** can separate!

Percent Composition by Mass

We can use molar masses to determine the % by mass of one element within a compound.

$$\% \ composition = \left(\frac{mass \ of \ part}{mass \ of \ whole}\right) * 100$$

For a compound...
- Mass of part = total mass of element
- Mass of whole = gram-formula mass

Example: % iron in iron (III) oxide

Formula = Fe_2O_3

$$\% \ Fe = \left(\frac{2*55.8}{(2*55.8)+(3*16)}\right)*100 = 69.9\%$$

Empirical Formulas

Empirical formulas represent the lowest ratio of element in a compound.

Molecular → Empirical: reduce subscripts by GCF

Ex: C_5H_{10} has empirical formula CH_2 (GCF = 5)

Empirical → Molecular: use molar mass to find a factor to multiply empirical subscripts by

Ex: Molecule with gfm of 82 g/mol and empirical formula C_3H_5
- C_3H_5 has mass of 41 g/mol
- 82 / 41 = 2 (GCF of molecular formula)
- Multiply subscripts by 2: C_3H_5 → C_6H_{10}

UNIT 9: PHYSICAL AND CHEMICAL CHANGE SUMMARY

Heat, Temperature, Phases

*Applying or removing heat changes a substance's particle arrangement, which changes its **physical** properties.*

Temperature: measurement of *average kinetic energy* (remains the same during phase change)

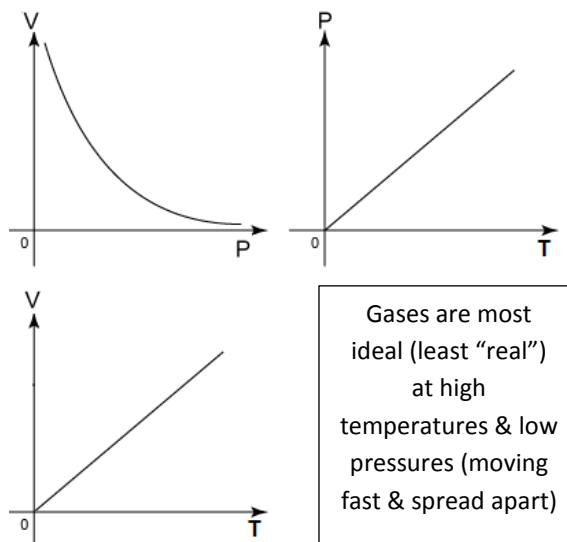**Heat**: thermal energy that moves from high temperatures to low temperatures
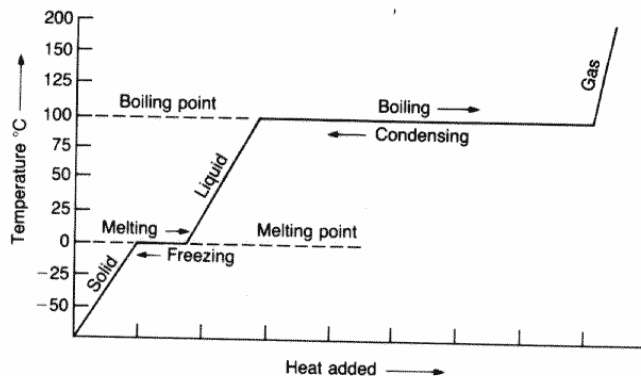
Don't forget sublimation (s→g) and deposition (g→s)!

Gas Laws/Relationships

Changes in gas particle movement result in measureable change.

Kinetic Molecular Theory says that **ideal** gases...

Gases are most ideal (least "real") at high temperatures & low pressures (moving fast & spread apart)
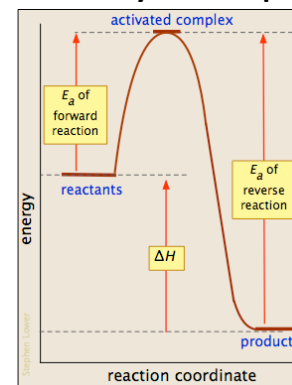
Chemical Change and Equilibrium

When particles collide and make/break new bonds, mass, energy, and charge are conserved.

Reactions are proportional "recipes"
- Coefficients represent **mole ratios** of different substances in the reaction

Some reactions are reversible
- In a **dynamic equilibrium** state:

→**Rate** of forward reaction = **rate** of reverse reaction
→ Concentration of reactants and Products remain **constant**
→Pressure and temperature are **constant**

Use Table I to find unknown **heats of reaction**

Math: Heat Calculations

Heat required to make a change can be calculated.

q = mCΔT | KINETIC ENERGY CHANGES

q = mH_f | POTENTIAL ENERGY CHANGES
- Melting or freezing (no temp change)

q = mH_v | POTENTIAL ENERGY CHANGES
- Boiling or condensing (no temp change)

➢ C, H_f, & H_v values for water are found on Table B.
➢ Relatively high C, H_f, or H_v values indicate *strong* intermolecular forces (IMFs)

Moles to kJ or kJ to moles: Use the equation given in the problem or Table I. *Example:*
2H₂ + O₂ →2H₂O + 571.6 kJ (1 mol H₂O = 285.8 kJ)

Math: Combined Gas Law

Changes in P, V, and T are proportional.

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

Temperature MUST be in Kelvin!

If a variable is not mentioned or is said to remain constant, you do not have to include it in your calculation.

Temperature and pressure at **STP** are recorded on Table A.

Stressing Equilibrium

LeChatelier's Principle predicts the direction a reversible reaction will shift in response to an applied stress.

↑ **Concentration**: shift to "use up" added particles
- More particles = more collisions

↑ **Temperature**: shift away from side with "+ heat"
- Faster moving particles = more collisions for endothermic process

↑ **Pressure**: shift to side with fewer moles of gas
- More pressure = more collisions where there are more gas particles

UNIT 10: SOLUTIONS SUMMARY

Electrolytes

Electrolytes can <u>conduct electricity</u> when <u>in their aqueous form</u> by releasing <u>mobile ions</u>.

Acids release/<u>DONATE</u> the **H⁺** (a.k.a. H_3O^+) ion
- Common acids are listed on **Table K**

Bases release the **OH⁻** ion in water (alternatively, *ACCEPT the H⁺ ion*)
- Common bases are listed on **Table L**

Conductive salts release soluble ions into a solution (+ and – "parts")
- Solubility of an ionic compound can be determined using **Table F**

Solids (insoluble) = no *mobile* ions = <u>no conductivity</u> (CAN be filtered)

Electrolyte <u>solutions</u> are **homogeneous mixtures**, requiring *distillation* or *evaporation* to be separated (NO FILTERING ☹)

Acids and Bases

Arrhenius acids release the H^+/H_3O^+ ion in solution while Arrhenius bases release the OH⁻ ion in solution. Their relative strength is indicated using the <u>pH scale</u>.

Indicators give us information about the relative pH of a solution
- **Table M** describes the color of different indicators as pH changes from 1 to 14.
- **Acids** have pH values below 7 (more H⁺)
- **Bases** have pH values above 7 (less H⁺)
- **Neutral** solutions have a pH of 7. (H⁺=OH⁻)
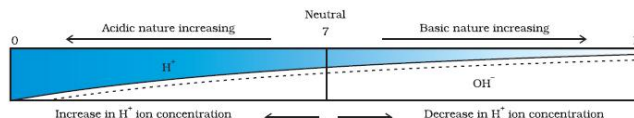
One "jump" on the pH scale indicates a 10x change in H⁺ (H_3O^+) concentration
More H⁺ = More acidic

Polarity: Like Dissolves Like

Polar <u>solvents</u> are best able to dissolve polar <u>solutes</u>.
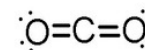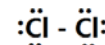
Polar MOLECULES: asymmetrical (uneven distribution of electrons)
- 0 or 1 line of symmetry in molecule
- Stronger IMFs
- Higher boiling points
- Lower vapor pressure
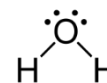
Nonpolar MOLECULES: symmetrical (electrons are distributed evenly throughout the molecule)
- 2 or more lines of symmetry
- Weaker IMFs
- Lower boiling points
- Higher vapor pressure

Strongest IMFs = hydrogen bonding, where polar molecules have H bonded to F, O, or N

Math: Concentration (M, ppm, %)

Concentration describes the "strength" of a solution: how much of a <u>solute</u> is dissolved into a particular amount of <u>solution</u>.

$$Molarity\ (M) = \frac{moles\ of\ solute}{Liter\ of\ solution}$$

- Moles may need to be converted from grams using $mol = \frac{g}{gfm}$
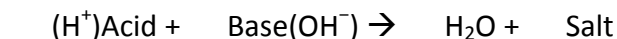- 1000 Milliliters (mL) = 1 Liter (L)

$$ppm = \frac{g\ solute}{g\ solution} * 1{,}000{,}000$$

$$\%\ by\ mass = \frac{g\ solute}{g\ solution} * 100$$

Math: Titrations

<u>Titrations</u> allow us to calculate the <u>concentration</u> of an acid or base using the concept of <u>neutralization</u>.

Because acids neutralize bases:

___ (H⁺)Acid + ___ Base(OH⁻) → ___ H_2O + ___ Salt

We can use known concentrations of an acid or base to figure out an unknown concentration of its "opposite"

$$M_A V_A = M_B V_B$$

- M represents the **molarity** of the acid (A) or base (B)
- V represents the **volume** of the acid (A) or base (B) in *either* mL or L

Solubility

Table G shows how much of a particular <u>solute</u> can dissolve in a specific (100 g) <u>amount of water</u> at a particular <u>temperature.</u>
(Solids increase, gases decrease.)

- Dissolving particles (many ions or 1 full covalent molecule) in water **increases the boiling pt** and **decreases the freezing pt**.
- More particles/higher M = bigger effect on bp/fp.

Unsaturated solutions: more solute can be dissolved (below Table G line)

Saturated solutions: the ideal amount of solute is dissolved (on Table G line)

Supersaturated solutions: "too much" solute has been dissolved (above Table G line)

UNIT 11: REDOX/ELECTROCHEMISTRY SUMMARY

## Identifying Redox Reactions

In a redox reaction, <u>electrons are transferred</u> from one species to another, resulting in changes in oxidation #.

<u>Redox reactions</u> are changes in which one species is *oxidized* and another is *reduced*

- ALL **single replacement** reactions are redox reactions.
- Double replacement reactions are <u>NEVER</u> redox reactions.
- **ALL changes** have conservation of mass, energy, and charge!

Example:

Oxidation states: $2\overset{+2\ -2}{CO} + \overset{0}{O_2} \rightarrow 2\overset{+4\ -2}{CO_2}$

Charge of C^{2+} goes up to C^{4+}: **C^{2+} is oxidized**

Charge of O^0 goes down to O^{2-}: **O^0 is reduced**
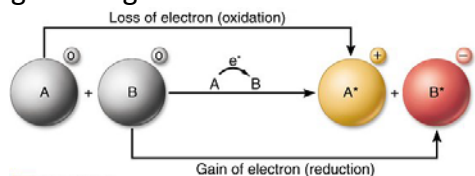
Half-Reactions and Net Ionic Equations

Half-reactions show oxidation and reduction processes separately. Combining half-reactions gives a net ionic equation where <u>the number of electrons lost is equal to the number of electrons gained</u>

<u>Example:</u> An <u>unbalanced</u> redox reaction is shown.

$$Zn + Ag(NO_3) \rightarrow Ag + Zn(NO_3)_2$$

Oxidation half-reaction: Zn^0 is oxidized to Zn^{2+}
$$Zn^0 \rightarrow Zn^{2+} + 2e^-$$

Reduction half-reaction: Ag^+ is reduced to Ag^0
$$Ag^+ + 1e^- \rightarrow Ag^0$$

To make e^- lost = e^- gained (conservation of charge), the whole reduction half-reaction needs to be multiplied by 2. The equations can then be added to give a net ionic equation:

$$Zn^0 + 2Ag^+ \rightarrow 2Ag^0 + Zn^{2+}$$

Voltaic Cells

Electrochemical cells that <u>spontaneously</u> convert <u>chemical energy into electrical energy</u> are called voltaic cells (used as batteries).

AN OX and a BIG RED CAT
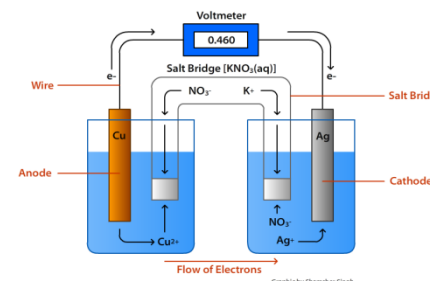
- Oxidation occurs at the anode (mass lost)— **more active** metal is the anode!
- Reduction occurs at the cathode (mass gained)
- e^- flow from the anode to the cathode (through the wire!)
- A **salt bridge** allows for the migration of ions

Oxidation and Reduction

The "transfer" of electrons results in changes in oxidation number (charge) of an element.

Assigning oxidation states allows you to see which, if any, elements get…

- **Oxidized** (charge UP) or
- **Reduced** (charge DOWN)

LEO↑ says **GER**↓ | **L**oss of **e**⁻ is **o**xidation
Gain of **e**⁻ is **r**eduction

OIL↑ **RIG**↓ | **O**xidation **i**s **l**oss (of e⁻)
Reduction **i**s **g**ain (of e⁻)

Spontaneity and Activity

<u>Table J</u> can be used to predict the reactivity of different elements.
*The most active elements are found at the **<u>top</u>** of Table J*

<u>More active **solid metal atoms**…</u>
- Are more likely to react by…
  - Losing electrons/getting oxidized
- Will spontaneously replace less active metal ions in a single replacement rxn

<u>Example</u>: Zn + Ag(NO₃) *will* react *spontaneosuly* because Zn is more active than Ag^+

<u>More active **nonmetal molecules**…</u>
- Are more likely to react by…
  - Gaining electrons/getting reduced
- Will spontaneously replace less active nonmetal ions in a single replace. rxn

Electrolytic Cells

Electrochemical cells that use batteries to <u>non-spontaneously</u> convert <u>electrical energy into chemical energy</u> are called electrolytic cells.
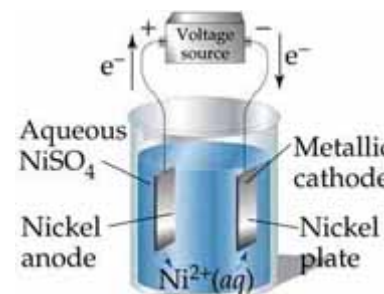
AN OX and a **BIG RED CAT** still applies, but an external **power source** is needed to oxidize the *less* active metal.
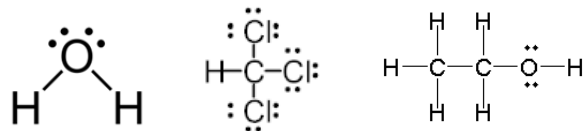
Electron flow is still from anode to cathode!

UNIT 12: ORGANIC CHEMISTRY, LEVEL 2

Types of Molecules

*Within a molecule, electrons can be distributed symmetrically or asymmetrically in chemical bonds.*

Polar MOLECULES: asymmetrical (electrons favor one side of the molecule over the other)
- 0 or 1 line of symmetry in molecule

Nonpolar MOLECULES: symmetrical (electrons are distributed evenly throughout the molecule)
- 2 or more lines of symmetry
- Bonds within the molecule could be polar or nonpolar
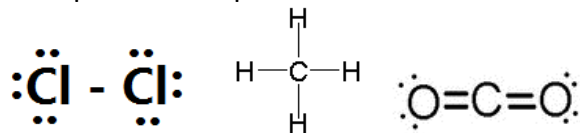
Naming Hydrocarbons

Tables P and Q provide prefixes and suffixes for naming saturated and unsaturated hydrocarbons.
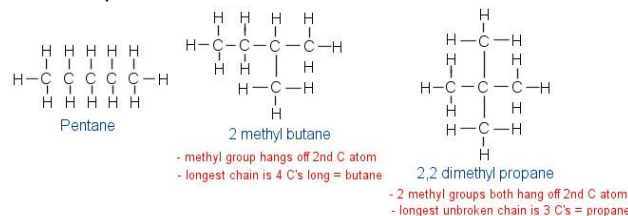
Saturated hydrocarbons contain only single bonds
- **Alkanes**: C_nH_{2n+2}

Unsaturated hydrocarbons contain at least one double or triple bond
- **Alkenes**: C_nH_{2n} (one double bond)
- **Alkynes**: C_nH_{2n-2} (one triple bond)

Isomers of a molecule have the same **molecular formula**, but a different structural formula

Pentane

2 methyl butane
- methyl group hangs off 2nd C atom
- longest chain is 4 C's long = butane

2,2 dimethyl propane
- 2 methyl groups both hang off 2nd C atom
- longest unbroken chain is 3 C's = propane

Empirical formulas show the lowest ratio of elements in a compound

Functional Groups

Distinct groups of atoms can be attached to organic molecules to give them similar chemical properties.

Any-carbon functional groups:
- **Halides** (-X where X is from Group 17)
- **Alcohols** (-OH)
- **Ketones** (-COC-)
- **Amines** (-NH$_2$)

Last-carbon functional groups:
- **Aldehydes** (-CHO)
- **Organic acids** (-COOH)
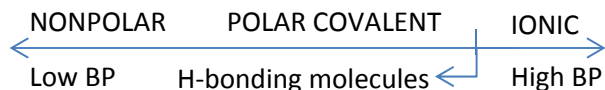- **Amides** (-CONH$_2$)

Interrupting the carbon-chain functional groups:
- **Ethers** (-O-)
- **Esters** (-COO-)

All classes, functional groups, and sample names/structures are on **Table R**!

Intermolecular Forces of Attraction

A molecule's polarity and size determines the strength of the IMFs that act between molecules.

NONPOLAR   POLAR COVALENT   IONIC

Low BP   H-bonding molecules   High BP

- **Hydrogen bonds** are strong IMFs that act between polar covalent molecules that contain **H** bonded to **N, O, or F**
- **Nonpolar hydrocarbons** have weak IMFs affected by mass/size
  - More carbons = higher BP
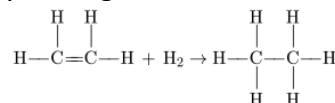 - Less branching = higher BP

Organic Reaction Types

Organic molecules react in predictable ways.

Substitution: Switch out a hydrogen for a halogen

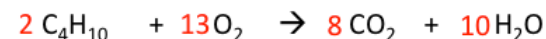$$CH_3CH_3 + Cl_2 \rightarrow CH_3CH_2Cl + HCl$$
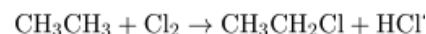
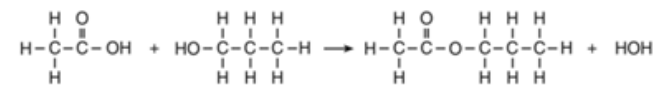Addition: Make a single bond out of a multiple bond by adding on a diatomic molecule

Polymerization: Make big molecules (polymers) out of the same small unit (monomer)

Combustion: Burn an organic molecule using O_2 to make CO_2 and H_2O

$$2\,C_4H_{10} + 13\,O_2 \rightarrow 8\,CO_2 + 10\,H_2O$$

Esterification: Make an ester and H_2O by smashing together an alcohol and an organic acid

Fermentation: Use an enzyme to make alcohol and CO_2 from sugar

Saponification: Make soap!
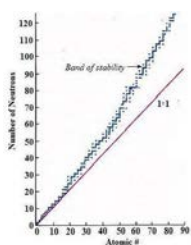
UNIT 13: NUCLEAR SUMMARY

Isotopes and Radioisotopes

The stability of a nucleus depends on its ratio of neutrons to protons.

Isotopes: atoms with the same number of **protons**, but different numbers of **neutrons**

Mass # = protons + neutrons → $^{14}_{6}C$

Atomic # = protons →

Atomic *mass*: *weighted average* mass of all *naturally occuring* isotopes of an element

Radioisotope: isotope with **unstable** *ratio of neutrons to protons* (& is therefore radioactive)
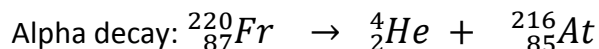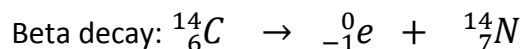
Natural Transmutation

Radioisotopes spontaneously decay by releasing particles from their nuclei, "transforming" their identities.

Table N tells us the particle that each radioactive nuclide emits when it decays

When a nuclide decays, the total mass and nuclear charge must remain the same from reactants to products:

Ex: Different types of decay equations

Beta decay: $^{14}_{6}C \rightarrow ^{0}_{-1}e + ^{14}_{7}N$

Alpha decay: $^{220}_{87}Fr \rightarrow ^{4}_{2}He + ^{216}_{85}At$

Positron decay: $^{137}_{55}Cs \rightarrow ^{0}_{1}e + ^{137}_{54}Xe$

Artificial Transmutation

We can "force" a stable nuclide to go undergo transmutation

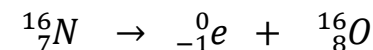- 2 reactants: stable nucleus bombarded by a nuclear particle

$$^{4}_{2}He + ^{14}_{7}N \rightarrow ^{17}_{8}O + ^{1}_{1}H$$
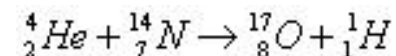
Note that total mass and nuclear charge still remain the same from reactants to products

Contrast:
Natural transmutation: spontaneous change
- 1 reactant: unstable nucleus

$$^{16}_{7}N \rightarrow ^{0}_{-1}e + ^{16}_{8}O$$

Fission, Fusion, & Risks/Benefits

*Fractional amounts of **mass** are converted into **huge** amounts of **energy** by **E = mc^2***

Fission: division of heavy nuclide
$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{141}_{56}Ba + ^{92}_{36}Kr + 3 \, ^{1}_{0}n$$
- Used for nuclear power (controlled) and nuclear weapons (uncontrolled)

Fusion: union of light nuclei
$$^{2}_{1}H + ^{2}_{1}H \rightarrow ^{4}_{2}He$$
- No radioactive byproducts!
- Occurs in the sun
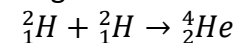
Benefits: nuclear energy/power, medical interventions (KNOW: ^{131}I & ^{123}I, ^{60}Co & ^{99}Tc), dating living & non-living things (KNOW: ^{14}C to ^{12}C and ^{238}U to ^{206}Pb), etc.
Risks: biological exposure/death, nuclear waste

Nuclear Particles

Table O lists a number of particles that can be released as radiation from an unstable nucleus or used in artificial processes.

Particles with *smaller* mass and charge have *higher* **penetrating power.**

Alpha	Positron	Beta	Gamma
$^{4}_{2}He$	$^{0}_{+1}e$	$^{0}_{-1}e$	$^{0}_{0}\gamma$

Half-Life

A radioisotopes takes a set amount of time to lose half of its remaining mass: its half life.

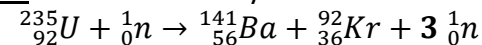At a regular interval (half-life, often on Table N), 50% of the remaining mass will be lost.

Ex: How long will it take for a 24-g sample of ^{131}I to decay to 3 grams? (HL of ^{131}I found on **Table N**)

Days	HLs	Grams
0	0	24
8	1	12
16	2	6
24 days	3	3