Year End Condensed Honors Chemistry Review

Important Notes: This handout is a summary of the concepts. Practice items for all concepts have not been provided in this packet. Additional practice may also be found in the Chemistry EOC Review Practice Handout, Example Problems by Goal prepared for the Foldable Project, website links, and in practice tests provided by your teacher. The majority of this information can be found on the teacher websites. I also recommend reviewing the chapter summaries for any chapters referenced in this handout. ***For Overall Vocab and Scientists/Atomic Theory review, see the separate handouts devoted to these topics!!***

****Be sure to practice the NCSCOS Goal 1 Questions Provided by your teacher as separate practice items for Goal 1 are not posted on the NCDPI website!!****

Unit 1 - Chapters 1 & 2; Laboratory Safety, Laboratory Equipment & Scientific Method, Dimensional Analysis, SI, etc.

Significant Figures (Rules in textbook p. 47):

If the decimal is PRESENT
Start at the Pacific.

Come to the first real digit and count all remaining digits

Ex.
a. 32.02 _________
b. 0.00235 _______

If the decimal is NOT PRESENT
Start at the Atlantic.

Come to the first real digit and count all remaining digits

Ex.
a. 42500 _________
b. 620350 ________

Calculations with Significant Figures

Addition or Subtraction with Significant Figures

- When adding or subtracting decimals, the answer must have the same number of digits to the right of the decimal point as there are in the measurement having the fewest digits to the right of the decimal point.

Limit and round your answer to the least number of decimal places in any of the numbers that make up your answer.

123.25 mL + 46.0 mL + 86.257 mL = 255.507 mL

The answer is expressed as 255.5 mL since 46.0 mL has only one decimal place.

Multiplication and Division with Significant Figures

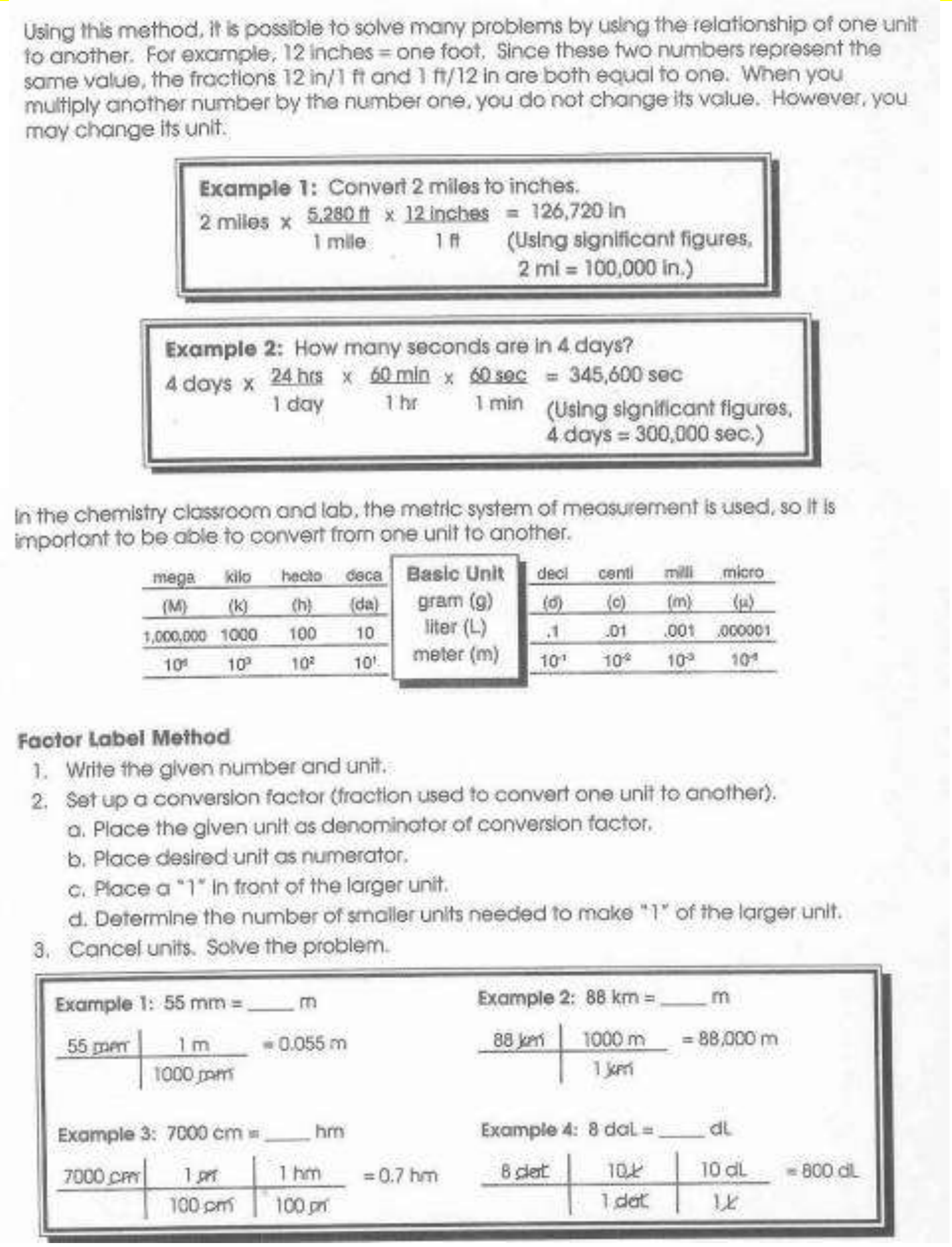
- For multiplication or division, the answer can have no more significant figures than are in the measurement with the fewest number of significant figures.

Limit and round to the least number of sig figs in any of the values.

23.0 cm x 432 cm x 19 cm = 188,784 cm^3

The answer is expressed as 190,000 cm^3 since 19 cm has only two sig figs.

Dimensional Analysis – Factor Label Method and the Metric System



Using this method, it is possible to solve many problems by using the relationship of one unit to another. For example, 12 inches = one foot. Since these two numbers represent the same value, the fractions 12 in/1 ft and 1 ft/12 in are both equal to one. When you multiply another number by the number one, you do not change its value. However, you may change its unit.

Example 1: Convert 2 miles to inches.

$2\text{ miles} \times \frac{5{,}280\text{ ft}}{1\text{ mile}} \times \frac{12\text{ inches}}{1\text{ ft}} = 126{,}720\text{ in}$

(Using significant figures, 2 mi = 100,000 in.)

Example 2: How many seconds are in 4 days?

$4\text{ days} \times \frac{24\text{ hrs}}{1\text{ day}} \times \frac{60\text{ min}}{1\text{ hr}} \times \frac{60\text{ sec}}{1\text{ min}} = 345{,}600\text{ sec}$

(Using significant figures, 4 days = 300,000 sec.)

In the chemistry classroom and lab, the metric system of measurement is used, so it is important to be able to convert from one unit to another.

mega	kilo	hecto	deca	**Basic Unit**	deci	centi	milli	micro
(M)	(k)	(h)	(da)	gram (g)	(d)	(c)	(m)	(µ)
1,000,000	1000	100	10	liter (L)	.1	.01	.001	.000001
10^6	10^3	10^2	10^1	meter (m)	10^{-1}	10^{-2}	10^{-3}	10^{-6}

Factor Label Method

1. Write the given number and unit.
2. Set up a conversion factor (fraction used to convert one unit to another).
   a. Place the given unit as denominator of conversion factor.
   b. Place desired unit as numerator.
   c. Place a "1" in front of the larger unit.
   d. Determine the number of smaller units needed to make "1" of the larger unit.
3. Cancel units. Solve the problem.

Example 1: 55 mm = _____ m

$\begin{array}{c|c} 55\,\cancel{\text{mm}} & 1\text{ m} \\ \hline & 1000\,\cancel{\text{mm}} \end{array} = 0.055\text{ m}$

Example 2: 88 km = _____ m

$\begin{array}{c|c} 88\,\cancel{\text{km}} & 1000\text{ m} \\ \hline & 1\,\cancel{\text{km}} \end{array} = 88{,}000\text{ m}$

Example 3: 7000 cm = _____ hm

$\begin{array}{c|c|c} 7000\,\cancel{\text{cm}} & 1\,\cancel{\text{m}} & 1\text{ hm} \\ \hline & 100\,\cancel{\text{cm}} & 100\,\cancel{\text{m}} \end{array} = 0.7\text{ hm}$

Example 4: 8 daL = _____ dL

$\begin{array}{c|c|c} 8\,\cancel{\text{daL}} & 10\,\cancel{\text{L}} & 10\text{ dL} \\ \hline & 1\,\cancel{\text{daL}} & 1\,\cancel{\text{L}} \end{array} = 800\text{ dL}$

Density Calculations: **Know how to do density by volume displacement and how to use Density as a Conversion Factor!**

density = mass/volume; $D = \frac{m}{V}$ **Be able to solve for any variable!**

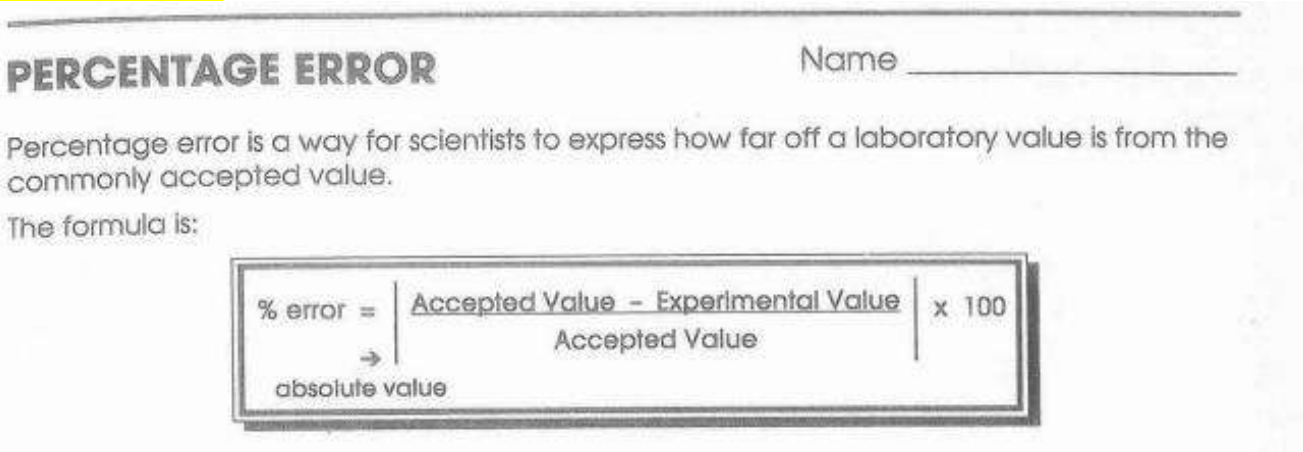
Accuracy & Precision

Accuracy refers to the closeness of measurements to the correct or accepted value of the quantity measured.
examples: baseball pitcher throwing strikes; basketball going in the hoop; lab data is the correct answer

Precision refers to the closeness of a set of measurements of the same quantity made in the same way.
examples: baseball pitcher throwing strikes in the same location or keeps throwing balls in the same location; basketball shots are all net every time or basketball shots are missed by bouncing of the rim in the same location; lab data give the same results over and over (possibly right or wrong)

****Questions regarding understanding of precision/accuracy are usually given in the form of interpreting data tables – practice those Goal 1 Questions mentioned above!!****

Percentage Error



PERCENTAGE ERROR Name ____________

Percentage error is a way for scientists to express how far off a laboratory value is from the commonly accepted value.

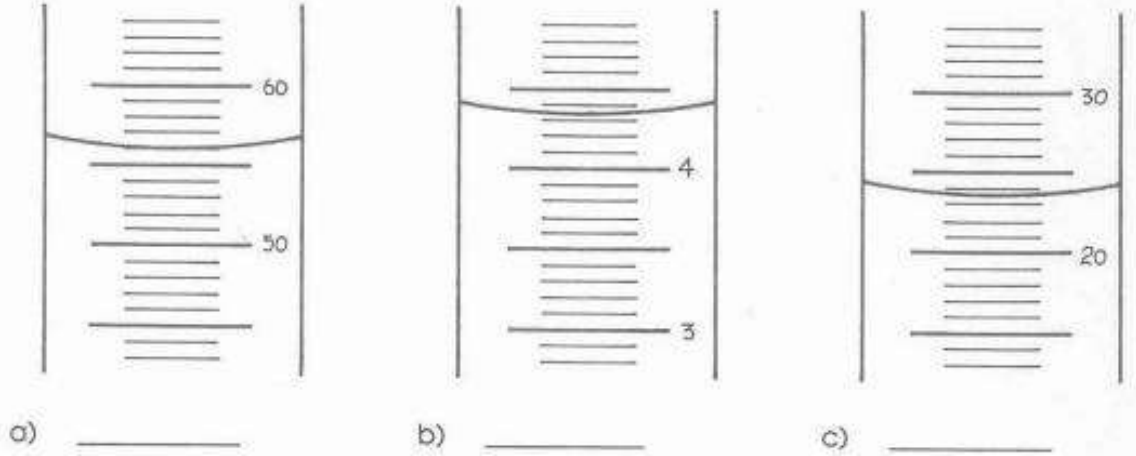
The formula is:

$$\boxed{\%\ \text{error} = \left| \frac{\text{Accepted Value} - \text{Experimental Value}}{\text{Accepted Value}} \right| \times 100}$$

→ absolute value

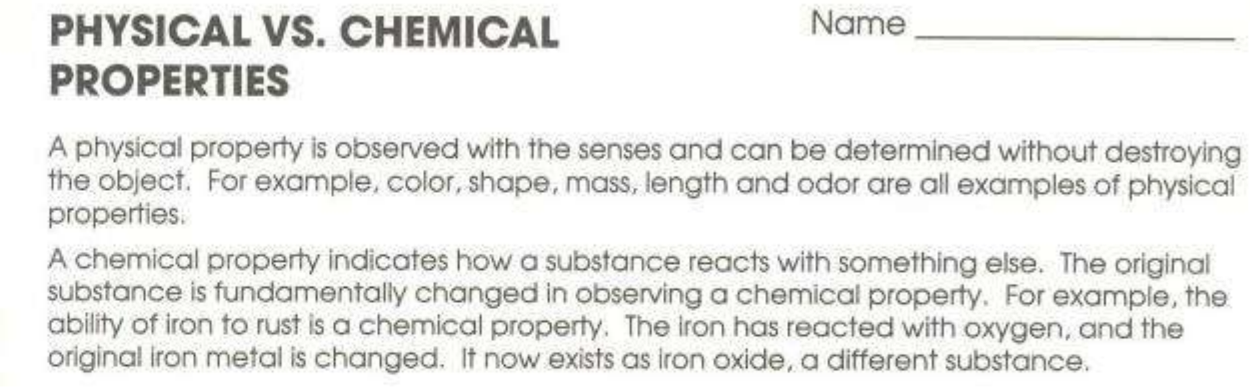
Practicing Measuring Liquid Volume

What volume is indicated on each of the graduated cylinders below? The unit of volume is mL.



Unit 2 - Selected Topics: Chapters 1, 3, & 4: Bohr's Model, Matter, Atomic Structure

1. Differentiate between a chemical and physical change.
   physical – does not involve a change in the identity of the material or substance
   chemical – occurs when one or more substances are converted into different substances



PHYSICAL VS. CHEMICAL PROPERTIES

Name ____________________

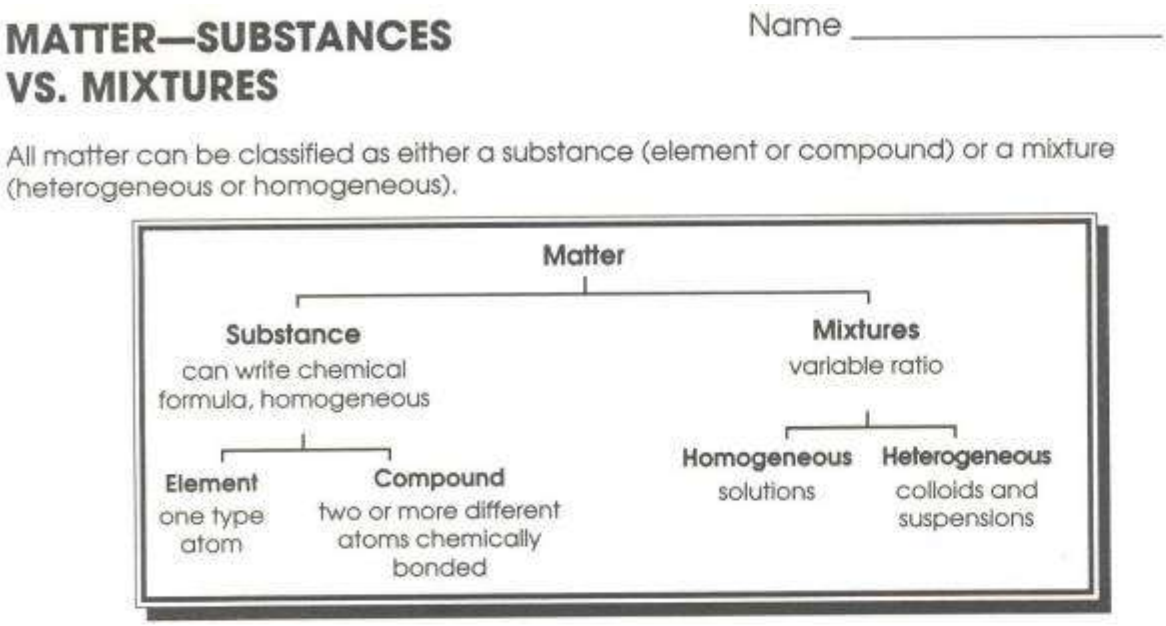
A physical property is observed with the senses and can be determined without destroying the object. For example, color, shape, mass, length and odor are all examples of physical properties.

A chemical property indicates how a substance reacts with something else. The original substance is fundamentally changed in observing a chemical property. For example, the ability of iron to rust is a chemical property. The iron has reacted with oxygen, and the original iron metal is changed. It now exists as iron oxide, a different substance.

The four signs of a chemical change are:

- change in temperature
- formation of a gas
- formation of a precipitate (a solid from two solutions)
- a color change (sometimes☺)

2. Describe the 3 states of matter.
   - Solid – definite shape and volume – particles packed very closely together
   - Liquid – no definite shape; definite volume – particles packed more loosely but still "together"
   - Gas – no definite shape; no definite volume – particles "far apart"
3. How is a pure substance different from a mixture?
   A pure substance has a fixed composition; every sample has the exact same characteristic properties and every sample has the exact same chemical composition.
   Element – pure substance of only one kind of atom
   Compound – substance is made from the atoms of two or more elements that are chemically bonded



MATTER—SUBSTANCES VS. MIXTURES

Name ____________________

All matter can be classified as either a substance (element or compound) or a mixture (heterogeneous or homogeneous).

Atomic Theory Development

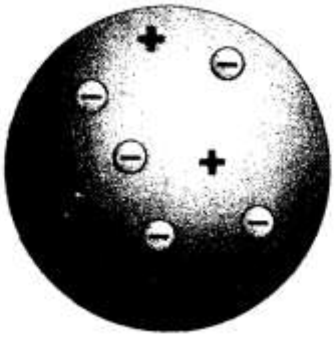
I. Democritus (400B.C.)

First used term “atom”

II. Dalton’s model (1803)

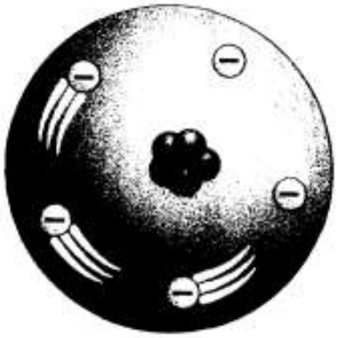
Credited with modern atomic theory
4 major ideas

III. Goldstein Discovers Protons (1886)



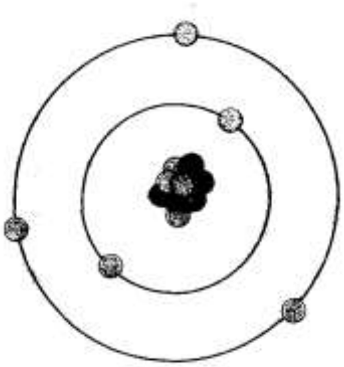
IV. Thomson's model (1897)

Credited with Discovery of the electron



V. Rutherford's model (1911)

-“gold foil” experiment
-Proposed the atom had a nucleus (mass concentrated in the nucleus) and that an atom is mostly empty space



VI. Bohr's model (1913)

- Electrons move in fixed orbits

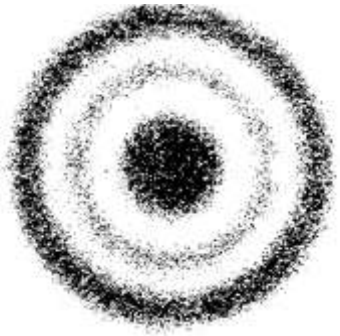
-Bohr's Model used for wavelength and frequency calculations
Model only works for the hydrogen (H) atom

VII. Millikan discovers electron charge and Mass of an electron (1909)

VIII. Chadwick discovers Neutron (1932)

IX. De Broglie (Frenchman) proposes particle wave behavior of Electron also known as particle-wave duality ((1923)

X. Schrodinger writes an equation to determine probability of electron location (quantum theory)



XI. Electron-cloud model (present)

Others:
Planck described packets of energy call quanta

Einstein described the photoelectric effect and the wave-particle duality of radiation (act as a wave and a particle) – deBroglie also is credited with the latter.

Atomic Structure

Atomic # (z) = # of protons

- Identifies (ID's) an element – UNIQUE for each element
- Elements in order on the periodic table by atomic #
- Because atoms must be neutral; (z) also = #e^-

Mass # - total # of p^+ and n^o

Ions = Charged Atoms

- charge occurs from gaining or losing e^- (<u>not</u> p^+)

(+) = lost e^- (−) = gained e^-

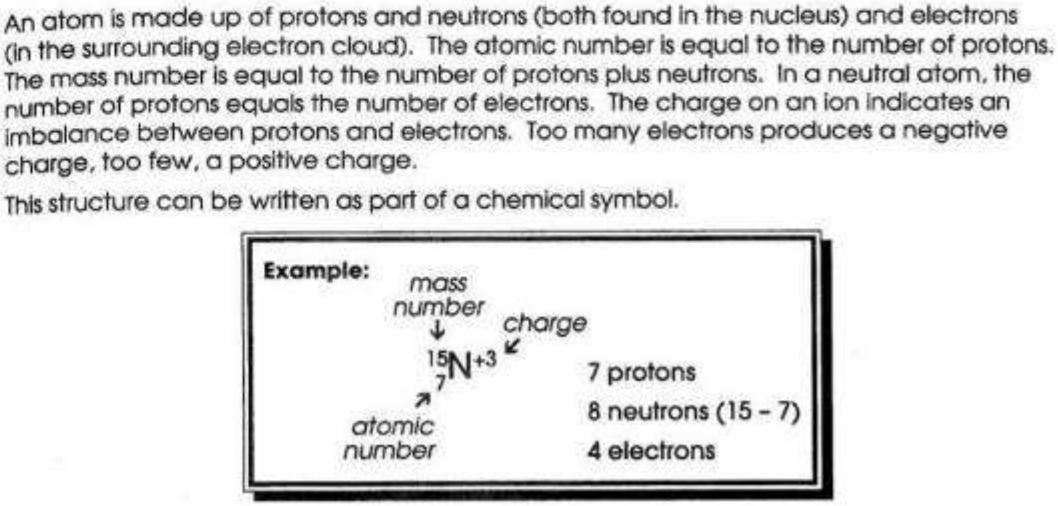
Isotopes and Average Atomic Mass

Isotope = (nuclide)

2 ways to represent isotopes: hyphen notation or nuclear symbol

$$\text{Uranium-235} \quad \text{or} \quad \overset{\text{mass \#}}{\uparrow}\ {}^{235}_{92}\text{U} \rightarrow \text{element symbol}$$

↓ element ↓ mass # ↓ atomic #



An atom is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud). The atomic number is equal to the number of protons. The mass number is equal to the number of protons plus neutrons. In a neutral atom, the number of protons equals the number of electrons. The charge on an ion indicates an imbalance between protons and electrons. Too many electrons produces a negative charge, too few, a positive charge.

This structure can be written as part of a chemical symbol.

Example:

mass number ↓ charge ↙

$^{15}_{7}N^{+3}$

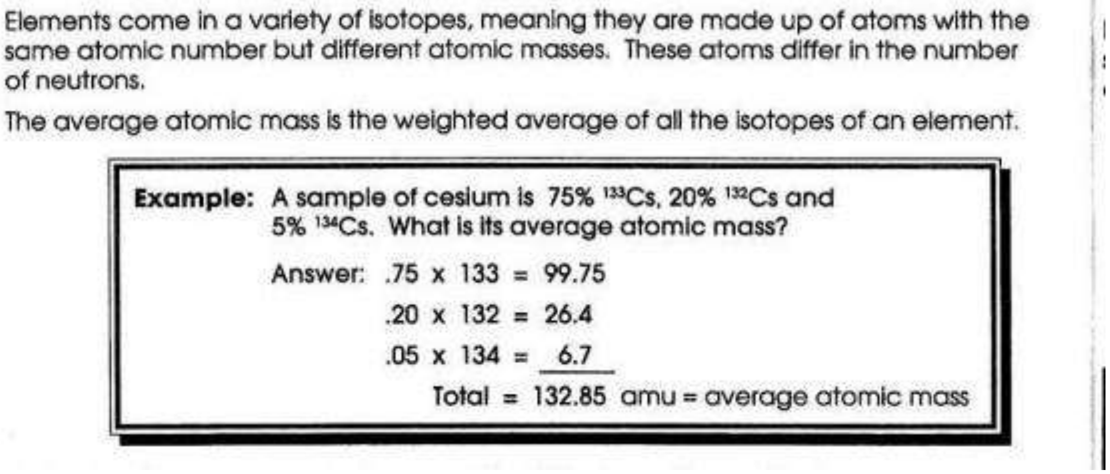
↗ atomic number

7 protons

8 neutrons (15 – 7)

4 electrons

Isotopes & Average Atomic Mass



Elements come in a variety of isotopes, meaning they are made up of atoms with the same atomic number but different atomic masses. These atoms differ in the number of neutrons.

The average atomic mass is the weighted average of all the isotopes of an element.

Example: A sample of cesium is 75% ^{133}Cs, 20% ^{132}Cs and 5% ^{134}Cs. What is its average atomic mass?

Answer: .75 x 133 = 99.75

.20 x 132 = 26.4

.05 x 134 = 6.7

Total = 132.85 amu = average atomic mass

Bohr's Model Summary: Practice!!

When to use it:

- When the problem has wording about e^-s jumping from one level to another, e.g., n=4 to n=2.

How to use it:

- Use it to find the wavelength of the electromagnetic radiation in nm or m (must use m in calculations!)

Know Relationships between Energy (E), Wavelength (λ), and Frequency (v):

Long Wavelength → Lower Frequency → Lower Energy
Short Wavelength → Higher Frequency → Higher Energy

$c = \lambda \upsilon$ and $E = h\upsilon$

Unit 3: Chapters 4 & 5: Atomic Structure, Electron Configuration, Periodicity

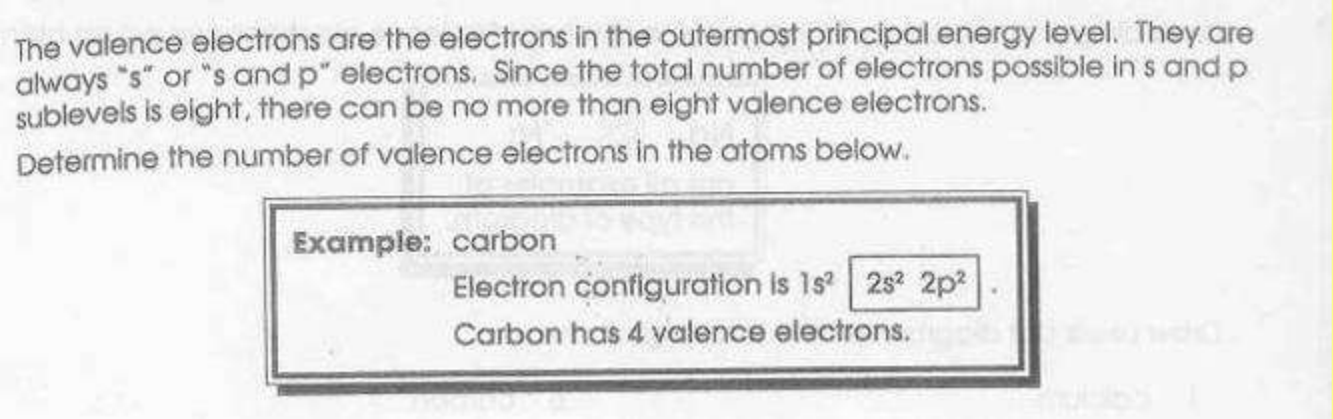
Review & Practice Electron Configuration (Full & Noble Gas)

s,p,d,f orbitals – shapes of s & p, how many e^-s can each sublevel hold, etc.

s orbital – sphere shaped

p orbital – dumbbell shaped

Valence Electrons: Be able to count them from the electron configuration and by looking at the periodic table.



The valence electrons are the electrons in the outermost principal energy level. They are always "s" or "s and p" electrons. Since the total number of electrons possible in s and p sublevels is eight, there can be no more than eight valence electrons.

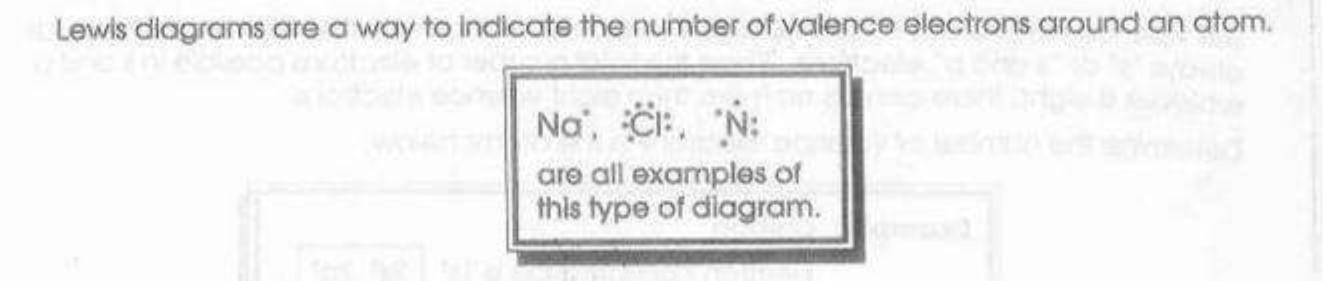
Determine the number of valence electrons in the atoms below.

Example: carbon

Electron configuration is $1s^2$ $\boxed{2s^2\ 2p^2}$.

Carbon has 4 valence electrons.

Lewis Dot Diagrams

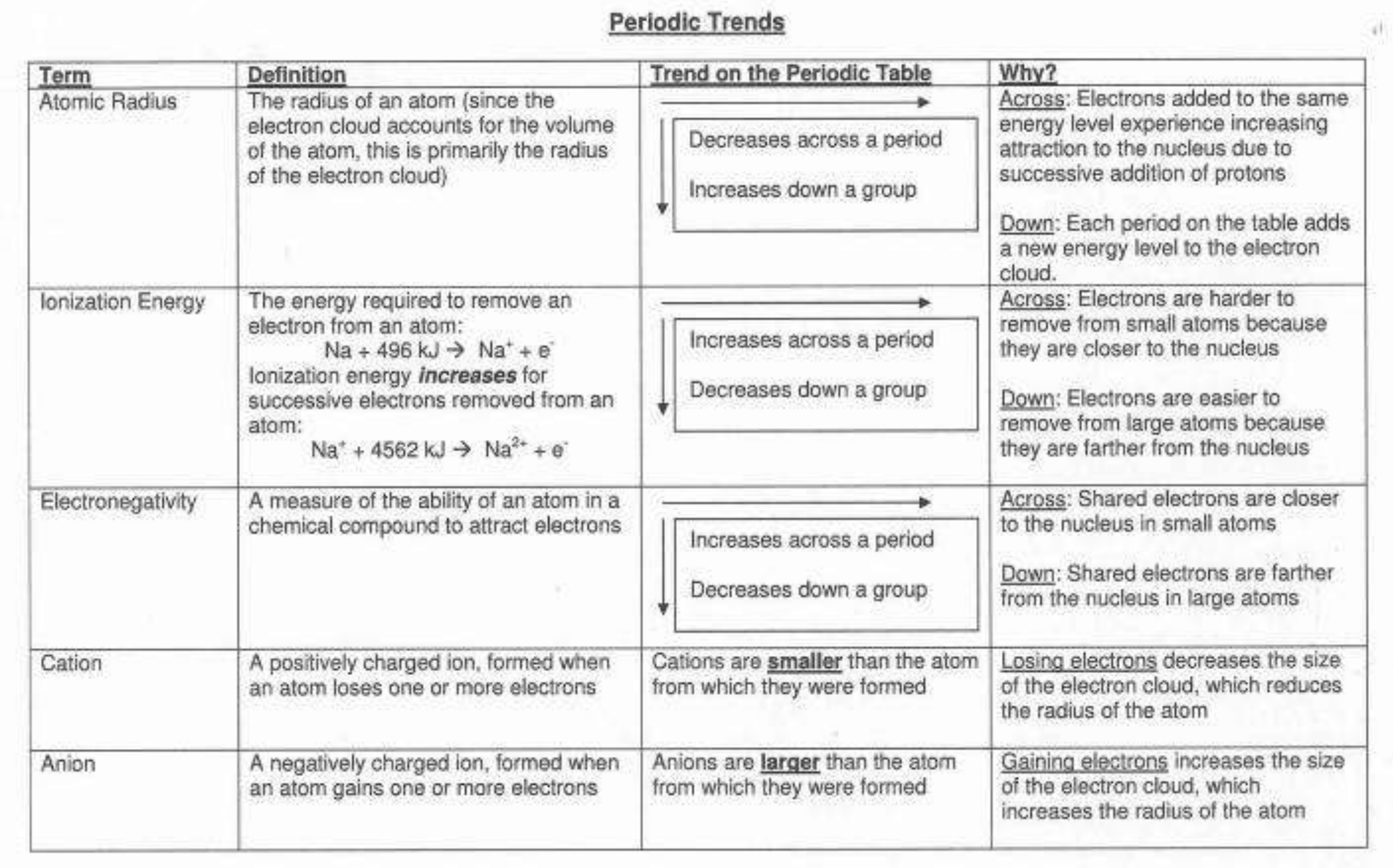


Lewis diagrams are a way to indicate the number of valence electrons around an atom.

Na·, :Cl:, ·N:

are all examples of this type of diagram.

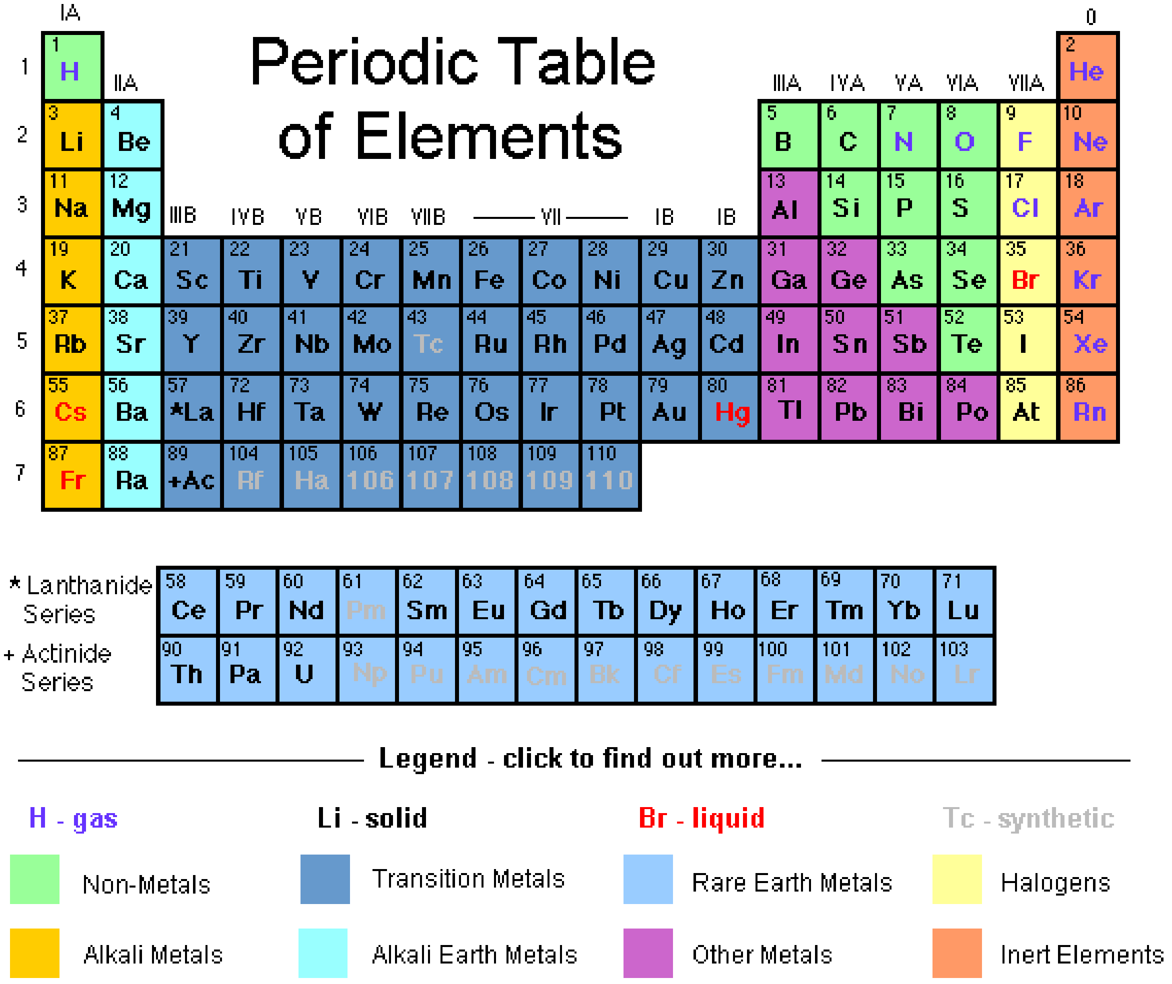
Periodic Trends & Properties – Practice!



Periodic Trends

Term	Definition	Trend on the Periodic Table	Why?
Atomic Radius	The radius of an atom (since the electron cloud accounts for the volume of the atom, this is primarily the radius of the electron cloud)	→ ↓ Decreases across a period Increases down a group	Across: Electrons added to the same energy level experience increasing attraction to the nucleus due to successive addition of protons Down: Each period on the table adds a new energy level to the electron cloud.
Ionization Energy	The energy required to remove an electron from an atom: $Na + 496\ kJ \rightarrow Na^{+} + e^{-}$ Ionization energy ***increases*** for successive electrons removed from an atom: $Na^{+} + 4562\ kJ \rightarrow Na^{2+} + e^{-}$	→ ↓ Increases across a period Decreases down a group	Across: Electrons are harder to remove from small atoms because they are closer to the nucleus Down: Electrons are easier to remove from large atoms because they are farther from the nucleus
Electronegativity	A measure of the ability of an atom in a chemical compound to attract electrons	→ ↓ Increases across a period Decreases down a group	Across: Shared electrons are closer to the nucleus in small atoms Down: Shared electrons are farther from the nucleus in large atoms
Cation	A positively charged ion, formed when an atom loses one or more electrons	Cations are **smaller** than the atom from which they were formed	Losing electrons decreases the size of the electron cloud, which reduces the radius of the atom
Anion	A negatively charged ion, formed when an atom gains one or more electrons	Anions are **larger** than the atom from which they were formed	Gaining electrons increases the size of the electron cloud, which increases the radius of the atom

Groups of the Periodic Table



Periodic Table of Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII			IB	IB	IIIA	IVA	VA	VIA	VIIA	0
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 106	107 107	108 108	109 109	110 110								

* Lanthanide Series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
+ Actinide Series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Legend - click to find out more...

H - gas	Li - solid	Br - liquid	Tc - synthetic
🟩 Non-Metals	🟦 Transition Metals	🟦 Rare Earth Metals	🟨 Halogens
🟨 Alkali Metals	🟦 Alkali Earth Metals	🟪 Other Metals	🟧 Inert Elements

Unit 4 - Chapter 7 – Nomenclature

Common Polyatomic Ions (most common listed from NCSCOS)

Acetate	$C_2H_3O_2{}^{1-}$	Ammonium	$NH_4{}^{1+}$
Carbonate	$CO_3{}^{2-}$	Nitrate	$NO_3{}^{1-}$
Sulfate	$SO_4{}^{2-}$		

Use reference tables for all others!!

Common Acids

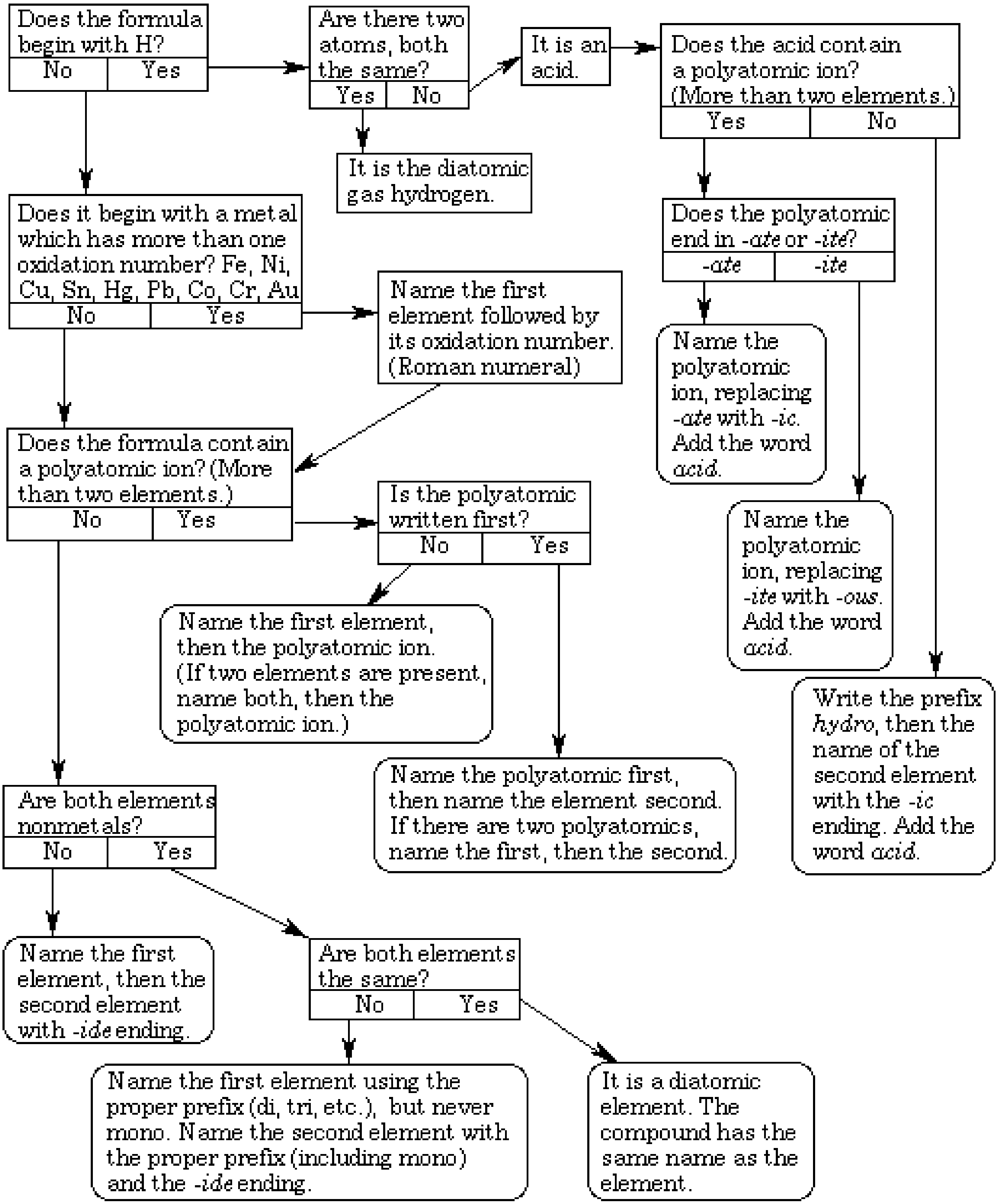
Hydrochloric Acid	HCl
Nitric Acid	HNO_3
Acetic Acid	$HC_2H_3O_2$ or CH_3COOH
Sulfuric Acid	H_2SO_4

Practice Naming & Formula Writing!!!

Be sure to notice if the compound is an acid, is ionic or is molecular before naming

Flow Chart for Naming Simple Inorganic Compounds

The flowchart is adapted from p. 131-132 of the February 1983 issue of the *Journal of Chemical Education*.



Unit 5 - Chapters 3 & 7 – The Mole – Math with Chemical Formulas

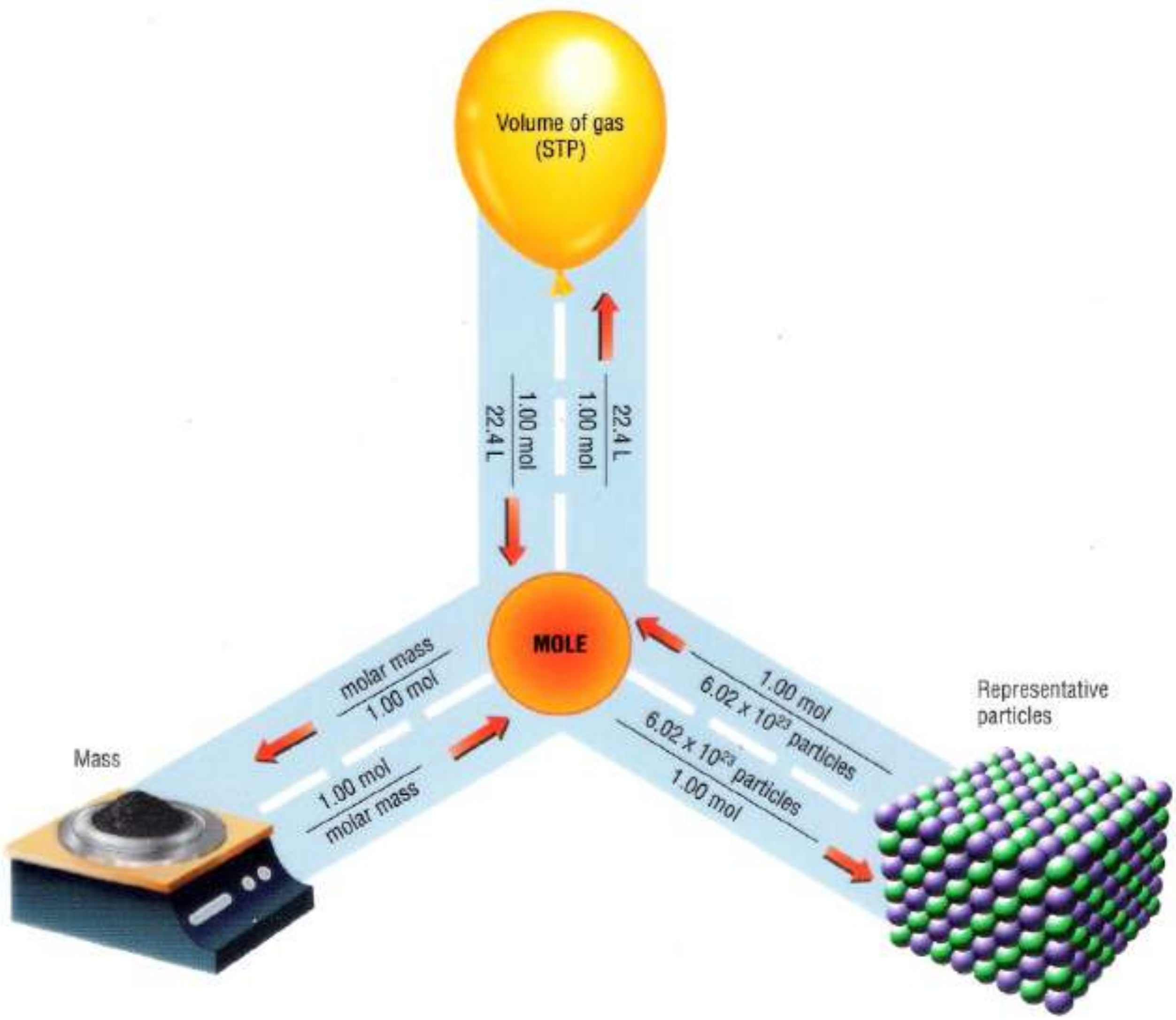
Avogadro's Number

One mole of a substance contains Avogadro's Number (6.02×10^{23}) of molecules.

Calculation of Molar Mass from atomic masses on the periodic table

Mole Conversions of All Types: (particle types: atoms, molecules, formula units, ions)

- g to mol
- mol to g
- particles to mol
- mol to particles
- g to particles
- particles to g



Practice Mole Conversion Problems!

Percent Composition

- Percent composition is the percent by mass of each element in a compound.
- Percent composition is the same, regardless of the size of the sample.

Percent Composition Calculations

$$\% \text{ comp} = \frac{\text{mass of element}}{\text{molar mass of cpd}} \times 100\% = \% \text{ element in the compound}$$

Empirical Formulas

- Empirical Formula = Simplest Formula

To find the empirical formula from data:

1. Assume 100% sample; change % to grams for each element ***(% to mass)***
2. Find moles from the grams of each element ***(mass to mol)***
3. Find the smallest whole # ratio by dividing by the smallest number of moles ***(divide by small)***
4. If necessary, multiply to get rid of fractions. ***(multiply to whole!)***

Molecular Formulas

- Molecular Formula = Actual Formula

Example:

C_2H_6	CH_3
molecular	empirical

$$MF = (EF)x \quad \text{where} \quad X = \frac{\text{Molecular mass}}{\text{Empirical mass}}$$

Moles in Solution (Molarity)

- **Molarity** is the term used for moles dissolved in solution
- Symbol for **Molarity = M**
- Definition – moles of solute per liter of solution
- Formula

$$M = \frac{\text{moles solute (mol)}}{\text{liter solution (L)}}$$

*****Be able to calculate Molarity or grams of solute needed.*****

Unit 6 - Chapter 6 – Bonding & Molecular Structures

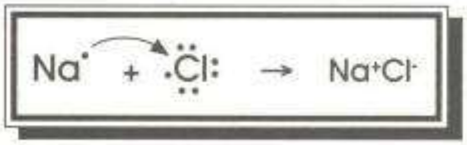
Bond Types Table

Bond Type	Type(s) of Atoms involved	Force	Properties	Examples
IONIC transfer of electrons	**Metal & Nonmetal**	Attraction between ions, opposite charges attract; transfer of electrons	High melting point Water soluble Crystalline Aquesous solutions conduct a current	NaCl MgO CaS
COVALENT Sharing of electrons	**Two Nonmetals** ***Polar*** = unequal sharing = partial charge ***Nonpolar*** = equal sharing = no charge	Sharing of electrons	Low melting point Brittle Nonconductors	Water CO_2 NH_3
METALLIC Free flow of Electrons	**Two Metals**	Sharing of electrons ***between all atoms***	Good conductors Malleable Ductile	Copper wire Iron bar

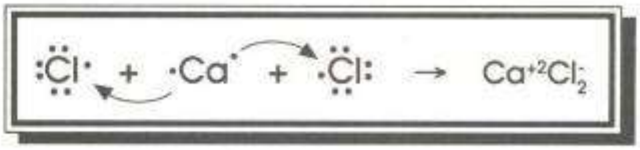
IONIC BONDING

Name ____________________

Ionic bonding occurs when a metal transfers one or more electrons to a nonmetal in an effort to attain a stable octet of electrons. For example, the transfer of an electron from sodium to chlorine can be depicted by a Lewis dot diagram.



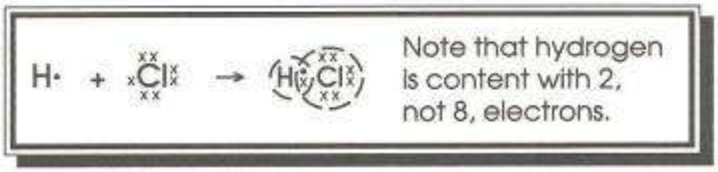
Calcium would need two chlorine atoms to get rid of its two valence electrons.



COVALENT BONDING

Name ____________________

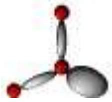
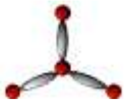
Covalent bonding occurs when two or more nonmetals share electrons, attempting to attain a stable octet of electrons at least part of the time. For example:



Note that hydrogen is content with 2, not 8, electrons.

*****Be able to use electronegativity values to determine bond type – see foldable! *****

VSEPR & Molecular Geometry

	Molecular Shape	Type of Molecule AB_yE_z	Atoms Bonded to Central Atom	Lone Pairs of e-s on Central Atom
Linear		AB_2	2	0
Bent		AB_2E	2	1
Trigonal Planar		AB_3	3	0
Tetrahedral		AB_4	4	0
Trigonal Pyramidal		AB_3E	3	1
Bent		AB_2E_2	2	2

Be able to draw the Lewis Dot Structures for the 7 diatomic molecules and know the # of bonds in each! (I_2, Br_2, Cl_2, F_2, O_2, N_2, H_2)

LEWIS STRUCTURE:

Element Symbol = nuclei and inner-shell electrons
Dashes = shared electron pairs in covalent bond
Dots = unshared electrons

Be able to draw Lewis Structure for molecules and determine bond & molecule polarity.

VSEPR Theory – be able to predict molecular shape

- valence-shell, electron pair repulsion
- Way to predict molecular geometry (shape)
- There is a repulsion between valence e^- pairs

Unit 7 - Chapter 8 – Chemical Equations

BALANCING EQUATIONS: 4 steps:

1. Start with a word equation
2. Convert to a formula equation (don't forget the diatomic molecules!)
3. Balance with coefficients:
    - balance each atom one at a time
    - balance polyatomic ions on each side of the equation as one unit
    - balance H and O last (they often appear in more than one compound)
4. Check; if coefficients are not the lowest possible; reduce down.

REACTION TYPES – Be able to Use Reference Tables – the equation types and subtypes are in there!

You must be able to predict products and write balanced chemical equations.

1. **Synthesis:** (or composition):
   2 or more substances combine to form 1 new substance

$$A + X \rightarrow AX$$

2. **Decomposition:**
   A single substance produces 2 or more simpler substances

$$AX \rightarrow A + X$$

3. **Single Displacement (Replacement):**
   1 element replaces a similar element in a cmpd

$$A + BX \rightarrow AX + B$$

Hint: In reactions write water as HOH.

4. **Double Displacement (Replacement):**
   The ions of 2 cmpds switch places to form 2 new cmpds.

$$AX + BY \rightarrow AY + BX$$

5. **Combustion:**
   When a substance combines with oxygen releasing a large amount of energy in the form of light and heat.

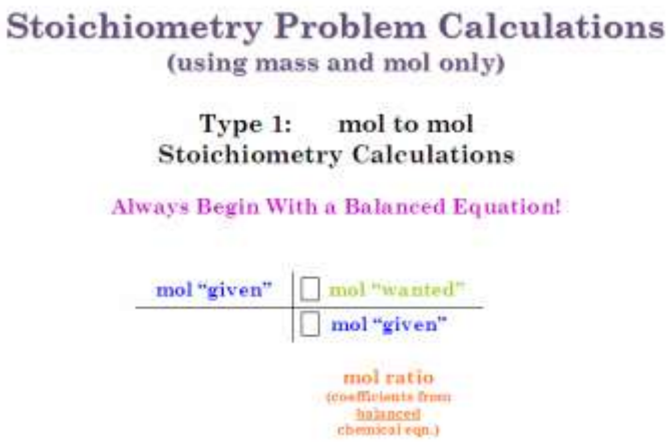
Often combustions involve a hydrocarbon : cmpd containing C and H

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O$$

Unit 8 Chapter 9 – Stoichiometry

You must be able to do stoichiometric calculations by using the mol ratios from a balanced chemical equation!

mol to mol:



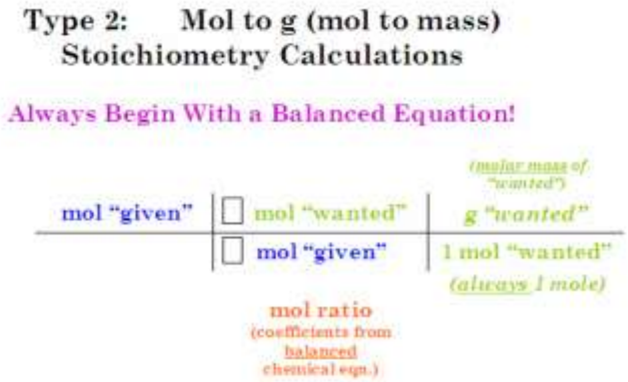
Stoichiometry Problem Calculations
(using mass and mol only)

Type 1: mol to mol
Stoichiometry Calculations

Always Begin With a Balanced Equation!

mol "given"	☐ mol "wanted"
	☐ mol "given"
	mol ratio (coefficients from balanced chemical eqn.)

mol to g:

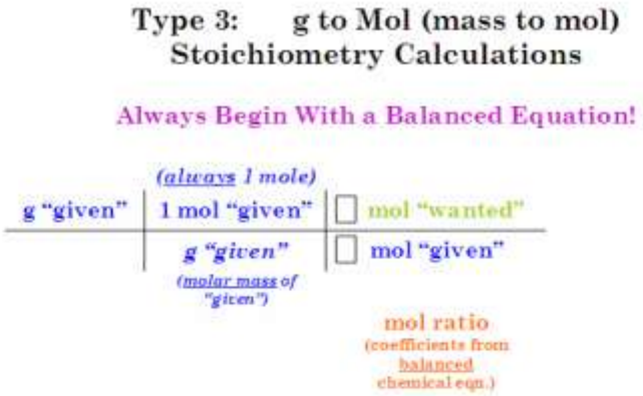


Type 2: Mol to g (mol to mass)
Stoichiometry Calculations

Always Begin With a Balanced Equation!

		(molar mass of "wanted")
mol "given"	☐ mol "wanted"	g "wanted"
	☐ mol "given"	1 mol "wanted"
	mol ratio (coefficients from balanced chemical eqn.)	(always 1 mole)

g to mol:

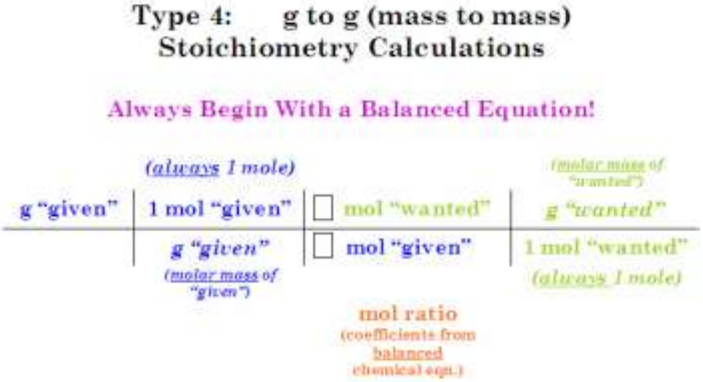


Type 3: g to Mol (mass to mol)
Stoichiometry Calculations

Always Begin With a Balanced Equation!

	(always 1 mole)	
g "given"	1 mol "given"	☐ mol "wanted"
	g "given"	☐ mol "given"
	(molar mass of "given")	mol ratio (coefficients from balanced chemical eqn.)

g to g:

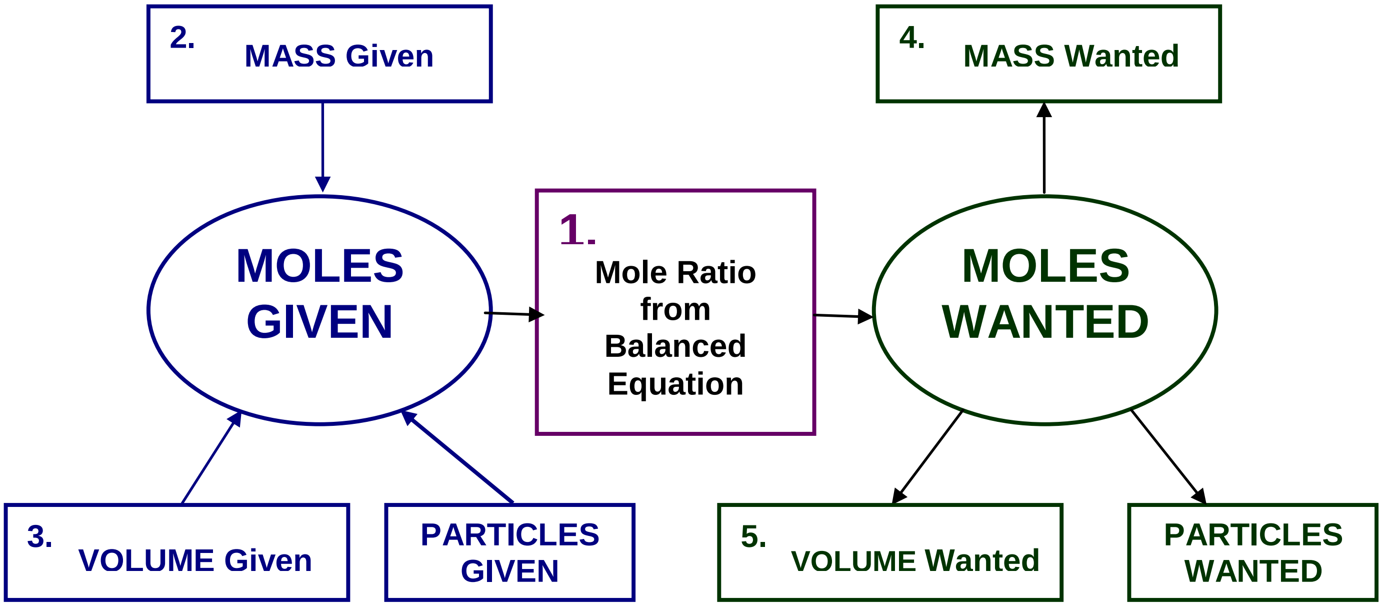


Type 4: g to g (mass to mass)
Stoichiometry Calculations

Always Begin With a Balanced Equation!

	(always 1 mole)		(molar mass of "wanted")
g "given"	1 mol "given"	☐ mol "wanted"	g "wanted"
	g "given"	☐ mol "given"	1 mol "wanted"
	(molar mass of "given")	mol ratio (coefficients from balanced chemical eqn.)	(always 1 mole)

Mole Roadmap



Limiting Reagent (Reactant) – (LR)

- Controls the amt. of product formed
- Completely consumed in the rxn.
- "runs out" first
- Example:
  People on plane → 300 people; 250 seats
  Seats are the limiting factor

Limiting Reagent Problems

- Do mass – mass (g to g) calc. for all reactants.
- Whichever reactant produces the least is the Limiting Reagent (LR)

Excess Reactant (ER)

To find the amount of excess reactant leftover after a rxn:

- Do two mass-mass (g to g) problems to find LR
- Use LR to calculate excess reactant used.
- Subtract excess reactant used from original amt. of excess reactant = leftover excess reactant

Percent Yield

- $$\%\text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \text{ X } 100\%$$

- theoretical yield:
  - maximum amt. of product (what you should "have gotten")
  - from mass-mass (g to g) problem

- actual yield:
  - actual amt. of product (what you "got")
  - from lab result; or given in a problem

Unit 9 - Chapters 10 & 12 – Kinetic Molecular Theory (KMT)

I. Unit Vocabulary

Absolute Zero- no molecular movement at this temperature (0 K, -273^{0}C)
Amorphous- non-crystalline substance such as glass that appears to be solid but is a super cooled liquid
Anhydrous - without water
Barometer- a manometer used to measure atmospheric pressure
Capillary Action - the attraction of the surface of a liquid to the surface of a solid
Condensation- change in state from a gas to a liquid
Deposition – change in state directly from a gas to a solid
Diffusion – mixing of 2 or more gases
Effusion – movement of a gas through a small opening
Evaporation- change in state from a liquid to gas
Fluids – gases and liquids, flow
Ideal gas – imaginary gas that fits all the assumptions of the kinetic molecular theory
Kelvin – SI unit of temperature
Kinetic Theory- group of ideas explaining the interaction of matter and energy due to particle motion
Melting – change in state from a solid to a liquid
Molar heat of fusion – heat needed to melt one mole a substance at its melting pt
Molar heat of vaporization – Heat needed to vaporize one mole of a substance at its boiling pt.
Plasma- high energy state of matter composed of ions that are knocked apart by collisions
Pressure - the number and speed of collisions on a wall of a container
States of Matter- solid, liquid, gas, and plasma are the four states of matter
STP – standard temperature and pressure
Sublimation -change in state directly from a solid to a gas
Surface tension- the apparent skin on surface due to forces holding a liquid together
Triple point – all three major states of matter are in equilibrium at this temperature and pressure
Vapor- gaseous state for substances that are normally a liquid or a solid at room temperature
Viscosity- resistance of liquids to flow
Volatile – a liquid that evaporates readily

II. KMT – Fundamental Concepts

What is an ideal gas?

~an ideal gas is an imaginary gas that perfectly fits all the assumptions of the kinetic molecular theory

What is the kinetic-molecular theory?

~a theory based on the idea that particles of matter are always in motion

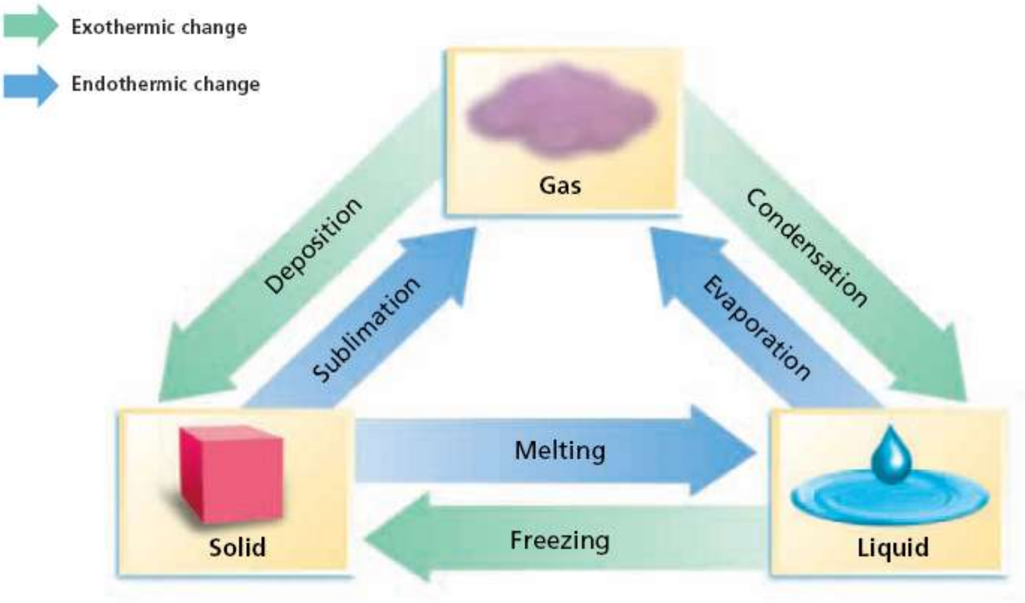
What are the five assumptions of the KMT?

1. **Gases consist of large numbers of tiny particles that are far apart relative to their size.**
2. **Collisions between gas particles and between particles and container walls are elastic collisions.**
3. **Gas Particles are in continuous, rapid, random motion. They therefore possess kinetic energy, which is energy of motion.**
4. **There are no forces of attraction or repulsion between gas particles.**
5. **The average kinetic energy of gas particles depends on the temperature of the gas.**

Explain the properties of expansion, fluidity, low density and compressibility.

1. **expansion - gases completely fill any container in which they are enclosed, and they take its shape.**
2. **fluidity - gas particles glide easily past one another because the attractive forces are insignificant; because liquids and gases flow, they are referred to as fluids.**
3. **low density – gases are about 1/1000 as dense as the same substance in the liquid or solid state.**
4. **compressibility – volume of a given sample of gas can be greatly decreased because the particles which are initially very far apart can be forced closer together (compressed).**

Three Phases of Matter



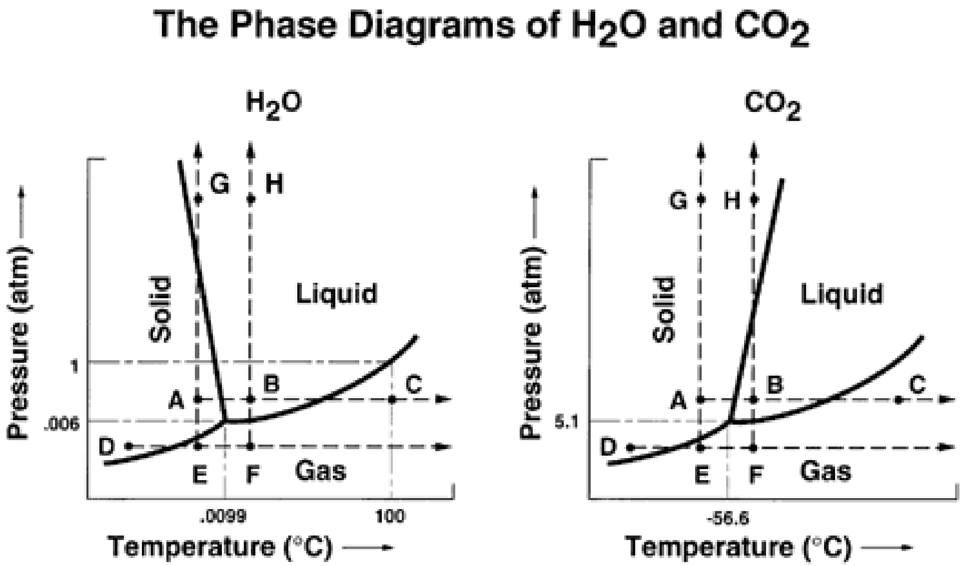
STP: Standard temperature and Pressure; *0^0 C(273K) and 1 atm*

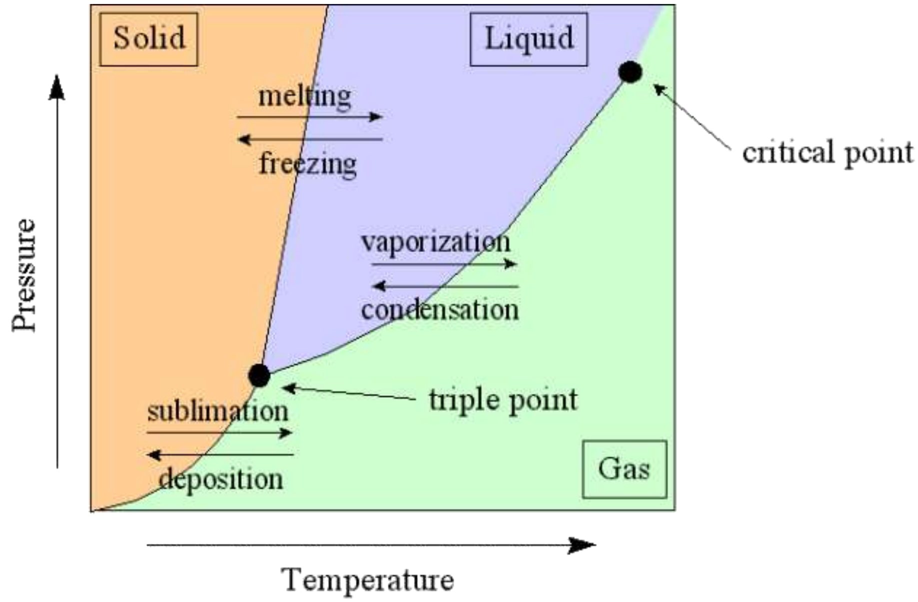
Pressure conversions (all of these values are equal to each other and any two can be set-up as a ratio to be used in a factor label problem! This info is also in the reference tables!!)

1 atm = 760 torr = 760 mm Hg = 101.3 kPa = 1.01 x 10^5 Pa

Molar Heat of Fusion & Molar Heat of Vaporization Problem Examples (covered in Unit 14):

Phase Diagrams for Water & Carbon Dioxide:





phase diagram – a graph of pressure versus temperature that shows the conditions under which the phases of a substance exist

triple point – the temperature and pressure conditions at which the solid, liquid, and vapor of substance can coexist at equilibrium

critical point – indicates the critical temperature and critical pressure of a substance

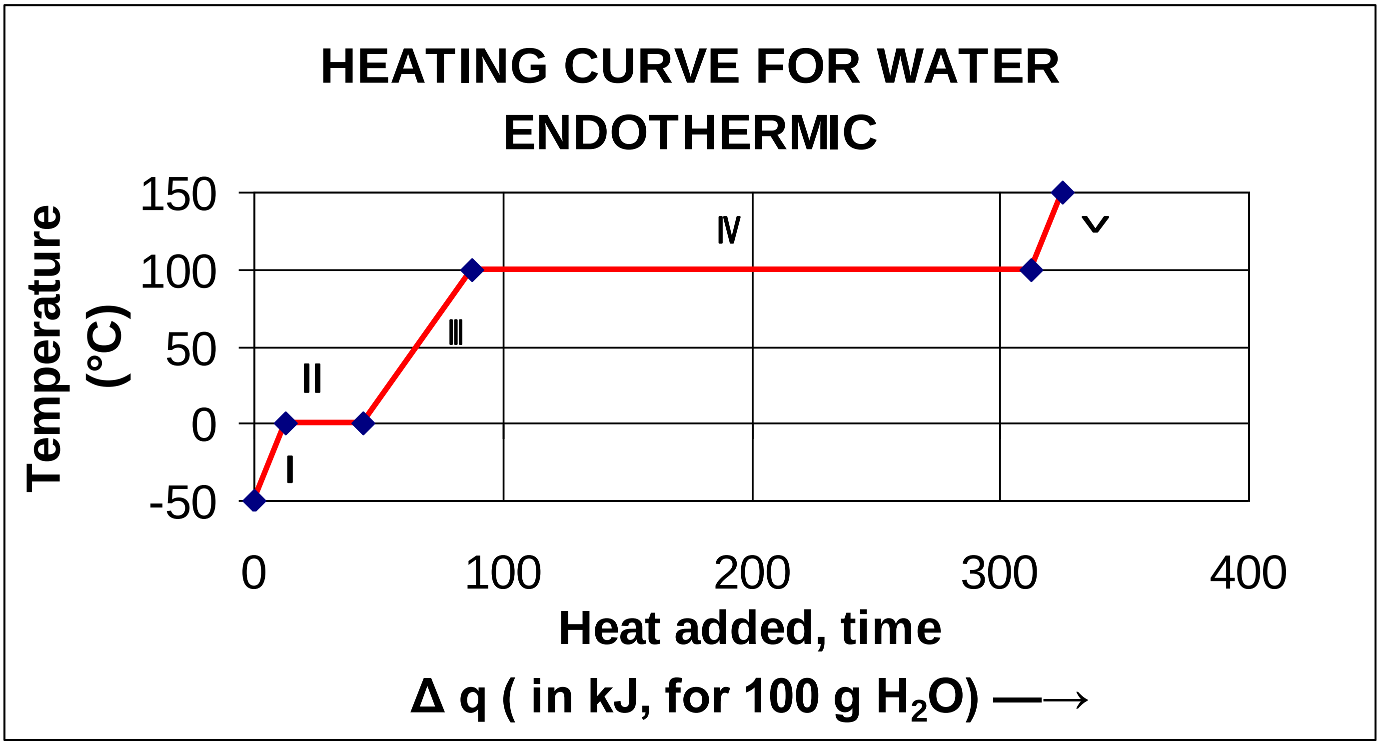
critical temperature – the temperature above which a substance cannot exist in the liquid state

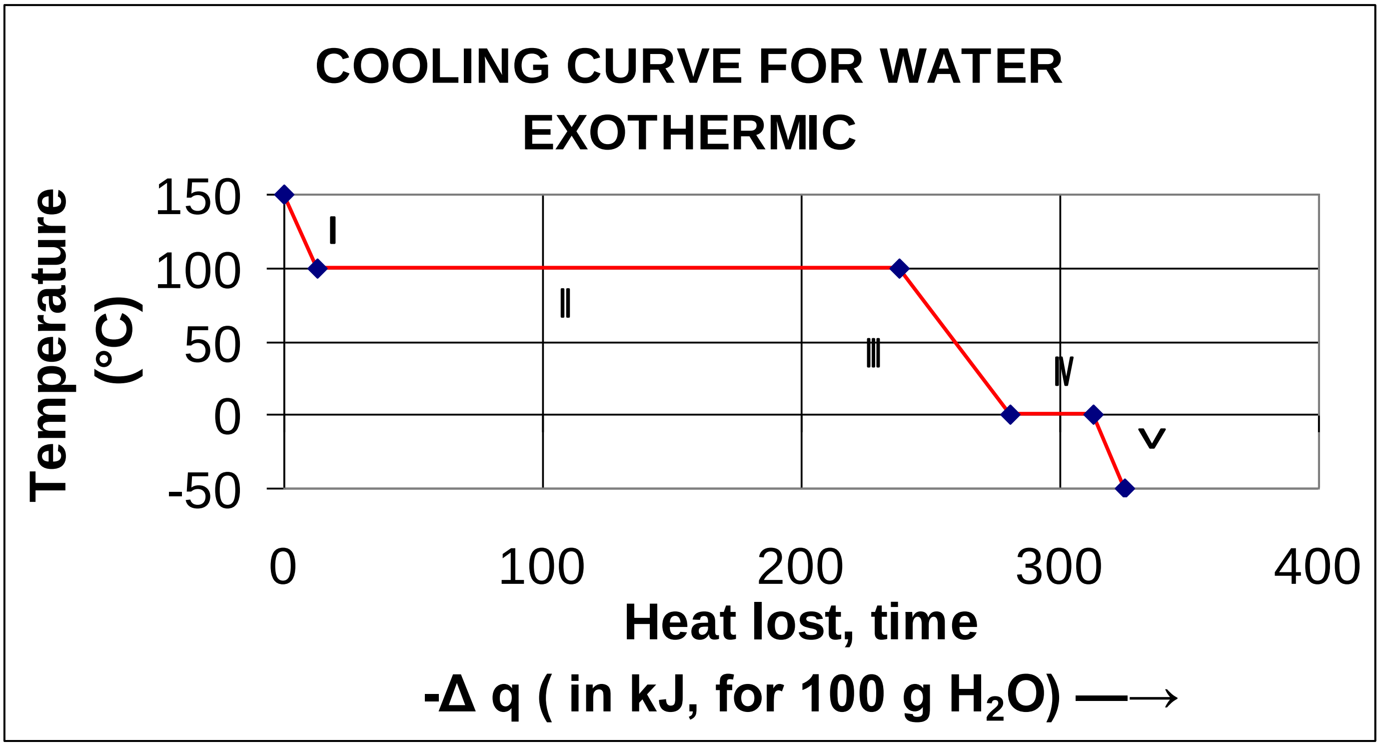
critical pressure – the lowest pressure at which a substance can exist as a liquid at the critical temperature

NOTE: Solid H_2O is less dense than the liquid indicated by the negative slope of the equilibrium line between the solid and liquid phases. The opposite is true for CO_2
(notice the positive slope of this same equilibrium line.

Heating & Cooling Curves

****See Unit 14 for Calculations along each step of the curve****





ENDOTHERMIC CURVE

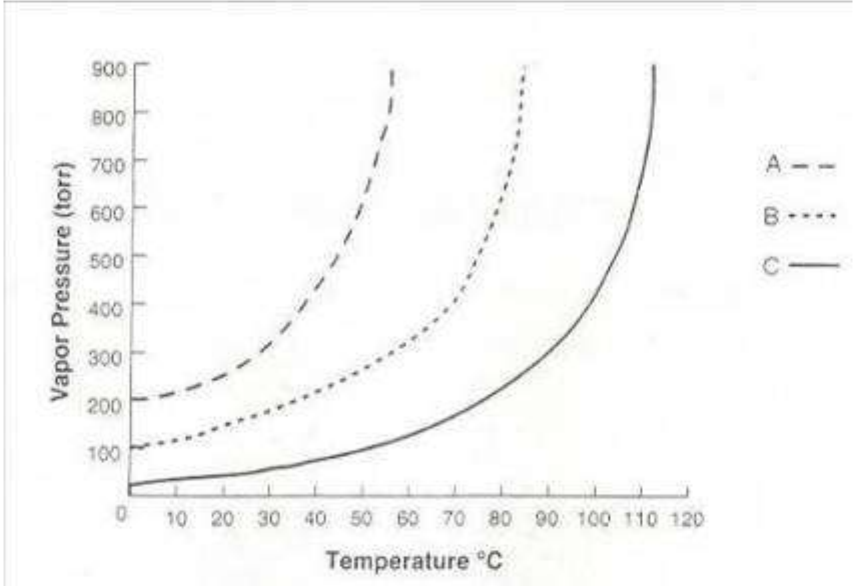
ZONE	I	II	III	IV	V
TEMPERATURE	-50 to 0°C	0°C	0 to 100°C	100°C	100 to 150°C
KINETIC ENERGY	Increasing	Constant	Increasing	Constant	Increasing
PHASE	solid	solid/liquid	liquid	liquid/gas	gas
POTENTIAL ENERGY	constant	increasing	constant	increasing	constant
PHASE CHANGE	no change (solid)	solid to liquid	no change (liquid)	liquid to gas	no change (gas)
DESCRIPTION	heating solid	melting solid	heating liquid	boiling/evaporating liquid	heating gas

EXOTHERMIC CURVE

ZONE	I	II	III	IV	V
TEMPERATURE	150 to 100°C	100°C	100 to 0°C	0°C	0 to -50°C
KINETIC ENERGY	decreasing	constant	decreasing	constant	decreasing
PHASE	gas	gas/liquid	liquid	liquid/solid	solid
POTENTIAL ENERGY	constant	decreasing	constant	decreasing	constant
PHASE CHANGE	no change (gas)	gas to liquid	no change (liquid)	liquid to solid	no change (solid)
DESCRIPTION	cooling gas	condensing the gas	cooling liquid	freezing the liquid	cooling solid

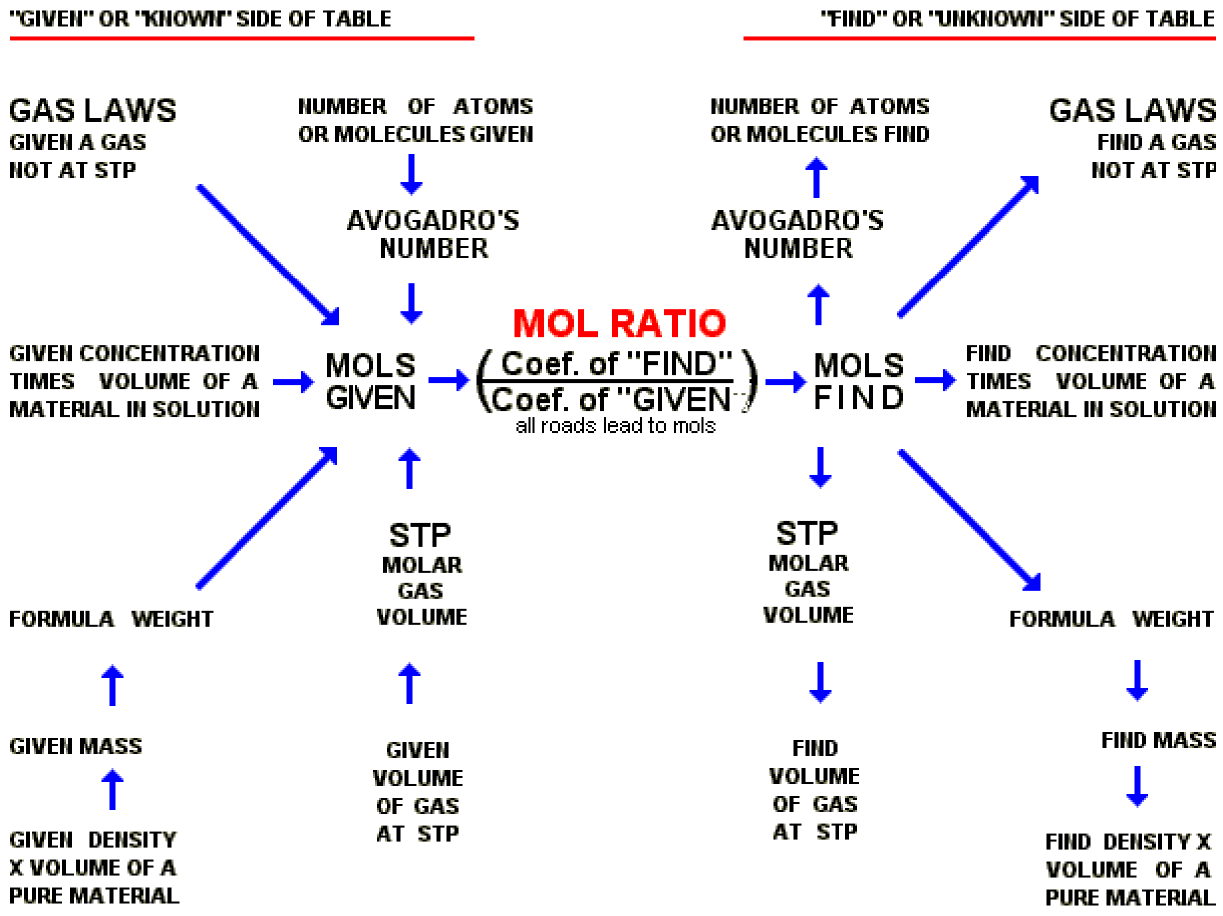
Vapor Pressure Curves:

A liquid will boil when its vapor pressure equals atmospheric pressure.



Unit 10 - Chapters 10 & 11 – Gases, Gas Laws, and Gas Stoichiometry

Stoichiometry Roadmap including gases.......

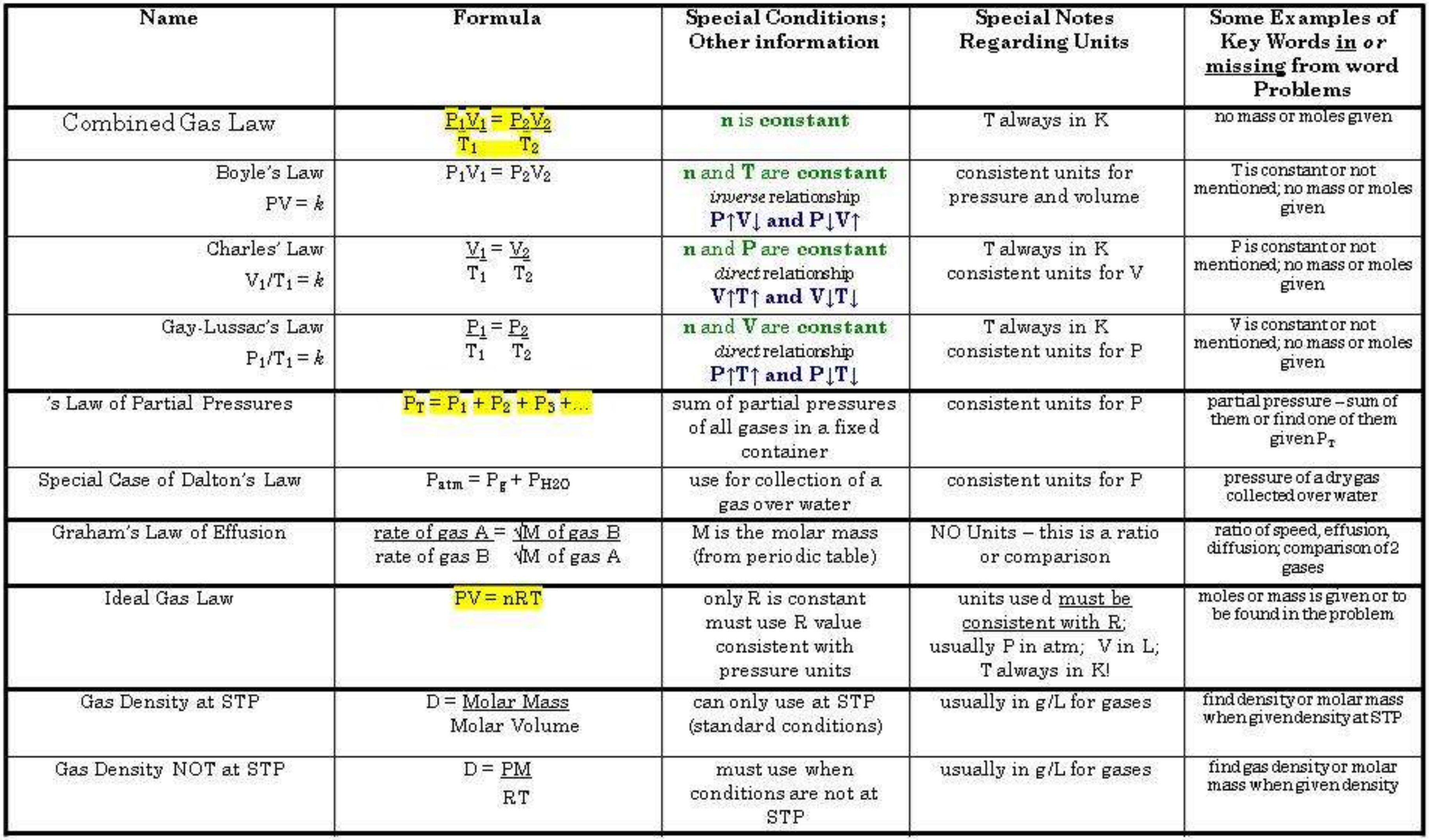


Gas Laws Summary Table

Definitions:

P = pressure
V = volume
T = temperature (in Kelvin always for gas laws!)
STP: standard temperature = 0^0C or 273 K
standard pressure = 1 atm = 760 mm Hg = 760 torr = 101.3 kPa

n = # of moles
R = gas constant (0.0821 L•atm/mol•K); this is the most common one
molar volume: 1 mol/22.4 L or 22.4L/1 mol of any gas at STP
K = ^{0}C + 273
M = Molar Mass (on EOC packet *M* is molarity – it is italicized!!)

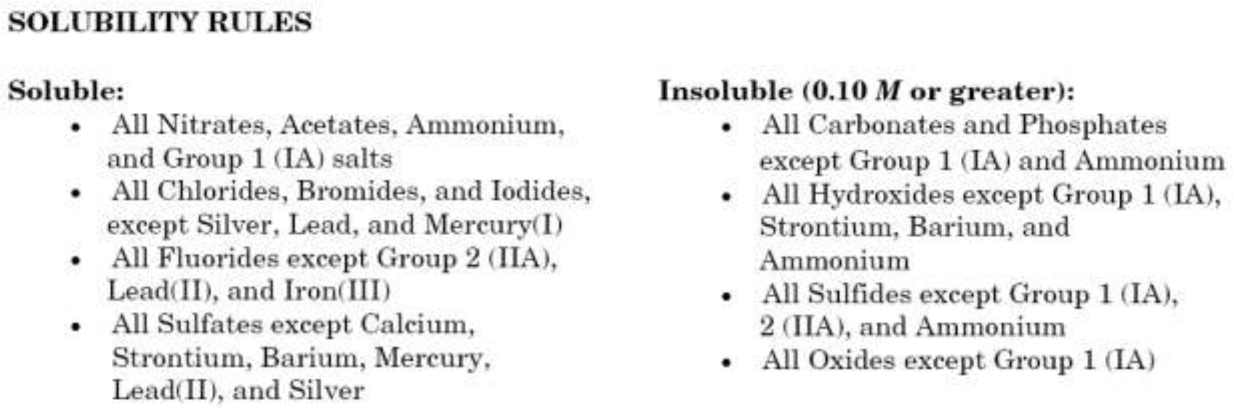


Name	Formula	Special Conditions; Other information	Special Notes Regarding Units	Some Examples of Key Words <u>in</u> *or* <u>missing</u> from word Problems
Combined Gas Law	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	**n is constant**	T always in K	no mass or moles given
Boyle's Law $PV = k$	$P_1V_1 = P_2V_2$	**n and T are constant** *inverse* relationship **P↑V↓ and P↓V↑**	consistent units for pressure and volume	T is constant or not mentioned; no mass or moles given
Charles' Law $V_1/T_1 = k$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	**n and P are constant** *direct* relationship **V↑T↑ and V↓T↓**	T always in K consistent units for V	P is constant or not mentioned; no mass or moles given
Gay-Lussac's Law $P_1/T_1 = k$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	**n and V are constant** *direct* relationship **P↑T↑ and P↓T↓**	T always in K consistent units for P	V is constant or not mentioned; no mass or moles given
's Law of Partial Pressures	$P_T = P_1 + P_2 + P_3 + \ldots$	sum of partial pressures of all gases in a fixed container	consistent units for P	partial pressure – sum of them or find one of them given P_T
Special Case of Dalton's Law	$P_{atm} = P_g + P_{H2O}$	use for collection of a gas over water	consistent units for P	pressure of a dry gas collected over water
Graham's Law of Effusion	$\frac{\text{rate of gas A}}{\text{rate of gas B}} = \frac{\sqrt{\text{M of gas B}}}{\sqrt{\text{M of gas A}}}$	M is the molar mass (from periodic table)	NO Units – this is a ratio or comparison	ratio of speed, effusion, diffusion; comparison of 2 gases
Ideal Gas Law	$PV = nRT$	only R is constant must use R value consistent with pressure units	units used <u>must be consistent with R</u>; usually P in atm; V in L; T always in K!	moles or mass is given or to be found in the problem
Gas Density at STP	$D = \frac{\text{Molar Mass}}{\text{Molar Volume}}$	can only use at STP (standard conditions)	usually in g/L for gases	find density or molar mass when given density at STP
Gas Density NOT at STP	$D = \frac{PM}{RT}$	must use when conditions are not at STP	usually in g/L for gases	find gas density or molar mass when given density

Highlighted Info Given in NCDPI EOC Reference Tables

Unit 11 - Chapters 13 & 14 – Solutions and Colligative Properties

From the Chemistry Reference Tables Packet:



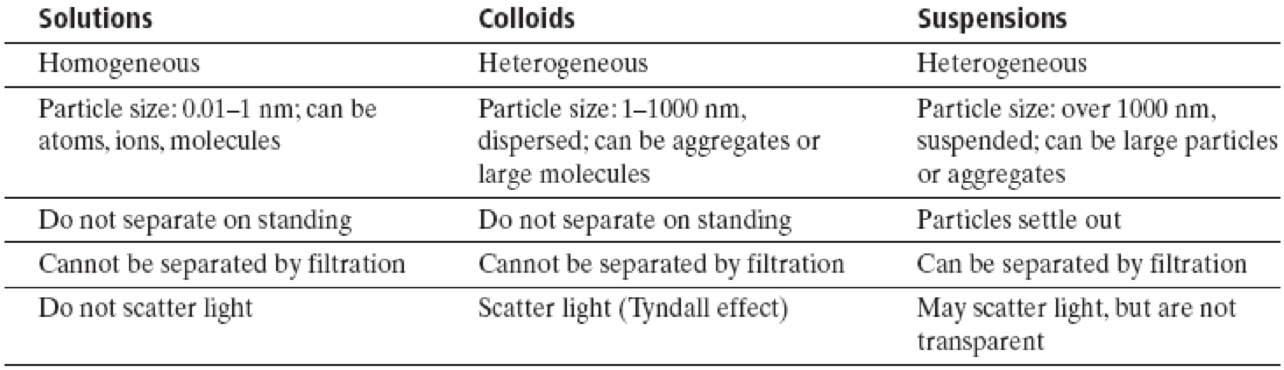
SOLUBILITY RULES

Soluble:

- All Nitrates, Acetates, Ammonium, and Group 1 (IA) salts
- All Chlorides, Bromides, and Iodides, except Silver, Lead, and Mercury(I)
- All Fluorides except Group 2 (IIA), Lead(II), and Iron(III)
- All Sulfates except Calcium, Strontium, Barium, Mercury, Lead(II), and Silver

Insoluble (0.10 *M* or greater):

- All Carbonates and Phosphates except Group 1 (IA) and Ammonium
- All Hydroxides except Group 1 (IA), Strontium, Barium, and Ammonium
- All Sulfides except Group 1 (IA), 2 (IIA), and Ammonium
- All Oxides except Group 1 (IA)



Solutions	Colloids	Suspensions
Homogeneous	Heterogeneous	Heterogeneous
Particle size: 0.01–1 nm; can be atoms, ions, molecules	Particle size: 1–1000 nm, dispersed; can be aggregates or large molecules	Particle size: over 1000 nm, suspended; can be large particles or aggregates
Do not separate on standing	Do not separate on standing	Particles settle out
Cannot be separated by filtration	Cannot be separated by filtration	Can be separated by filtration
Do not scatter light	Scatter light (Tyndall effect)	May scatter light, but are not transparent

Solute = what is being dissolved
Solvent = dissolving medium (often water)
Solutions = solute and solvent together

Dilute = ?
Concentrated = ?

Molarity - M A way to measure solution concentration. It's the most common chemistry concentration unit.

$$\mathbf{M} = \frac{\text{moles solute}}{\text{liters of solution}}$$

Making a Molar Solution


1

Start by calculating the mass of $CuSO_4 \bullet 5H_2O$ needed. Making a liter of this solution requires 0.5000 mol of solute. Convert the moles to mass by multiplying by the molar mass of $CuSO_4 \bullet 5H_2O$. This mass is calculated to be 124.8 g.


2

Add some solvent to the solute to dissolve it, then pour it into a 1.0 L volumetric flask. Rinse the weighing beaker with more solvent to remove all the solute, and pour the rinse into the flask. Add water until the volume of the solution nears the neck of the flask.


3

Put the stopper in the flask, and swirl the solution thoroughly.


4

Carefully fill the flask to the 1.0 L mark with water. Restopper the flask and invert it at least 10 times to ensure complete mixing.


5

The resulting solution has 0.5000 mol of solute dissolved in 1.000 L of solution, which is a 0.5000 M concentration.

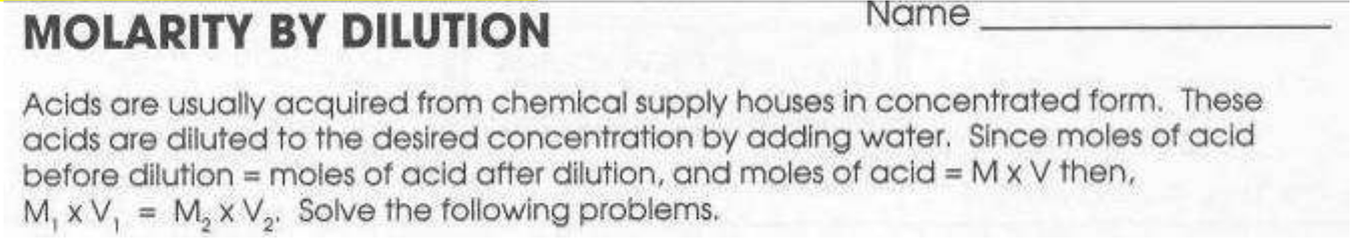
Molarity Calculations (also in Unit 5):

- **Molarity** is the term used for moles dissolved in solution
- Symbol for **Molarity = M**
- Definition – moles of solute per liter of solution
- Formula

$$M = \frac{\text{moles solute (mol)}}{\text{liter solution (L)}}$$

*****Be able to calculate Molarity or grams of solute needed.*****

Molarity By Dilution Calculations:



MOLARITY BY DILUTION Name ____________

Acids are usually acquired from chemical supply houses in concentrated form. These acids are diluted to the desired concentration by adding water. Since moles of acid before dilution = moles of acid after dilution, and moles of acid = M x V then, $M_1 \times V_1 = M_2 \times V_2$. Solve the following problems.

Molality = *m* - another way to measure solution concentration; it is independent of temperature *(volume can change with temperature, so M can be affected by changes in temperature)*

$$m = \frac{\text{moles of solute}}{\text{Kg of solvent}}$$

Making a *Molal* Solution

1

Calculate the mass of $CuSO_4 \cdot 5H_2O$ needed. To make this solution, each kilogram of solvent (1000 g) will require 0.5000 mol of $CuSO_4 \cdot 5H_2O$. This mass is calculated to be 124.8 g.

2

Add exactly 1 kg of solvent to the solute in the beaker. Because the solvent is water, 1 kg will equal 1000 mL.

3

Mix thoroughly.

4

The resulting solution has 0.5000 mol of solute dissolved in 1 kg of solvent.

Net Ionic Equations:

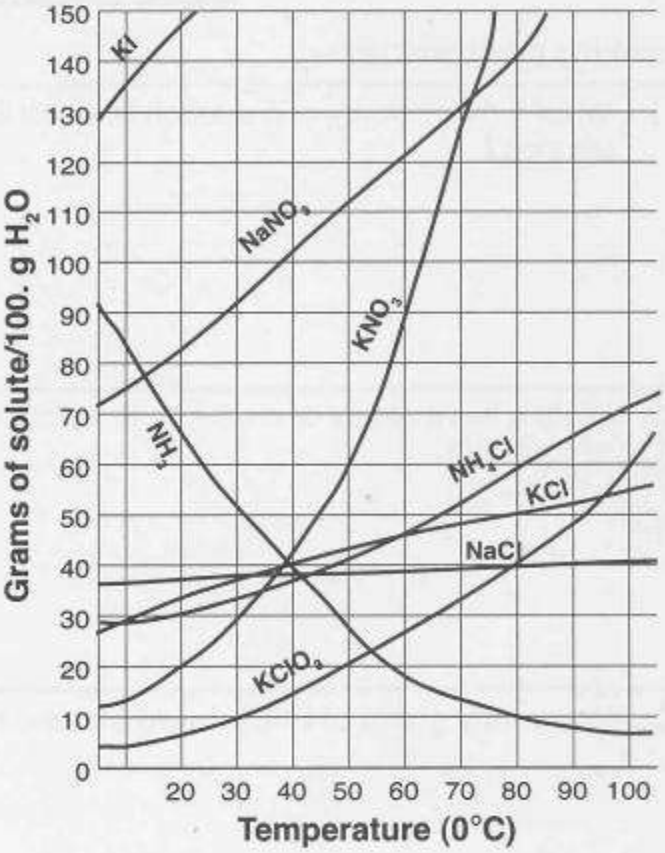
Steps for writing a net ionic equation
1. Write a **balanced** equation.
2. Use your solubility rules to determine solubility and break soluble compounds into ions (with charges!). Leave insoluble compounds together without any charges!
3. Cancel out spectator ions – if everything cancels write no net reaction.
4. Write the final equation – ***include state symbols!***

    1. sodium hydroxide + zinc (II) nitrate → sodium nitrate + zinc (II) hydroxide

SOLUBILITY CURVES

Name ____________________

Answer the following questions based on the solubility curve below.

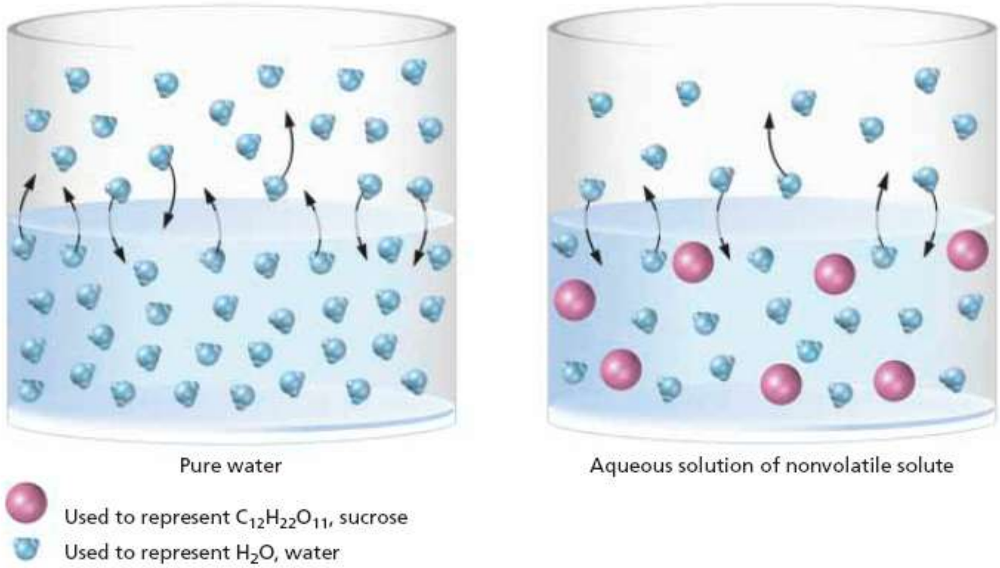


1. Which salt is least soluble in water at 20° C? ____________________
2. How many grams of potassium chloride can be dissolved in 200 g of water at 80° C? ____________________
3. At 40° C, how much potassium nitrate can be dissolved in 300 g of water? ____________________
4. Which salt shows the least change in solubility from 0° – 100° C? ____________________
5. At 30° C, 90 g of sodium nitrate is dissolved in 100 g of water. Is this solution saturated, unsaturated or supersaturated? ____________________
6. A saturated solution of potassium chlorate is formed from one hundred grams of water. If the saturated solution is cooled from 80° C to 50° C, how many grams of precipitate are formed? ____________________
7. What compound shows a decrease in solubility from 0° to 100° C? ____________________
8. Which salt is most soluble at 10° C? ____________________
9. Which salt is least soluble at 50° C? ____________________
10. Which salt is least soluble at 90° C? ____________________

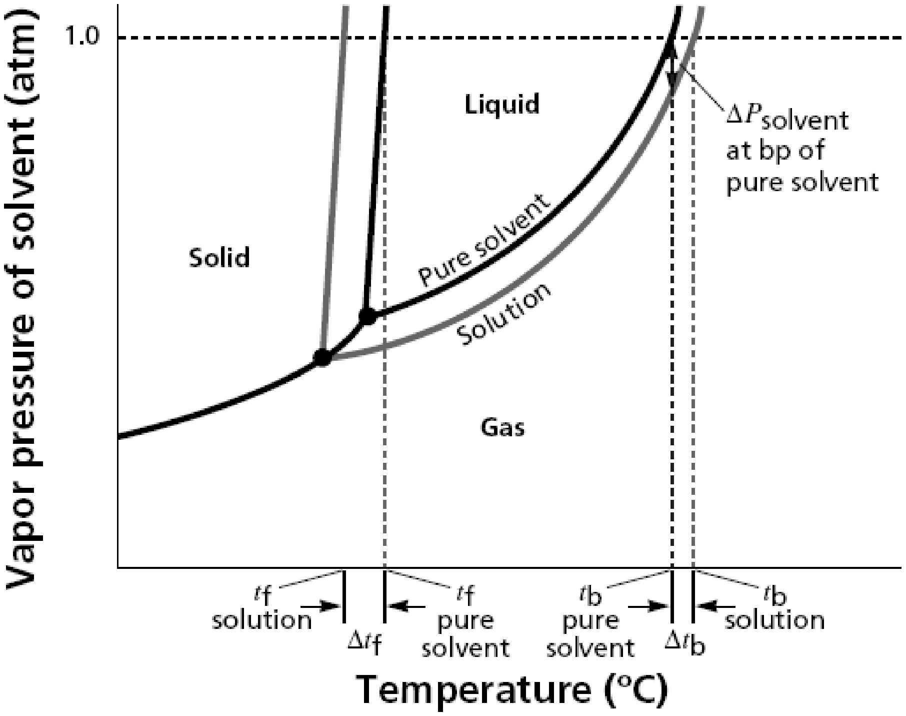
The 3 Colligative properties:

A. Vapor pressure lowering=

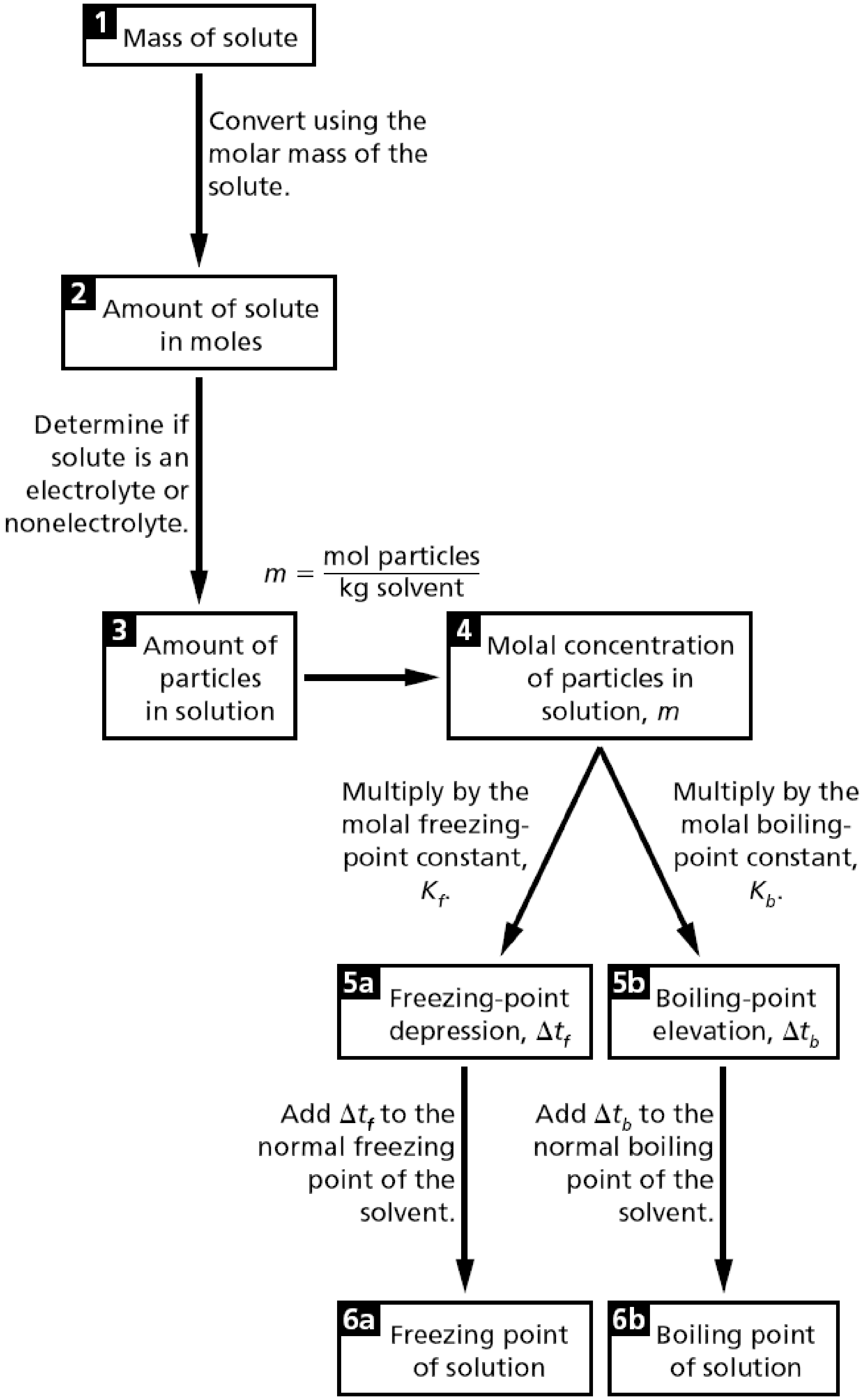
The vapor pressure of water over pure water is greater than the vapor pressure of water over an aqueous solution containing a nonvolatile solute.



B. Freezing point depression and Boiling Pt. Elevation:

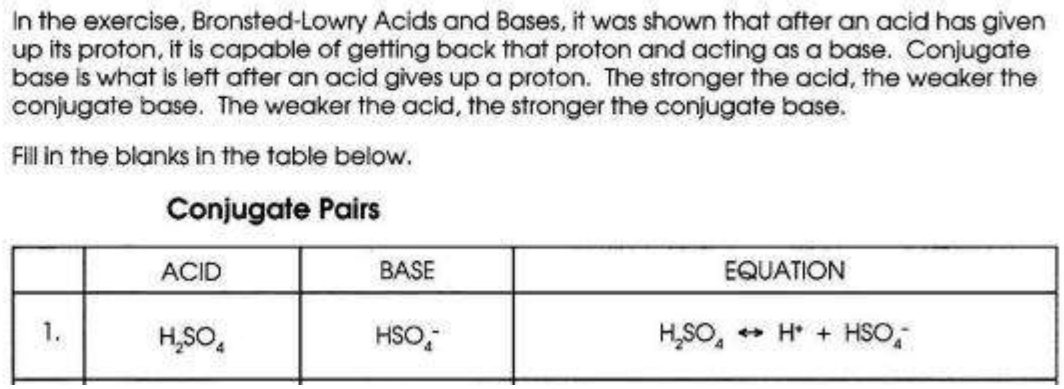


General Plan for Solving Problems Involving Freezing-Point Depression and Boiling-Point Elevation



Unit 12 - Chapter 15 – Acids & Bases

Conjugate Acid-Base Pairs



In the exercise, Bronsted-Lowry Acids and Bases, it was shown that after an acid has given up its proton, it is capable of getting back that proton and acting as a base. Conjugate base is what is left after an acid gives up a proton. The stronger the acid, the weaker the conjugate base. The weaker the acid, the stronger the conjugate base.

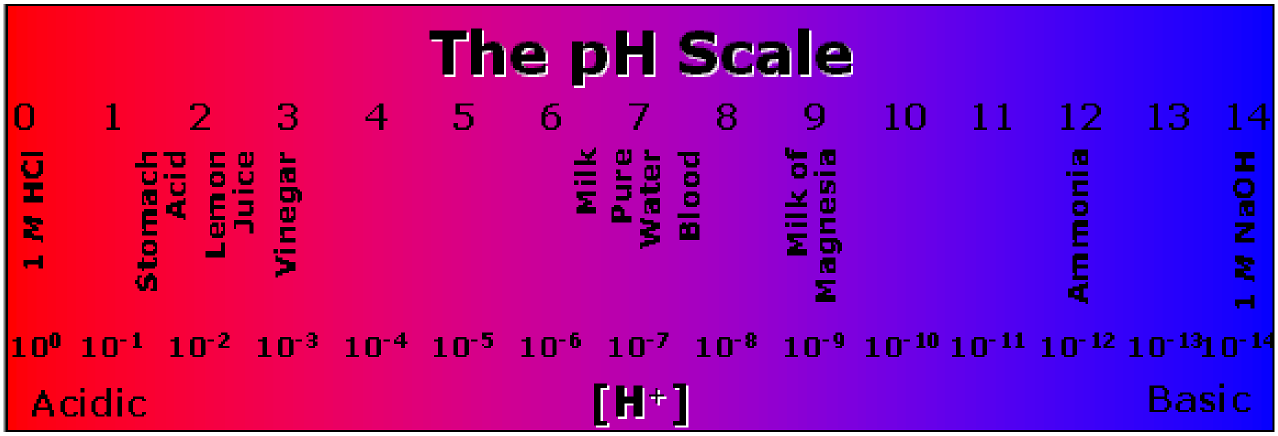
Fill in the blanks in the table below.

Conjugate Pairs

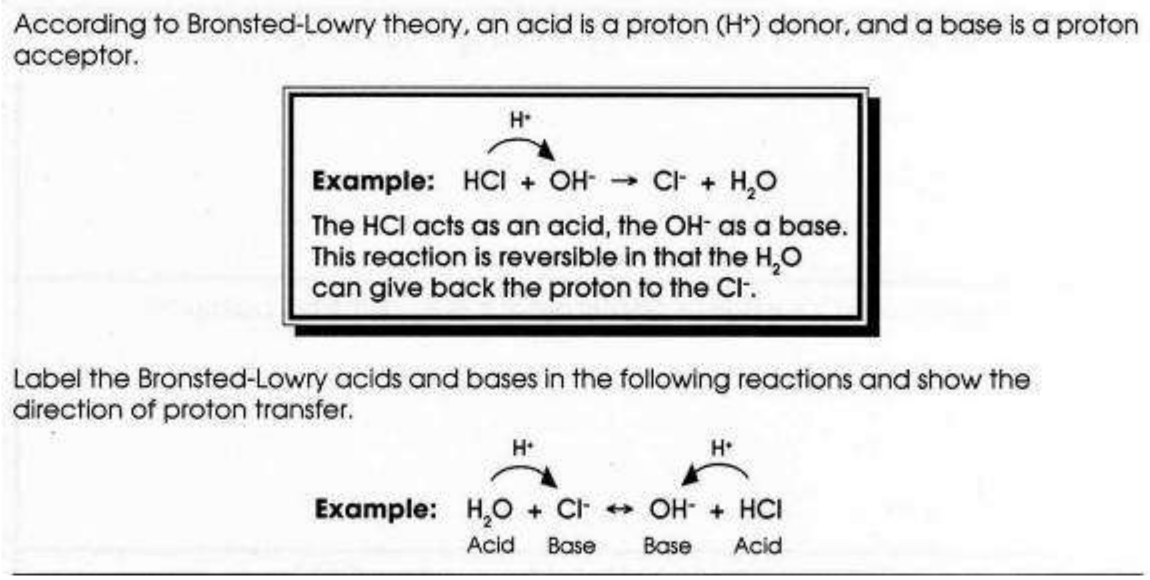
	ACID	BASE	EQUATION
1.	H_2SO_4	HSO_4^-	$H_2SO_4 \leftrightarrow H^+ + HSO_4^-$

COMMON ACIDS TO BE MEMORIZED!

Hydrochloric Acid	HCl
Nitric Acid	HNO_3
Acetic Acid	$HC_2H_3O_2$ or CH_3COOH
Sulfuric Acid	H_2SO_4



Bronsted-Lowry Acids & Bases



According to Bronsted-Lowry theory, an acid is a proton (H^+) donor, and a base is a proton acceptor.

$$\overset{H^+}{\curvearrowright}$$

Example: $HCl + OH^- \rightarrow Cl^- + H_2O$

The HCl acts as an acid, the OH^- as a base. This reaction is reversible in that the H_2O can give back the proton to the Cl^-.

Label the Bronsted-Lowry acids and bases in the following reactions and show the direction of proton transfer.

$$\overset{H^+}{\curvearrowright} \qquad \overset{H^+}{\curvearrowleft}$$

Example: $\underset{\text{Acid}}{H_2O} + \underset{\text{Base}}{Cl^-} \leftrightarrow \underset{\text{Base}}{OH^-} + \underset{\text{Acid}}{HCl}$

Unit 13 - Chapter 16 – pH & Titrations

Be able to Use These Six Equations – In Reference Tables:

$K_w = [OH^-][H^+] = 1 \text{ X } 10^{-14}$ $pH + pOH = 14$

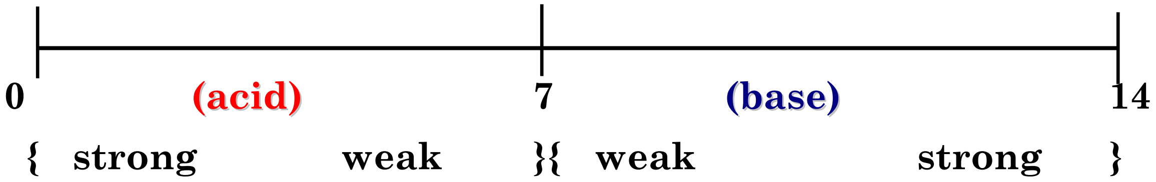
$pH = -\log[H^+]$ $[H^+] = 10^{-pH}$

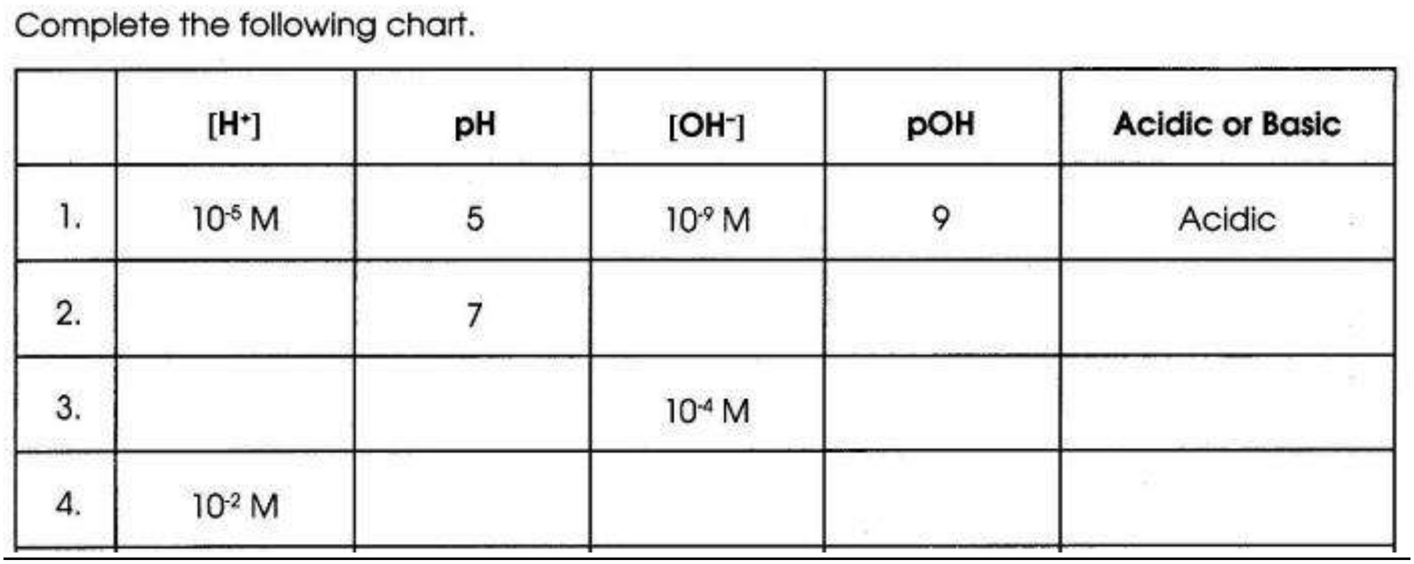
$pOH = -\log[OH^-]$ $[OH^-] = 10^{-pOH}$

- Pure H_2O has:

  1 X 10^{-7} M H^+ (same as H_3O^+) and 1 X 10^{-7} M OH^{-1}

- Because $[H^+]=[OH^-]$ in water = neutral solution
  (remember the brackets stand for concentration in molarity)
- When $[H^+] > [OH^{-1}]$; the solution is **acidic**
- When $[OH^{-1}] > [H^+]$; the solution is **basic**
- pH is an easier way to express concentration; remember:

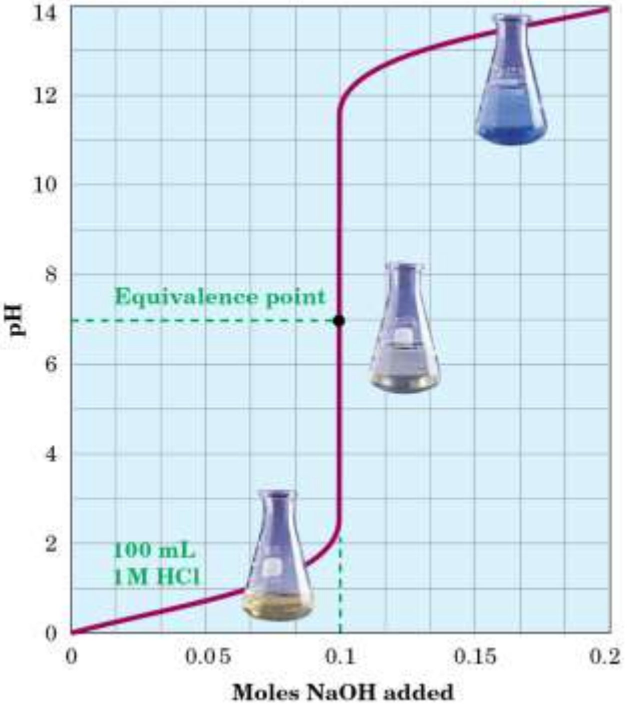




Complete the following chart.

	$[H^+]$	pH	$[OH^-]$	pOH	Acidic or Basic
1.	10^{-5} M	5	10^{-9} M	9	Acidic
2.		7			
3.			10^{-4} M		
4.	10^{-2} M				

Titration Curve for the Titration of a Strong Acid with a Strong Base



Write the Balanced Equation		
Acid Data	Base Data	
M (mol/L) = V (mL) =	M (mol/L)= V (mL) =	
(start with the side that has the MOST data)		
	Mole Ratio	
Mole Conversion from “given” to “unknown”		
Calculate Molarity or Volume of Unknown:		

Using Titration Data to Determine Molarity

Use the 4 following steps:

1) Use balanced equation (you may have to write it yourself!) for the neutralization reaction to determine the ratio of moles of acid to base.
2) Determine the moles of standard solution (acid or base) used during the titration.
3) Determine the moles of solute of unknown solution used during the titration.
4) Determine the molarity (or the volume) of the unknown solution.

*****Be able to use data from buret readings to find the volume of titrant used in titration calculations!*****

Unit 14 Chapter 17 – Thermochemistry, Reaction Rates, Entropy, Enthalpy

Heat/Energy –

Remember heat is one form of energy and the terms are often used interchangeably.

- Heat (**q or Q**) is the energy (**E**) transferred due to the difference in temp. (**T**).
  Unit = J (joule) or cal (calorie)
- 1 calorie (cal) = 4.184 J
- Food calories are actually kilocalories → 1 food calorie = 1000 "chemical" calories

Example:
How many calories are in a potato with 686000 J of energy? (Copy Work!)

Specific Heat (***c or*** c_p): the amount of heat energy required to raise the temperature of one gram of a substance by 1 ^{0}C (one degree Celsius) or 1 K (one kelvin) – *because the sizes of the degree divisions on both scales are equal.*

***Specific heat is usually measured under constant pressure conditions – the subscript p* (C_p) *– is used as a reminder.*

Specific heat is a constant for a substance. You must notice the state of matter (s,l,g) when selecting the correct constant value from the **specific heat from the reference packet**.

For water *(l)*, c = 4.18 J/g^{0}C - given in packet! Also given for ice and steam!

$$Q = mC_p\Delta T$$ *(EOC packet)*

Use when there is a change in Temperature

where
- C= specific heat
- q = heat gained or lost in J
- m = mass in g
- ΔT = change in temperature; IMPORTANT: $\Delta T = T_f - T_i$, in ^{0}C or K

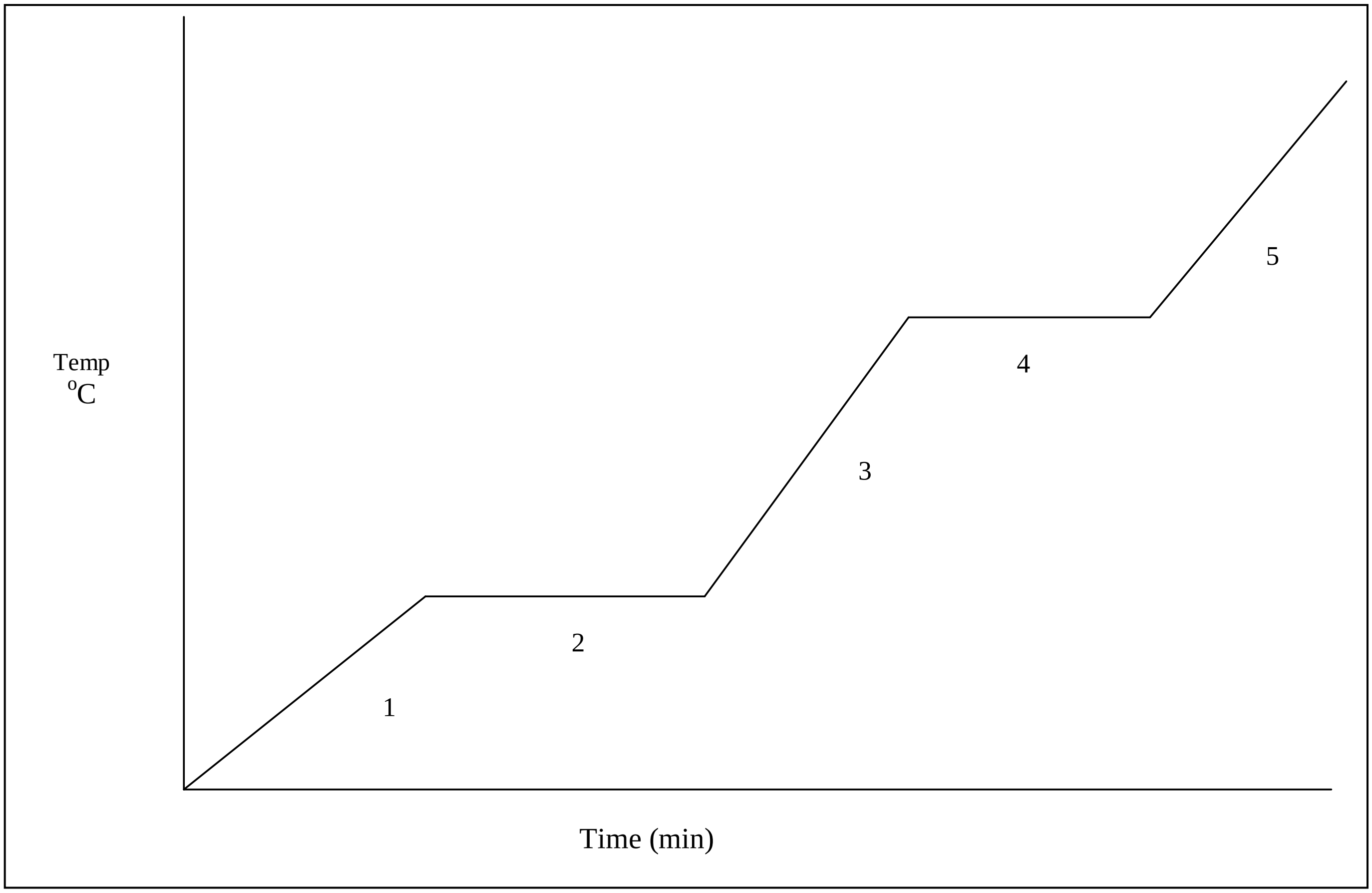
Equation can be rearranged to $C_p = \dfrac{q}{m\,\Delta T}$

To calculate the heat required for changes of state (**at constant temperature**), use

$q = mH_v$ for boiling/condensing and $q = mH_f$ for freezing/melting

for water: H_v = 2,260 J/g and H_f = 334J/g (all of this is in the packet!)

Heating/Cooling Curves



Legend for Calculations:

1. **$Q = m\Delta TC_{p(solid)}$**
2. **$Q = mH_f$**
3. **$Q = m\Delta TC_{p(liquid)}$**
4. **$Q = mH_v$**
5. **$Q = m\Delta TC_{p(gas)}$**

Using your Reference Tables, list the values for the following variables:

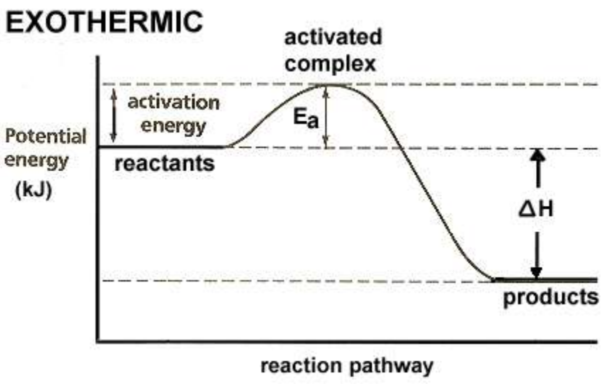
$C_{p(solid)}$
$C_{p(liquid)}$
$C_{p(gas)}$
H_f

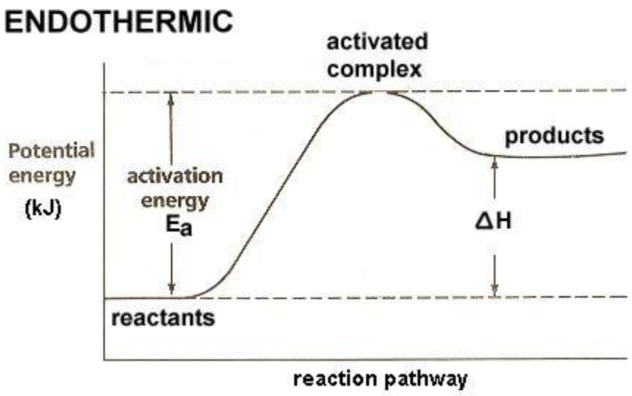
Phase Diagrams – See Unit 9

Reaction rate depends on five things:

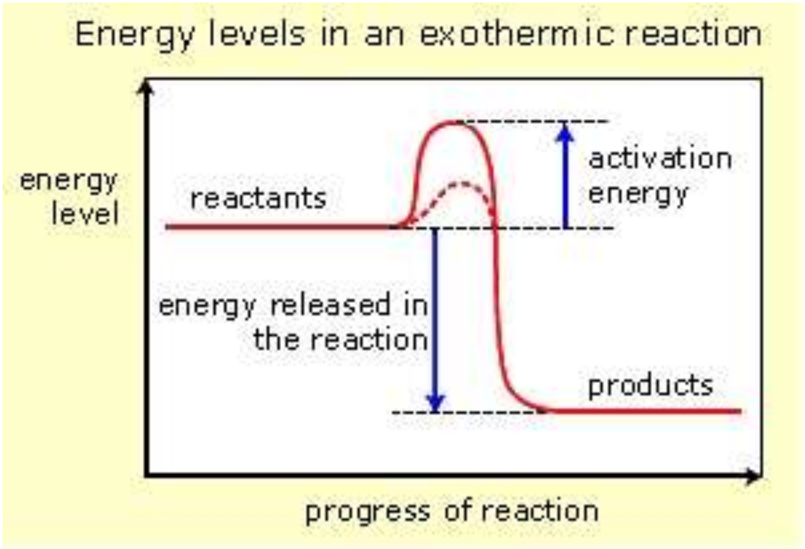
1. nature of reactants
2. surface area
3. temperature
4. concentration
5. presence of a catalyst

Energy Diagrams





http://www.saskschools.ca/curr_content/chem30_05/graphics/2_graphics/exo.gif
http://www.saskschools.ca/curr_content/chem30_05/graphics/2_graphics/endo.gif



Red dashed line demonstrates the reduction of the activation energy through the use of a catalyst.
http://www.bbc.co.uk/schools/gcsebitesize/chemistry/chemicalreactions/2energychangesrev3.shtml

Unit 15 Chapter 18 – Equilibrium, LeChatelier's Principle, K

Chemical Equilibrium

Reversible reaction :

Use ←→

Chemical equilibrium: state of balance in which the rates of opposing rxns are exactly equal

- Dynamic state: reactions are continually happening

- **Example:** students changing rooms

Equilibrium constant: K or K_{eq} or K_c

$$K = \frac{[\text{products}]^{\text{coefficients}}}{[\text{reactants}]^{\text{coefficients}}}$$

Why use K?

K = 1	reactions are equal → at *equilibrium*!!
K< 1	reverse reaction favored, more reactants
K> 1	forward reaction favored, more products

Very Important:

- *K includes only gases and aqueous solutions; liquids and solids do not have a concentration*

Acids, Bases, and Salts

- ***Weak acid:*** ionizes (breaks down) partially
  - K_a = *acid ionization or dissociation constant*
- ***Weak base:*** slightly dissociate just like a weak acid
  - K_b = *hydrolysis constant*

K_w = *dissociation constant for water* = $[OH^{-1}][H^{+}]$

Buffer: solution that can resist changes in pH, usually made up of a weak acid or base and a salt of the weak acid or base.

Le Chatelier's Principle

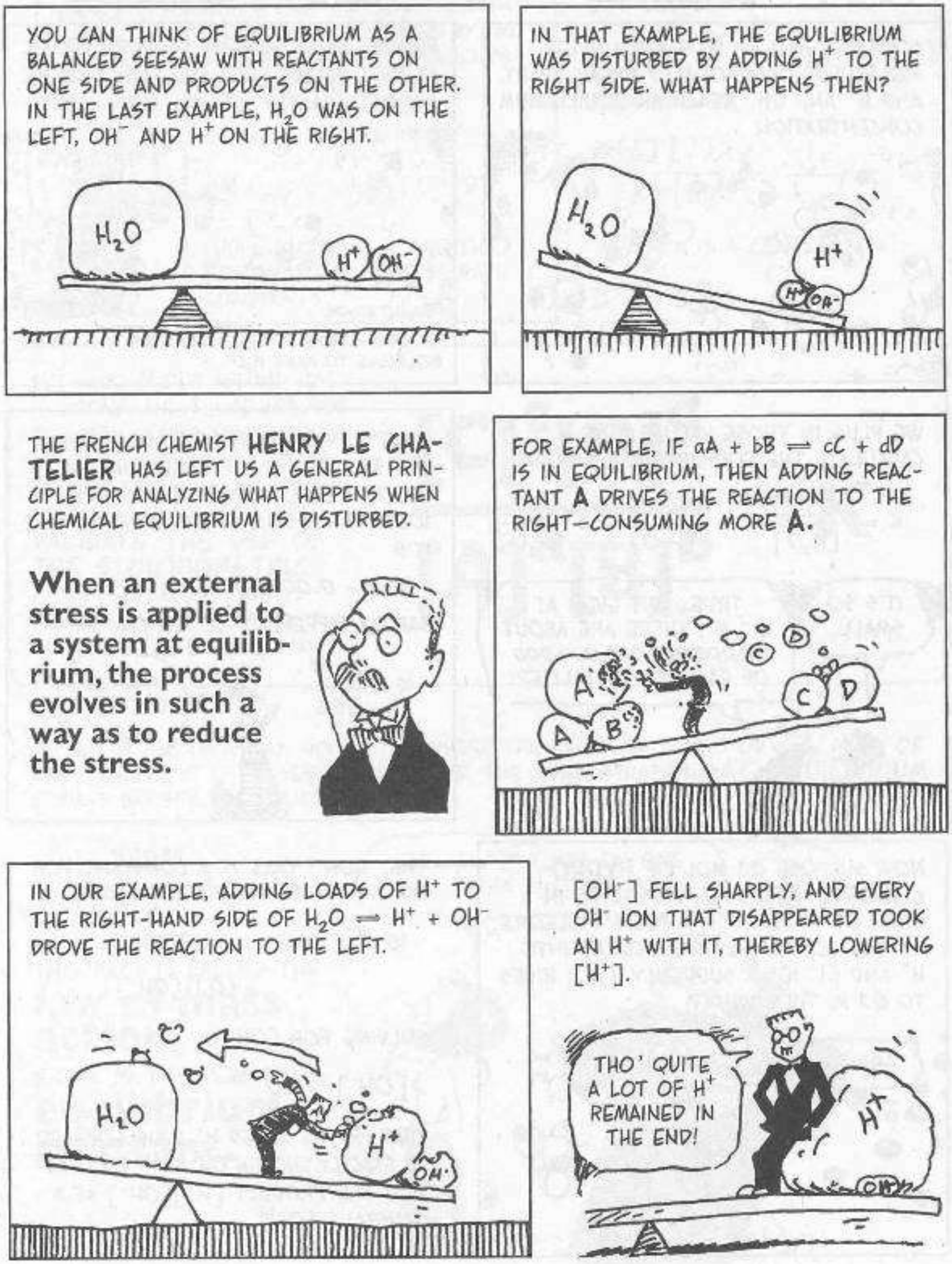
A system at equilibrium will shift to adjust to changes to stay at equilibrium.

Conditions:

1. ***Concentration:*** A. Increase/add → shifts away (to consume excess)
   B. Decrease/Remove → shifts towards (to replace)
2. ***Pressure***: A. Increase → shifts to side with less gaseous moles
   B. Decrease → shifts to side with more gaseous moles
3. ***Temperature:*** – depends on if rxn is exothermic or endothermic

Temperature is the Only condition that changes the value of K

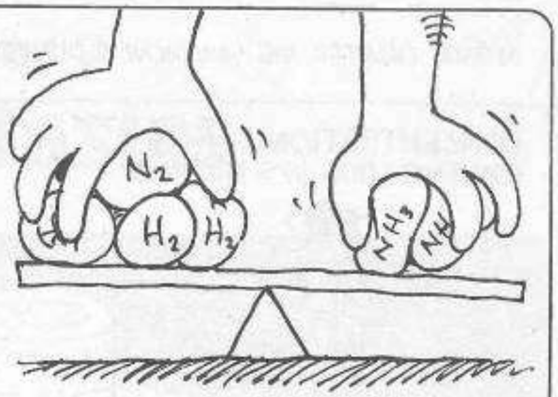
Le Chatelier's Principle



LE CHATELIER VERY CLEVERLY APPLIED HIS OWN PRINCIPLE TO THE SYNTHESIS OF **AMMONIA,** NH_3, A KEY INGREDIENT OF COUNTLESS PRODUCTS, FROM FERTILIZER TO EXPLOSIVES.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

INCREASING **PRESSURE,** SAID HIS PRINCIPLE, WILL DRIVE THE REACTION IN THE DIRECTION THAT **REDUCES PRESSURE.**



THERE ARE FOUR MOLES OF GAS ON THE LEFT, BUT ONLY TWO ON THE RIGHT. BY THE GAS LAW, PRESSURE IS DIRECTLY PROPORTIONAL TO THE NUMBER OF MOLES. SO PRESSURE IS RELIEVED WHEN THE REACTION GOES IN THE DIRECTION OF **FEWER MOLES,** THAT IS, TO THE RIGHT.

IN 1901, LE CHATELIER ATTEMPTED THE SYNTHESIS AT A PRESSURE OF *200* atm IN A STEEL "BOMB" HEATED TO *600°C*. UNFORTUNATELY, AN AIR LEAK CAUSED THE BOMB TO EXPLODE...

...AND THE CHEMIST GAVE UP THIS FERTILE LINE OF INVESTIGATION.

FIVE YEARS LATER, THE GERMAN **FRITZ HABER** SUCCEEDED WHERE LE CHATELIER HAD FAILED, AND EVER SINCE, AMMONIA SYNTHESIS HAS BEEN KNOWN AS THE

Haber process.

"I LET THE DISCOVERY OF THE AMMONIA SYNTHESIS SLIP THROUGH MY HANDS. IT WAS THE GREATEST BLUNDER OF MY SCIENTIFIC CAREER."
—LE CHATELIER

K_{sp} – Solubility Constant Expression

Solubility: the amount of substance required to form a saturated solution with a specific amount of solvent at a specified temperature.

- Units = ? ***Units – g/L or M***
- **Remember**, salts ionize (break down) in water.
- When a solution is saturated, it is said to be at equilibrium →

A saturated solution of salt and water is at equilibrium.

Unit 16 Chapters 19 & 22 – Redox Rxns, Electrochemistry & Nuclear Chemistry

Electrochemistry:

Students should be able to:

- Determine oxidation number of each element in a REDOX reaction, including peroxides.
- Determine elements oxidized and reduced.
- Write half reactions indicating gain or loss of electrons and identify the reaction as either reduction or oxidation.

Students should be aware of some practical applications of oxidation/reduction reactions. Some examples include: simple wet cell, dry cell, bleaching, and electroplating.

Oxidation Numbers:

<u>Summary of Rules for Oxidation Numbers:</u>

- **Rule 1:** Atoms in a pure element have an oxidation number of zero.
- **Rule 2:** The more electronegative element in a binary compound is assigned the number equal to the negative charge it would have as an anion. The less-electronegative atom is assigned the number equal to the positive charge it would have as a cation.
- **Rule 3:** Fluorine has an oxidation number of -1 in all of its compounds because it is the most electronegative element.
- **Rule 4:** Oxygen has an oxidation number of -2 in almost all compounds.
  Exceptions:
  Peroxides, such as H_2O_2, in which its oxidation # is -1
  When oxygen is in compounds with halogens, such as OF_2, its oxidation # is +2.
- **Rule 5:** Hydrogen has an oxidation # of +1 in all compounds that are more electronegative than it; it has an oxidation # of -1 in compounds with metals.
- **Rule 6:** The algebraic sum of the oxidation numbers of all atoms in a neutral compound is zero.
- **Rule 7:** The algebraic sum of the oxidation numbers of all atoms in a polyatomic ion is equal to the charge of the ion.
- **Rule 8:** Rules 1-7 apply to covalently bonded atoms; however, oxidation numbers can also be assigned to atoms in ionic compounds.

Oxidation & Reduction Notes – Chapter 19

oxidation-reduction reactions: reactions which involve changes in oxidation states due to an exchange of e^-s.

Also called ***redox reactions***.

oxidation: reaction where atoms or ions become ***more positive*** (***less negative***) by **losing e^-s**.

example:	$2Na(s)$	+	$Cl_2(g) \rightarrow$	$2NaCl(s)$
	↓			↓
oxidation #	**0**			**+1**

(oxidation # is more positive due to loss of 1e^-. Na has been oxidized.)

reduction: reaction where atoms or ions become ***more negative*** (***less positive***) by **gaining e^-s**.

example:	$2Na(s)$	+	$Cl_2(g) \rightarrow$	$2NaCl(s)$
			↓	↓
oxidation #			**0**	**-1**

(ox# is more negative due to gain of 1e^-. Cl_2 has been reduced.)

Use **OIL RIG** to Remember!!

*O*xidation *I*nvolves *L*oss of e^- (= more positive "+")
*R*eduction *I*nvolves *G*ain of e^- (= more negative "-")

Key Points:

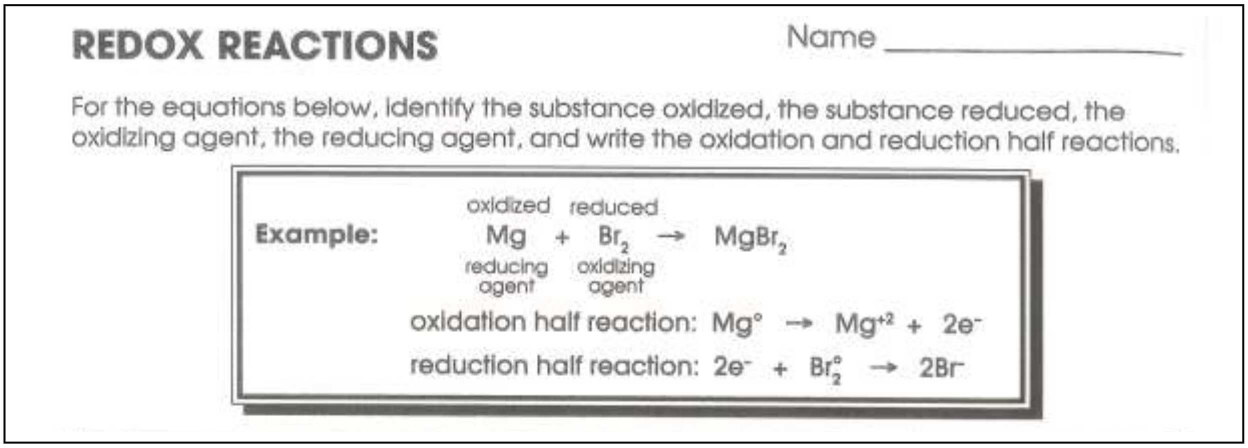
- **Oxidation** & **Reduction** always **occur together**.
- **e^-s** **lost** & **gained** must be ***equal.***
- If **oxidation #'s do not change**, it is ***NOT*** a **redox reaction**!!

example: $\underset{+4\quad -2}{\underset{\swarrow\ \searrow}{SO_2}} + \underset{+1\quad -2}{\underset{\swarrow\ \searrow}{H_2O}} \rightarrow \underset{+1\quad +4\quad -2}{\underset{\swarrow\ \downarrow\ \searrow}{H_2SO_3}}$

This is not a redox reaction!

The compound or element on the reactant side containing the oxidized element is the reducing agent (it causes reduction).

The compound or element on the reactant side containing the reduced element is the oxidizing agent (it causes oxidation).



REDOX REACTIONS Name ____________

For the equations below, identify the substance oxidized, the substance reduced, the oxidizing agent, the reducing agent, and write the oxidation and reduction half reactions.

Example: $\underset{\text{reducing agent}}{\overset{\text{oxidized}}{Mg}} + \underset{\text{oxidizing agent}}{\overset{\text{reduced}}{Br_2}} \rightarrow MgBr_2$

oxidation half reaction: $Mg^{\circ} \rightarrow Mg^{+2} + 2e^{-}$

reduction half reaction: $2e^{-} + Br_2^{\circ} \rightarrow 2Br^{-}$

Nuclear:

A student should be able to:

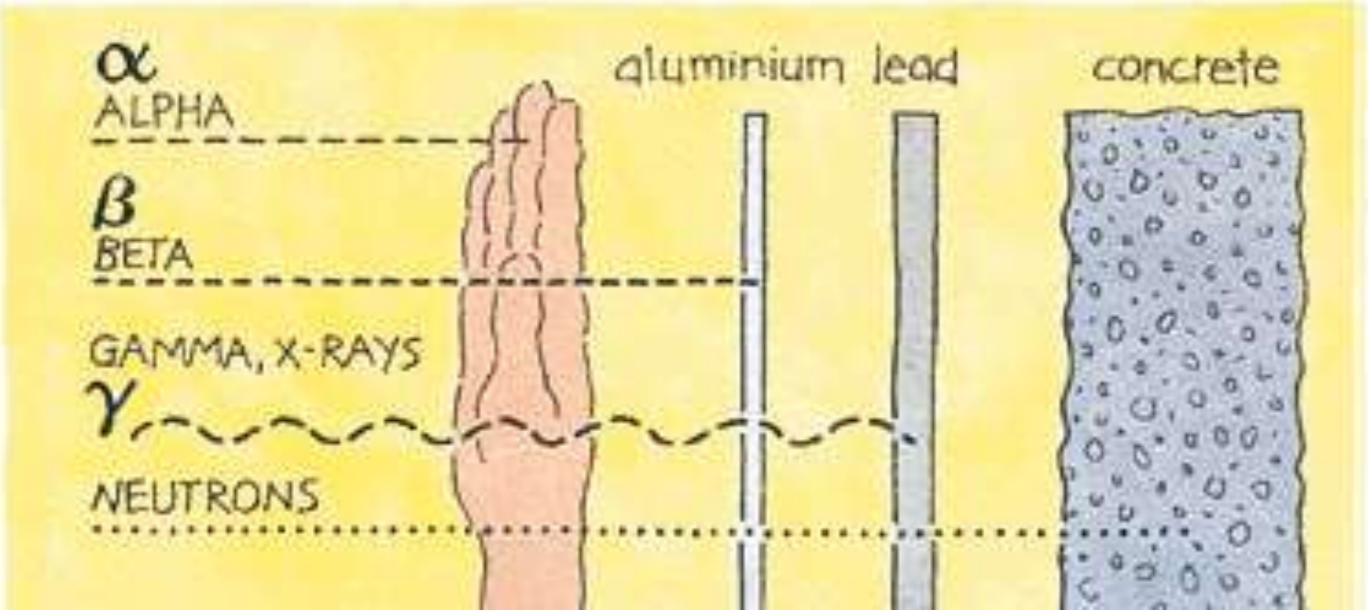
- Use the symbols for and distinguish between alpha ($_2^4He$), and beta ($_{-1}^0e$) nuclear particles, and gamma (γ) radiation include relative mass).
- Use shorthand notation of particles involved in nuclear equations to balance and solve for unknowns. Example: The neutron is represented as ($_0^1n$) .
- Discuss the penetrating ability of alpha, beta, and gamma radiation.
- Conceptually describe nuclear decay, including:
    - Decay as a random event, independent of other energy influences
    - Using symbols to represent simple balanced decay equations
    - Half-life (including simple calculations)
- Contrast fission and fusion.

Cite illustrations of the uses of nuclear energy, including, but not limited to: electricity, Carbon-14 dating, and radioisotopes for medicine (tracers, ionizing radiation, gamma sterilization, etc).

Nuclear Decay

Types of decay:

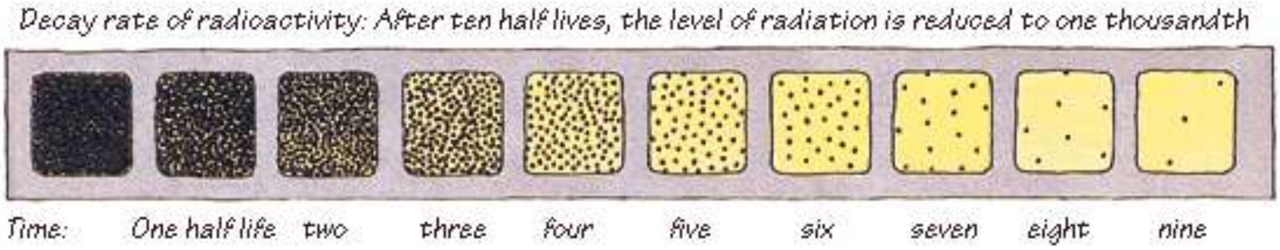
Name	Symbol	Mass	Shielding/Penetrating ability
Alpha	α $^{4}_{2}He$ (Helium nucleus)	largest mass	stopped by a sheet of paper
Beta	ß $^{0}_{-1}e$ (electron emission)	relatively small mass	stopped by thin metal
Gamma	γ $^{0}_{0}\gamma$ (energy, not written in the equation)	no mass; energy	stopped by thick lead or Concrete
Neutron	n $^{1}_{0}n$ (neutron)	relatively small mass	stopped by thick concrete



Nuclear equations:

Must be balanced on both sides by mass and atomic number!

HALF-LIFE



Half-life, $t_{1/2}$ is the time required for half the atoms of a radioactive nuclide to decay.

Radioactive dating uses knowledge of half-lives to approximate the age of an object. Radioactive dating includes using Carbon -14 to date organic material.

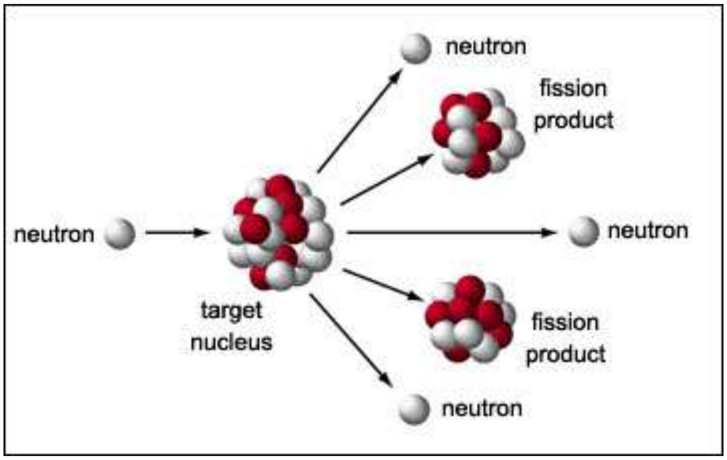
Half-life calculations:

Use ratio's to solve: $$\frac{\text{1half-life}}{\text{Time of half life}} = \frac{\text{\# of half-lives}}{\text{given amount of time}}$$

Must remember what a half-life means – ½ of the original amount decays

Time	amount
Original	100 %
1 half-life	50 %
2ne half-life	25 %
3rd half-life	12.5 %
4th half-life	6.25 %

Fission:



Fusion: