

Per a request, I offer an outline below which is a "study guide" for topics covered and things to know/learn for doing well on the exam

## **=== CHAPTER 1 ===**

- Quantities/Measurements
  - Number with Units
- Numbers:
  - standard notation (ordinary numbers)
  - scientific notation - how to express them with significand & exponent (power of 10)
  - verify your number: multiply significand and exponent
- Significant Digits (Significant Figures)
  - set by multiplication/division operations
  - Why? Indicator of precision
  - What makes significant digits?
    - from first non-zero (on the left) to last non-zero digit (on the right)
    - zeroes after decimal point after last non-zero digit
    - result: the factor with fewest significant digits sets the number of sig digits
- Decimal Places
  - set by addition/subtraction operations
  - Why? Part of the determination of precision process too
  - What makes decimal places?
    - determine count of digits after decimal point for all numbers in operation
    - result: the number with fewest decimal places sets number of digits after decimal point
- Priority of Math Operations (talked about in PEMDAS)

- 0. No intermediate rounding -- carry full precision until final result
- 1. Parentheses first — evaluate grouped expressions without rounding
- 2. Exponents — apply powers and roots next (roots are actually exponent operations)
- 3. Multiplication / Division — follow sig fig rules *after* full evaluation.
- 4. Addition / Subtraction — follow decimal place rules *after* full evaluation.
- Recap:
  - Addition/Subtraction → Round to the least number of decimal places.
  - Multiplication/Division → Round to the least number of significant digits.
- Best Practices
  - Avoid intermediate rounding unless explicitly required (e.g., reporting intermediate results).
  - Track precision mentally or with annotations, but apply rounding only at the final step.
  - Use parentheses liberally to clarify grouping and avoid ambiguity in mixed operations.
  - Label units and precision in each step to reinforce dimensional analysis and sig fig awareness.
- Conversion
- Conversion Factors - natural or "fixed" constants
  - They just express a number with its units in a different way
    - 2.5 g/mL and 2.5 kg/L are different ways of saying matter with the same density
    - 60 mi/h and ~26.8224 m/s are both the same velocity!
  - They do NOT affect the PRECISION in your calculations because they are NOT quantities that measurements: they are considered to have "infinite" significant digits & decimal places. Do not use rounded conversion factors within reason

- Conversions in chemistry are really products: the multiplication of one or more factors which are used as "factor" or "1/factor" (inverted) so that their units cancel the units of your quantity to get the units of the final/desired result
- Special name prefixes for powers of 10
  - Know kilo ( $10^3$ ), deci ( $10^{-1}$ ), centi- ( $10^{-2}$ ), milli-, micro-, nano-
- Density
  - Mass divided by volume
  - If you know any of the two values (mass, volume, density) you can calculate the third
  - Simple algebra to calculate
- Temperature
  - Know how to go from Celsius to Kelvin and back
    - $C = K - 273$
    - $K = C + 273$
  - You should understand also going from Fahrenheit and Celsius (testing your conversion ability). Possible to derive the equation but formula would be given
    - $F = 9/5 C + 32$
    - $C = 5/9 (F - 32)$

## === CHAPTER 2 ===

- Scientific Method
  - Terms and meaning
    - Observation
    - Hypothesis
    - Experiment
    - Theory (model)
    - Law

- Cycles of following the method show that science is re-evaluating its understanding of nature
- Atomic Theory
  - What did ancient Greeks not do in the scientific method (go to the "Experiment" stage)
  - Dalton's Atomic Theory
    - Be able to recognize what Dalton said about atoms and what he did not
      - atoms are smallest, nothing smaller (did not know about subatomic particles)
      - atoms of same element have same mass properties (did not know about isotopes)
      - compounds are combination of atoms of different elements (by definition)
      - chemical reactions are rearrangements of atoms, not the creation or destruction of them
- Atomic Mass Unit (amu)
  - useful way of measuring mass of subatomic particles
  - defined as 1/12th of the mass of the carbon-12 isotope
- Subatomic Structure
  - What's a nucleus in an atom?
  - Protons: mass = 1 amu (1.007 amu), charge = +1, part of nucleus
  - Neutrons: mass = 1 amu (1.009 amu) (only slightly bigger than proton), charge = zero, part of nucleus
  - Electrons: mass = 0.00055 amu (1800 times smaller than proton), charge = -1, "orbit" the nucleus
- Atomic Number
  - The number an element has on the Periodic Table, its identity by this number
  - also called the Z value

- also is the number of protons the element has
- Mass Number
  - It is the number (sum) of protons and neutrons in an atom
  - It is not the number of protons and neutrons in an element, since an element can have isotopes
  - also called the A value
  - This number identifies the isotope of an element
- Names and Symbols of Elements
  - These identify an element in the Periodic Table
  - Sodium => Na (comes from "natrium")
  - Potassium => K (comes from "kalium")
  - Carbon => C
  - Lithium => Li
  - Your memory won't be tested, but it is good idea to learn the symbols and what name they point to, at least for the first 3 periods, and many of the 4th and 5th period elements
- Isotopes
  - Elements with atoms can have more than one mass number (A value)
  - The count of protons and electrons does not affect what an isotope is: determined by different numbers of neutrons (this affects the mass number)
  - IMPORTANT: The chemical properties of elements with different isotopes is generally identical
  - Isotopes have a relative abundance: a percentage (fraction of 1) of the isotope's presence in nature
  - Abundance of all isotopes used to calculate the Atomic Weight (Atomic Mass) of an element as the "weighted average" of all the abundances
  - Calculation is: sum of the (relative abundance [fraction of 1]) \* amu of isotope

- What affects isotope stability? What is a sign of isotopic instability: radioactivity
- Atomic Notation
  - Has 3 things:
    - - the element symbol
    - - the mass number (A) as a superscript number to the left of symbol
    - - the atomic number (Z) as a subscripted number to the left of symbol
    - - one can omit the Z value because the element symbol will identify the Z value
- Periodic Table features
  - Periods (the rows)
  - Groups (the columns), sometimes group called "Family"
  - Element classes: metals, non-metals, metalloids (semi-metals)
  - special groups: alkali metals (Group 1), alkaline earth metals (Group 2), halogens (Group 17), noble gases (Group 18), transition metals (Groups 3 to 10 on right), the lanthanides and actinides (these are inserted after Group 2 and 3 in the 6th and 7th period elements)
- Terms: cations, anions (in addition to many terms being used in chemistry)

### === CHAPTER 3 ===

- Waves
  - THERE IS A LOT OF DETAIL in the points below: I am trying to connect the relationship between waves and what happens in the atom. I will indent another level to indicate you don't need to know this but it helps make sense. These are finer points in physics that are to be learned by chemistry students, engineers and physicists
  - components of wave and what they signify: wavelength ( $\lambda$ ), frequency, ( $\nu$ ), speed of wave ( $v$ ), equation  $v = \lambda * \nu$
  - other wave features: crest, trough (named parts of a wave to indicate its sinusoidal description)

- waves also have amplitude: amplitude is a measure of the energy of a wave
- why waves? electrons have a charge (this is an electric field). when electrically charged particles have motion like an orbit about a nucleus, they generate a magnetic field perpendicular to the electric field
  - These oscillating perpendicular waves of electric and magnetic fields caused by a moving (orbiting) electron are what is called "electromagnetic radiation": this "radiation" has ENERGY; electrons with this energy can "transfer" this energy to another electron in another atom by emitting (releasing this energy) in the form of an EM wave (a "photon") which can be absorbed by another electron which takes its energy
- These energetic EM waves (EM radiation) form the EM spectrum, which includes visible light
- the ENERGY of this EM wave, this EM radiation can be determined mathematically using Planck's Equation:  $\text{Energy} = \text{Planck's constant} \times \text{the FREQUENCY } (\nu) \text{ of the wave}$
- Since the EM wave/radiation ("photon") energy is determined by frequency, HIGHER frequency (ipso facto SHORTER wavelength) EM waves have MORE ENERGY than LOWER frequency (ipso facto LONGER wavelength) EM waves (photons)
- In the spectrum (KNOW/RECOGNIZE these regions or zones of EM radiation/rays)
  - Gamma radiation/rays have higher frequency / shorter wavelength than X-rays
  - X-radiation/rays have higher frequency / shorter wavelength than UV (ultraviolet) radiation
  - UV radiation/rays have higher frequency / shorter wavelength than visible light radiation/rays
  - *UV, X- and gamma radiation/rays are called "ionizing radiation" because they have the energy to cause atoms to lose one or more electrons in their orbits, affecting the chemistry of the atoms, and in living individuals, that affected chemistry can lead to radiation sickness*

- Visible light radiation/rays have higher frequency / shorter wavelength than infrared radiation/rays
  - Infrared (IR) light radiation/rays have higher frequency / shorter wavelength than microwave radiation/rays
  - Microwave radiation/rays have higher frequency / shorter wavelength than radio radiation/rays/waves that allow you to hear music from a local transmitter on the FM band
- The visible light part of the EM spectrum goes from violet to blue to green to yellow to red
  - The color list shown above go from higher frequency/shorter wavelength to lower frequency/longer wavelength from left to right
- Know how to calculate wave speed, wavelength, frequency from the equation if you have any two of the three values
- Know how also to use conversion factors in doing this. Practice problems to follow
- History of Discovering the Atom and its Structure
  - What did Thomson say about atomic structure
  - How did Rutherford add to knowledge with gold foil experiments (a nucleus that has positive charges, atoms have particles but "space" also)
  - How did Bohr refine model: electrons, orbits, fixed energy levels, absorption and emission of EM waves/radiation ("photon"), concept of ground state and excited state, used his model to explain hydrogen (H) atom emission spectrum
- Energy level transitions: Can just any amount of energy be absorbed by an electron? or a certain amount is required?
  - continuous change in energy level versus "quantized" change in energy level
  - How do observations in an emission spectrum explain transitions of electrons between fixed energy levels?
- Relationship between energy level transitions of electrons and the first (principal) quantum number ( $n$ )



- atoms of periodic table can have electrons represented as orbits of "shells" have integer  $n$  values ( $n = 1, 2, 3, \dots$ )
- The number of electrons in these shells has a fixed maximum that follows the way the Periodic Table is constructed
- Electron "identity"
  - each electron in atom has four quantum numbers ( $n, l, m_l, m_s$ )
  - an "orbital" is a region of space that an electron can occupy in the atom. One orbital can hold two electrons only, and they must be of opposite spin. (see the  $m_s$  quantum number)
  - the  $n$  quantum number really describes the (relative) energy--the energy level--of the electron in the atom
  - the  $l$  quantum number describes the "shape" or the orbit--which is a geometric representation of the space an electron can be found--in that energy level given by  $n$ 
    - the shape of the orbital is described by the letter designations  $s$  (spherical),  $p$  (bilobed shape),  $d$  (more lobes, one which is bilobed with a torus ["donut"]), and  $f$  (a really complex formation of lobes). The  $l$  values of those  $s, p, d,$  and  $f$  subshells are 0, 1, 2, 3 respectively
  - The  $m_l$  quantum number applies only to  $l = 1$  and greater (that is, to  $p, d,$  and  $f$  type orbitals) and they pertain to how these orbital types are oriented when the space is viewed from a 3D  $x-y-z$  coordinate system space
  - The  $m_s$  quantum number describes that an electron can have one of two spins, which happen to be given values of  $1/2$  and  $-1/2$  (important for the math of quantum mechanics). This just shows electrons of opposite spins can be paired into one orbital, and that pairing requires opposite spins. Electrons would rather be unpaired during orbital filling of electrons in the Periodic Table, but they can be paired when required
  - Be familiar with the shapes (geometric representations) of the  $s, p,$  and  $d$  orbitals
- Know the different terms of the quantum mechanics model: shell, subshell, energy level, energy sublevel, orbital, and how they are used synonymously or interchangeably

- Adding electrons to atoms to get configurations
  - Electrons are always added to the atom to achieve the greatest stability, that is to lower the energy
  - Understand the Aufbau Principle (which is about recognizing finding the lowest energy, most stable configuration)
    - The orbitals involved are 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, and so on
    - The n value is the number part: 1, 2, 3...
    - The l value is the integer representation of the letter part: s = 0, p = 1, d = 2, f = 3 (that's it)
    - There are two parts
      - the n + l rule fills first
        - So 2s = 2 + 0 = 2 and 2p = 2 + 1 = 3, so 2s will fill before 2p
        - So 3d is 3 + 2 = 5 and 4s is 4 + 0 = 4. So the 4s orbital will get electrons before 3d (!)
      - Breaking the tie: if two orbitals have equal n + l, for instance, 4f and 5d. Both have n + l = 7. In that case the orbital with the smaller n fills first, so 4f before 5d.
  - **Pauli Exclusion Principle:** this says that each electron has in effect an “identity”: one set of those four quantum numbers  $n$ ,  $l$ ,  $m_l$ , and  $m_s$
  - **Hund's rule:** electrons will occupy degenerate orbitals (orbitals of the same energy) singly and with parallel spins before pairing up. This minimizes **electron-electron repulsion** and maximizes **total spin**, leading to a more stable, lower-energy configuration.
    - Applies to **p, d, and f orbitals**, which contain multiple orbitals of the same energy.
    - Electrons prefer to **spread out** before they **double up**.
    - All unpaired electrons in a subshell will have the **same spin direction** (either all  $\uparrow$  or all  $\downarrow$ ).

- Think of electrons like passengers on a bus: Everyone grabs their own seat first. Only when seats run out do they start sitting together. This keeps things “less crowded” and energetically favorable.
- The atom of every element has an electron configuration. This will appear as  $nl^e$  where  $n = 1, 2, 3, \dots$  to indicate the shell number which corresponds to the period (row) of the Periodic Table; the  $l$  value will actually be  $s, p, d$ , or  $f$ , to indicate the subshell of the shell ( $l = 0, 1, 2, 3$  by correspondence with  $s, p, d, f$ ); and that superscript  $e$  will be a number to indicate the number of electrons in the subshell of the shell. The  $s$ -type ( $l = 0$ ) has only one orbital of its type, and thus an  $s$ -type orbital can have maximum two electrons. The  $p$ -type ( $l = 1$ ) orbital can have three such orbitals, and thus the  $p$ -type can range from 1 to 6 (three orbitals, always 2 electron per orbital, thus 1 through 6). The  $d$ -type ( $l = 2$ ) orbitals can have five such orbitals, and  $d$ -type range in number from 1 to 10 (five orbitals, always 2 electrons max per orbital, thus 1 through 10). Finally the  $f$ -type ( $l = 3$ ) can have seven such orbitals, and thus  $f$ -type range from 1 to 14 (seven orbitals, always 2 electrons max per orbital, thus 1 through 14).

Important to know: no orbital can exist where  $l$  is not less than  $n$ , i.e.,  $l < n$  always.

This is why there are no  $1p$  ( $n = 1, l = 1$  for  $p$ -type) type orbitals, no  $2d$  ( $n = 2, l = 2$  for  $d$  type), and no  $3f$  ( $n = 3, l = 3$  for  $f$  type) orbitals

- How many  $s$ -type orbitals? One. Electron configuration look: thus you can only see  $ns^1$  or  $ns^2$  in a configuration (where  $n = 1, 2, 3, \dots$ )
- How many  $p$ -type orbitals? Three. Electron configuration look: thus you can only see  $np^1, np^2, np^3, np^4, np^5, np^6$  in a configuration (where  $n = 2, 3, \dots$  1p orbitals do not exist)
- How many  $d$ -type orbitals? Five. Electron configuration look: thus you can only see  $nd^1, nd^2, nd^3, nd^4, nd^5, nd^6, nd^7, nd^8, nd^9, nd^{10}$  in a configuration (where  $n = 3, 4, \dots$  2d orbitals do not exist)
- How many  $f$ -type orbitals? Seven. Electron configuration look: thus you can only see  $nf^1, nf^2, nf^3, nf^4, nf^5, nf^6, nf^7, nf^8, nf^9, nf^{10}, nf^{11}, nf^{12}, nf^{13}, nf^{14}$  in a configuration (where  $n = 4, 5, \dots$  3f orbitals do not exist)
- These patterns correspond to what we see in the Periodic Table  
 Period 1- H:  $1s^1$ , He:  $1s^2$   
 Period 2-Li:  $1s^2 2s^1$  or  $[\text{He}]2s^1$ , Be:  $[\text{He}]2s^2$ , B:  $[\text{He}]2s^2 2p^1$ , C:  $[\text{He}]2s^2 2p^2$ ,  
 N:  $[\text{He}]2s^2 2p^3$ , O:  $[\text{He}]2s^2 2p^4$ , F:  $[\text{He}]2s^2 2p^5$ , Ne:  $[\text{He}]2s^2 2p^6$   
 Period 3-Na:  $[\text{Ne}]3s^1$ , Mg:  $[\text{Ne}]3s^2$  Al:  $[\text{Ne}]3s^2 3p^1$ , Si:  $[\text{Ne}]3s^2 3p^2$ ,

P: [Ne]3s<sup>2</sup>3p<sup>3</sup>, S: [Ne]3s<sup>2</sup>3p<sup>4</sup>, Cl: [Ne]3s<sup>2</sup>3p<sup>5</sup>, Ar: [Ne]3s<sup>2</sup>3p<sup>6</sup>

Period 4-K: [Ar]4s<sup>1</sup>, Ca: [Ar]4s<sup>2</sup>, Sc: [Ar]4s<sup>2</sup>3d<sup>1</sup>, Ti: [Ar]4s<sup>2</sup>3d<sup>2</sup>, V: [Ar]4s<sup>2</sup>3d<sup>3</sup>,  
Cr: [Ar]4s<sup>2</sup>3d<sup>4</sup>, Mn: [Ar]4s<sup>2</sup>3d<sup>5</sup>, Fe: [Ar]4s<sup>2</sup>3d<sup>6</sup>, Co: [Ar]4s<sup>2</sup>3d<sup>7</sup>, Ni: [Ar]4s<sup>2</sup>3d<sup>8</sup>,  
Cu: [Ar]4s<sup>2</sup>3d<sup>9</sup>, Zn: [Ar]4s<sup>2</sup>3d<sup>10</sup>, Ga: [Ar] 3d<sup>10</sup>4s<sup>2</sup>4p<sup>1</sup>, Ge: [Ar] 3d<sup>10</sup>4s<sup>2</sup>4p<sup>2</sup>,  
As: [Ar] 3d<sup>10</sup>4s<sup>2</sup>4p<sup>3</sup>, Se: [Ar] 3d<sup>10</sup>4s<sup>2</sup>4p<sup>4</sup>, Br: [Ar] 3d<sup>10</sup>4s<sup>2</sup>4p<sup>5</sup>, Kr: [Ne] 3d<sup>10</sup>4s<sup>2</sup>4p<sup>6</sup>

- In the electron configurations for the first 4 periods above, in the 4<sup>th</sup> period, you will note that 3d<sup>10</sup> was written before 4s<sup>2</sup> although the filling order is that 4s will be filled before 3d (under the  $n + l$  rule: which also states 3d is filled before 4p)
- Know what **valence** and **core** electrons are. Valence electrons are in the outermost shell, and they will be involved in bonding, no matter what kind of bonding. The core electrons are all the electrons that are closer to the nucleus, not involved in bonding.
- Be able to recognize the s-, p-, d-, and f-block of elements on the Periodic Table. Their orbitals start to fill with electrons in those special blocks.
- Be aware that the p-block elements start in period 2 (there are no 1p orbitals)
- Be aware that the d-block elements start in period 4, starting the fill with 3d (there are no 1d or 2d orbitals), and they are offset by one period from the 3s and 3p orbital elements
- Be aware that the f-block elements start in period 6, starting the file with 4f (there are no 1f, 2f, or 3f orbitals) and they are offset from 4s and 4p orbital elements by two rows/periods
- Understand the Aufbau Principle two rules: 1.  $N + L$  2. When  $N+L$  are equal for two orbital types, like 3d and 4p or 5d and 4f, then filling starts with orbital type with lower  $n$  (3d before 4p, 4f before 5d)
- Atomic Radius: atomic radii get bigger going DOWN the table and moving from right to left of table
- Ionization Energy: the energy need to push one electron out and make the atom have a positive charge—gets takes LESS energy to ionize going DOWN the table and moving from right to left of table
- Multiple ionization: it take progressively MORE energy to push more electrons away from the atom. Suppose pushing the 1<sup>st</sup> electron out takes X energy (kJ/mol). Pushing the 2<sup>nd</sup> electron out won't take X kJ/mol, but maybe 2-3X kJ/mol, and get even higher pushing the 3<sup>rd</sup> electron out. The creation of high charge separation of increasing magnitude is very energetically destabilizing
- Electron Affinity: this is the energy release when an atom takes on an electron. The larger the kJ/mol (usually much less than zero as energy release), the more there is

an affinity. Atoms on the right of the table want to add electrons to themselves because they complete their valence shell maximum electron number, which is a very stable situation for atoms. In this way, they want to be like the Group 18 at the very right end Noble Gas elements. Of course in taking on electrons, they acquire a net NEGATIVE charge, and that too influences their chemistry. (Noble gases have complete valence shells so they don't have to take on electrons and thus have zero charge). ON the left side of the table, those elements want complete valence shells, but taking on electrons in their period is no good for them since they will not complete those valence shells: instead, they want to get RID of electrons by which they will have complete valence shells, but they look like the elements of the previous period (one row up of the table). This is why their ionization energies and not electron affinities are small by comparison.

#### === CHAPTER 4 ===

- Matter has properties
  - Physical properties
  - chemical properties

Matter can undergo changes of state: solid, liquid, gas

Know the terms: vaporization, freezing, condensation, distillation, precipitation, filtration, solvent extraction, chromatography

- Law of conservation of mass: relationship to chemical reactions
- Mixtures: homogeneous, heterogenous
- What's a compound?
- What's a
  - Chemical formula
  - Molecular formula
  - Empirical formula
- Elements in their pure natural states
  - Atomic elements existing as individual atoms: usually the metals like Au (gold), silver (Ag), zinc (Zn), sodium (Na)...countless examples
  - Molecular elements: usually the nonmetals. Of the nonmetals, 7 of them exist as diatomic elements:  $H_2$ ,  $N_2$ ,  $O_2$ ,  $F_2$ ,  $Cl_2$ ,  $Br_2$ ,  $I_2$ . Some like sulfur and phosphorus exist in a very polyatomic arrangement.
- What's a formula unit?
- Ionic compounds: where the atoms bond with each other electrostatically, where one atom gives up an electron to become positively charged (typically a metal) to

another atom (typically a nonmetal) which becomes negatively charged, and by that transfer of one or more electrons, they become attracted to each other because of the creation of these electric charges. Calcium can give up TWO electrons to two fluorine atoms, which each become negatively charged (-1 charge state), and form the  $\text{CaF}_2$  salt. Potassium can give up one electron to chlorine which readily takes it, to form KCl salt. Vanadium can lose 3 and 4 electrons to three and four chlorine atoms to form  $\text{VCl}_3$  and  $\text{VCl}_4$

- What is the Stock system? It is used in the name of compound where the cation—typically a metal—has MULTIPLE potential ionic states. Copper can lose one or two electrons to have a  $\text{Cu}^+$  or  $\text{Cu}^{2+}$  ionic (cationic) charge. So the compound “copper bromide” can refer to copper(I) bromide  $[\text{CuBr}]$  and to copper(II) bromide  $[\text{CuBr}_2]$ . Stock system → the way of naming elements with multiple stable ionic charge states using the Roman numeral system
- Criss-cross method: the way of figuring out how many cations (monoatomic or polyatomic) and anions (monoatomic or polyatomic) are involved forming the compound, enabling you to determine the subscript of the cations and anions in the ionic compound. (This is a kind of stoichiometric determination). It includes dividing the subscript by the greatest common divisor until any one of the subscript = 1, with the 1 not being written for the subscript.
- **Polyatomic** ions are MOLECULES of typically several atoms bonded together (as generally wholly or partially covalent [“shared”] bonds that will form a STABLE ion. (Some molecules that are ions form unstable ions as the molecule is reactive and will not maintain that ionic state—those are “molecular ions” and not “polyatomic ions”). These polyatomic ions are typically composed of nonmetal elements like B, C, N, O, P, S, Cl, Br, I, and a few others, and the H atom is also involved in these polyatomic groups [hydrogen is not considered a nonmetal but is also not fully a metal, but is “metal-like”]. Polyatomic molecules can form ions with single element cations  $[\text{Na}_3\text{PO}_4, \text{K}_2\text{SO}_4, \text{Ca}(\text{NO}_3)_2, \text{etc}]$  or anions  $[\text{NH}_4\text{Cl}, \text{NH}_4\text{Br}]$  or with each other  $[(\text{NH}_4)_3\text{PO}_4, (\text{NH}_4)_2\text{SO}_4]$  (did you notice there are not too many polyatomic cations? typically it’s ammonium)
- What is an ionic compound and what is likely NOT an ionic compound?
  - Is cation a monoatomic metal element and anion a monoatomic nonmetal element? Likely ionic
  - Are all the elements in the compound nonmetal elements? Likely NOT ionic.  $\text{CO}_2, \text{NH}_3, \text{NO}_2, \text{N}_2\text{O}_2, \text{SOCl}_2$
  - Is there a clearly polyatomic molecule/group that is recognizably ionic? Then ionic
  - The T43 Method is used to show you the following

- How many oxygen atoms will bond to the nonmetal element in forming that polyatomic ion
- What the magnitude of the negative charge will be on that polyatomic ion

- NAMING

- For metal cations, there are Type I and Type II cations.
- With Type I, that metal only ever has one form of a cation, Na is 1+ ( $\text{Na}^+$ ), K is 1+ ( $\text{K}^+$ ), Ca is 2+ ( $\text{Ca}^{2+}$ ), Mg is 2+ ( $\text{Mg}^{2+}$ ), Li is 1+ ( $\text{Li}^+$ ), Al (aluminum) is 3+ ( $\text{Al}^{3+}$ ). The charge never varies.
- With Type II, that metal can have more than one form of the magnitude of the positive charge state. Cu (copper) can be 1+ or 2+ ( $\text{Cu}^+$ ,  $\text{Cu}^{2+}$ ), Fe (iron) is 2+ or 3+ ( $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$ ), Pb (lead) is 2+ or 4+ ( $\text{Pb}^{2+}$ ,  $\text{Pb}^{4+}$ )
- For type I, take the name of the metal, take the stem or base name of the nonmetal (“ox” for oxygen, “chlor” for chlorine, “sulf” for sulfur, “fluro” for fluoride, “nitr” for nitrogen and more) and put suffix “-ide” on it.
- For type II, you take the metal element name as for type I, add the Roman numeral for the charge variant, and then for the nonmetal, you do the naming as for type I (base + “-ide”)
- For polyatomic ions, the cationic form is usually just “ammonium” ( $\text{NH}_4^+$ ) although there are a couple of other rare others. For anionic forms, there are many you should recognize:  $\text{CO}_3^{2-}$  is “carbonate”,  $\text{NO}_3^-$  is “nitrate”,  $\text{PO}_4^{3-}$  is “phosphate”,  $\text{SO}_4^{2-}$  is “sulfate”,  $\text{BO}_3^-$  is “borate” and the halogen forms too:  $\text{ClO}_3^-$  (chlorate),  $\text{BrO}_3^-$  (bromate),  $\text{IO}_3^-$  (iodate). Other polyatomic ions and their names (look for a pattern):  $\text{NO}_2^-$  (nitrite),  $\text{PO}_3^{3-}$  (phosphite),  $\text{SO}_3^{2-}$  (sulfite)