You Know You Wanted It...Chemistry Exam 1 Review Sheet!

Need to know the following concepts:
- mole
- using balanced equations (to find amount of product made/other reagent needed)
- Unit conversion
- Molecular weight
- how to use Density/ Molarity (mol of stuff/ L of solution) as unit conversion

-Democritus (5th cent. B.C.): matter composed of small, individual, indivisible particles; theory not believed at the time

-Antoine Lavoisier (1775): Law of Conservation of Mass-- mass cannot be created or destroyed, but it can change forms

-Provst (1799): Law of Definite Proportions—any sample of a pure compound has the same mass percentage of its elements (example—H₂O is 89% H and 11% O)

-John Dalton (1808): Atomic theory of matter
  1. elements are made up of individual/indivisible parts called atoms
  2. atoms of the same element have same mass and chemical properties
  3. in order to form compounds, atoms combine in whole numbers (you never see ½ H in a compound)
  4. Chemical reactions only change the way the atom is arranged, not the atoms itself.

-J.J. Thompson (late 1800s): later discovered the electron, and that it was a particle common to all atoms

-Rutherford’s Gold Foil Experiment (1911): In this experiment, Rutherford fired alpha particles at a gold foil, behind which was a fluorescent screen. Every time an atom passed through the foil, it would show up on the screen. Although most atoms passed directly through, some deflected at great angles, when hitting the foil. As a result, Rutherford proposed that the atom had a solid core, and the idea of the nucleus was born.

-Mass number (A) = number of protons + number of neutrons
-Atomic number (Z) = number of protons = number of electrons

-Niels Bohr (1913)- Bohr model of atom; e⁻ have specific orbits, therefore, specific energies; implied that e⁻ could only absorb/emit specific amts. of energy; energy is quantized

-Emission and absorption spectra: these differ from gas to gas- only certain colors on the line spectrum are observed, and the emission and absorption spectra are related, because…
  - The electrons absorb energy to jump to a higher level. They will only absorb a quantized amount of energy. When this happens, the electrons jump to a higher energy level within the atom—the absorption spectrum shows certain colors. When the electron jumps back down, it emits energy in the form of a photon. Although the emission and absorption spectra are related, you only see one side of it. That is, when looking at emission, excitation is not caused by light, and when viewing absorption, emission is not how e⁻ return to their original state.

-Fun equations (and constants)!
  - \( c = 3.00 \times 10^8 \text{ m/s} \) (speed of light), \( h = 6.626 \times 10^{-34} \text{ Js} \) (Planck’s constant), \( 1.097 \times 10^7 \text{ m}^{-1} \) (Rydberg constant)
  - \( \lambda \) is wavelength (measured in meters), \( \nu \) is frequency (measured in Hertz or s⁻¹), \( E \) is energy (measured in Joules)
  - Energy equations: \( c = \lambda \nu \quad E = h\nu = hc/\lambda \)
  - Energy of orbital: \( E = -Rhc/n^2 \); \( n \) = energy level = integer > 1
    - Also: \( E = -Rhc((1/n^2_{\text{final}})-(1/n^2_{\text{initial}})) \)
    - \( Rhc = 2.179 \times 10^{-18} \text{ J/atom or 1312 kJ/mol} \)
  - Rydberg equation: \( 1/\lambda = R(1/2^2-1/n^2) \) but only for H₂, and n>2

-Electrons want to be close to center, further away they are, more energy is needed to keep them there…so, electrons jumping from low to high energy level need energy, therefore absorb energy; electrons jumping from high to low energy levels have extra energy, therefore emit energy as photon.

-DeBroglie (1925)- e⁻ behave as waves and particles; thus, cannot know both position and velocity of e⁻ at a given time

-Heisenberg Uncertainty Principle: \( \Delta x (\Delta \nu) \geq \hbar / 4\pi \); \( \Delta x = \) uncertainty in position, \( \Delta \nu = \) uncertainty in velocity
- Schrödinger (late 1920s): e⁻ could be described as wave function (ψ); gives energy of e⁻; only specific values of ψ possible; ψ² gives probability function—probability of finding e⁻ in given space

- Quantum Numbers:
  - n= principle QN; n>0 (integer); n=1,2,3,4,…; specifies energy level where e⁻ is found; lower n = lower energy level = closer to nucleus
  - l= angular momentum QN; any integer from 0 to n-1; number of possible l’s corresponds to # of subshells in given shell; determines shape of orbital (“S” orbital: l=0, “P” orbital: l=1, “d” orbital: l=2, “f” orbital: l=3)
  - m_l= magnetic QN; -l ≤ m_l ≤ l; # of possible m_l determines that type of orbital
  - m_s= magnetic spin QN; m_s= +½ or -½; allows 2 e⁻ per orbital

- Example: nitrogen (N)—1s²2s²2p³  Quantum #s of valence e⁻: n=2, l=1, m_l=1, m_s= +½

- Pauli Exclusion Principle: no 2 e⁻ can have the same 4 quantum numbers

- Ground state is the lowest energy configuration; we want to fill lowest energy orbitals
- Energy determined by: distance from nucleus (lower n= lower energy); presence of other e⁻ (e⁻ tend to repel each other due to like charges); shape (s<p<d<f in increasing energy order—fewer nodal pts. = lower energy)
- Lowest energy orbital is 1s—filling order is as follows (diagonal rule):

- Valence e⁻—full and half-filled subshells are most stable

- Examples:
  - potassium (K) Z=19  1s²2s²2p⁶3s²3p⁶4s¹ OR [Ar] 4s¹ (noble gas config.)
  - bromine (Br) Z=35  1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁵ OR [Ar] 4s²3d¹⁰4p⁵

- Hund’s Rule: e⁻ configuration will maximize number of unpaired electrons
  - spinning e⁻ creates magnetic field; unpaired e⁻ have mag. moment and react to mag. field
  - unpaired e⁻ = paramagnetic; no unpaired e⁻ = diamagnetic

** You will also find it useful to know how to convert between units (like on the first quiz) and from grams to moles, and vice versa. Keep in mind that many problems will require unit conversions, calculations of molecular and atomic mass, and percent composition. Most of this can be found by looking at chapters 1 and 2 of the Chemistry textbook. Here is some important information…

- Avogadro’s Number = 6.02 x 10²³ (1 mole of anything has this many units)
- Moles to mass: Moles x (grams/1 mol) = grams
Mass to moles: Grams x (1 mol/grams) = moles
✓ Density = mass/volume
✓ Atomic weight=
  ((% abundance isotope one)/100)(mass isotope 1)+((% abundance isotope two)/100)(mass isotope 2)+...
✓ Percent abundance= (# of atoms of given isotope/total # of all isotopes of element)(100)

Learn “Sig. Fig.” rules (pp. 38-41 in textbook)

| Common Prefixes form Metric System (more can be found on pg. 27 of textbook) |  |
|---|---|---|
| mega- | M | $10^6$ |
| kilo- | k | $10^3$ |
| deci- | d | $10^{-1}$ |
| centi- | c | $10^{-2}$ |
| milli- | m | $10^{-3}$ |
| micro- | µ | $10^{-6}$ |
| nano- | n | $10^{-9}$ |

Practice Problems:

1) What are the mass number (A), the atomic number (Z), and the number of electrons in $^{31}$P?
2) You have 0.539 mol of Au, and each gram of Au sells at $15.00, how much money can you make from selling the Au?
3) Name the following compounds:
   a. MgS
   b. NaCl
   c. CoCl$_2$
4) Determine the electron configuration of the following elements (regular and noble gas configuration). Which ones are considered paramagnetic? Diamagnetic?
   a. Si
   b. Mn
   c. Sb
   d. Pb
   e. Ca
5) Blue light has a wavelength of about 400 nm. What is the frequency of blue light? Calculate the energy of one photon of blue light. Now calculate the energy of 1.0 mol of photons from the blue light.

Problems are from the chemistry textbook, class notes, and me!