## Unit 0 – Basics

Honors Unit 1 – Chemistry Basics

1. When multiplying/dividing numbers, round based on smallest number of sig figs in original numbers.
2. When adding/subtracting numbers, round based on smallest number of decimal places in original numbers.
3. Zeros - Leading = not significant, trapped/captive = significant, trailing w/decimal = significant, trailing w/o decimal = not significant.
4. Dimensional analysis! For so many things all year!!!
5. Mass is conserved during chemical and physical changes.
6. Compounds can be separated into elements by chemical Δs, mixtures can be separated by physical Δs.
7. Filtering separates mixtures based on differences in particle size...the large particles are trapped on the filter paper while the soluble component goes through the filter paper and stays in the "filtrate".
8. Distillation separates mixtures based on differences in boiling point.
9. Chromatography separates mixtures based on differences in polarity.
10. Paper chromatography - component that is most similar in polarity to the "mobile phase" moves up the farthest.
11. Always read a graduated cylinder from eye level.
12. When reading a volume of a liquid in a container, you can estimate by reading in between the graduated markings. That can give you one more sig. fig. in your volume.
       You would read the volume on this buret as 37.30 mL.... NOT 38.70 mL.
13. Ranking measuring devices from least precise to most precise→ beaker, graduated cylinder, volumetric flasks, burette
       (The volumetric flask only has ONE line on it to measure one specific volume.)
14. Density = mass/volume
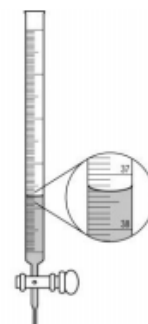
Honors Unit 2 – Atomic Structure

15. Isotopes = different number of neutrons, ions = different number of electrons, different number of protons = new element with new name.
16. Dalton's postulates, mostly right but wrong about – atoms being indivisible smallest particles (protons, neutrons, electrons exist!), all atoms of same element identical (isotopes exist!)
17. Rutherford's Gold Foil Exp. – discovered dense, positively charged nucleus, and that atom is mostly empty space.
18. J.J. Thomson's Cathode Ray Tube – discovered negatively charged electron.

Honors Unit 3 – Electrons

19. Each orbital can hold two electrons. 1 s orbital per energy level, 3 p orbitals, 5 d orbitals, 7 f orbitals.
20. Aufbau principal = fill lowest energy orbitals first, Hund's = if more than one orbital at the same energy level put one electron in each orbital before doubling up, Pauli Exclusion = one electron spin up, one spin down.

Honors Unit 4 – Periodic Trends

21. Rows = periods, columns = groups/families
22. Everything in the same group have similar properties because they have the same number of valence electrons.
23. Properties are largely based on valence electrons.

Honors Unit 5 – Bonding and Structure

24. MEMORIZE YOUR COMMON IONS!!!!! Seriously...
25. Count your valence electrons before you start drawing a Lewis Structure! Place atoms, single bond everything, give everything a full shell, recount electrons used. Too few – add lone pairs to central atom. Too many – erase a pair from two neighboring atoms and put another bond between them to make a double bond (repeat as necessary, can do again between same atoms to make a triple bond).
26. Strength of IMFs and Bulk Solids – LDF < Dipole-Dipole < H-bond < Ionic < Metallic < Network Covalent
27. Increase in IMFs = increase boiling point, melting point, surface area, viscosity

Honors Unit 6 – Reactions

28. Write good valid, neutral formulas before trying to balance an equation.
29. Main types of reactions – synthesis, decomposition, single replacement, double replacement, combustion.

Honors Unit 7 – Stoichiometry

30. Molar mass = sum of each atom in the formula, g/mol
31. Mole highway pathways – grams A → moles A → moles B → grams B, etc. Lots of pathways!
32. Mole Ratio to convert between moles A and moles B.

Honors Unit 8 – Advanced Chemical Ratios

33. The % composition by mass for a pure compound does not change.
34. Empirical formula rhyme→ % to mass, mass to mole, divide by small, times until whole...Get the simplest whole # ratio of the moles (or atoms) in the compound.
35. The molecular formula for a compound is a whole # multiple of the empirical formula ratio.
36. % yield = (experimental / theoretical) x 100
37. % error = [|(experimental - theoretical)| / theoretical] x 100
38. The amount of product for a reaction is determined by the limiting reactant.

Honors Unit 9 – Gas Laws

39. Ideal Gas Law, $PV = nRT$ (all other gas laws can be derived from this one!)
40. Dalton's Law of Partial Pressures, $P_{total} = P_1 + P_2 + P_3...$
41. When collecting gas over water, subtract out the water vapor pressure to get the pressure of the dry gas.

Honors Unit 10 – Thermochemistry

42. $Q = mC\Delta T$
43. Calorimetry, $Q_{in} = Q_{out}$
44. Heating/cooling = $mC\Delta T$, phase changes = mL
45. Takes energy to break a bond = endothermic, energy is released when bonds form = exothermic, ALWAYS!!!!

Honors Unit 11 – Solutions

46. Molarity = mol/L

Honors Unit 12 – Kinetics

47. Rate affecting factors = temperature, concentration/pressure, surface area, catalyst
48. Rate laws, orders are found experimentally by comparing initial rate data for different trials.
49. Rate constant has units based on the overall order of the reaction

Honors Unit 13 - Equilibrium

50. Le Chatelier's Principal can only be used for systems that are ALREADY at equilibrium. A system at equilibrium will shift to lessen the effect of a stressor being applied. It will reach a new equilibrium position.
51. Solids and liquids do not affect the position of equilibrium.
52. Equilibrium expression also called Law of Mass Action, products over reactants $K = [C]^c[D]^d/[A]^a[B]^b$
53. Exponents are the coefficients in the balanced equation.
54. Each reaction has its own K value, and the K value changes ONLY with a change in temperature.
55. Compare Q to K to decide if you are at equilibrium yet, and if not, which way you need to shift $Q < K$, not enough products, forward rate increases to make more products. $Q > K$, too many products, reverse rate increases to make more reactants, $Q = K$, at equilibrium already.
56. Using ICE tables to find various concentrations at equilibrium, or to determine the rate constant.
57. Almost always we can use the 5% rule. If $K < 1$, and x < 5% of the starting concentrations.

Honors Unit 14 – Acid Base

58. $pH = -log[H^+]$, $14 = pH + pOH$, $[H^+] = 10^{-pH}$
59. Strong acids/bases dissociate completely. Weak acids/bases only dissociate a little bit.
60. Use ICE tables for weak acids/bases.
61. Acid + Base → H_2O + a salt. Some salts are acidic, basic, neutral.
62. When titrating, equivalence point is where the moles of acid = moles of base. Can use this to find concentration of an unknown acid or base.

Summer Assignment

63. Oxidation = loss of e- , Reduction = gain of e-
64. Gravimetric analysis is a clever way of using dimensional analysis and stoichiometry to find out how much of an unknown substance was originally present.

No question examples provided this time because you should have all your Honors Chem work as examples!!!

Red Packet