Summary of Basic Chemistry

Significant Figures	Tell you about the precision of a measurement	➢ All nonzero #s and zeros between non-zero #s are significant. 708 has 3 sig figs ➢ If a # is less than one, count all the #s <u>after</u> the first non-zero #. EX: 0.0009870 has 4 sig figs. ➢ If a # is greater than 1 and no decimal point is written, count only the non-zero #s and zeros in-between non-zero #s. Ex: 408000 has 3 sig figs. ➢ If a # is greater than one and a decimal point is written, count all #s. Ex: 9487.000 has 7 sig figs
	Adding and Subtracting (Round to the Significant figure farthest to the left of all of your givens)	➢ Add or subtract the numbers. Round your answer to the place of the least decimal places given Example: 15.00 cm + 4.352 cm = 19.352 = 19.35 cm; 5100 cm + 4.3 cm = 5104.3 cm = 5100 cm
	Multiplying and Dividing (Round to the least number of significant figures of all of your givens)	➢ Multiply or divide the numbers. Count the total number of significant figures in each number. Your final answer should have no more significant figures than the lowest number of significant figures you started with Example: 6.7 moles * 1.101 g/mol = 7.3767 = 7.4 g

Prefixes

Prefix	Symbol	Meaning	Conversion
giga	G	Billion (1 000 000 000 X)	10^9 b = 1 Gb
mega	M	Million (1 000 000 X)	10^6 b = 1 Mb
kilo	k	Thousand (1 000 X)	* 1000 b = 1 kb
hecto	H	Hundred (100 X)	100 b = 1 Hb
deca	D	Ten (10 X)	10 b = 1 Db
Base Unit			
deci	d	Tenth (1/10)	1 b = 10 db
centi	c	Hundredth (1/100)	* 1 b = 100 cb
milli	m	Thousandth (1/1000)	* 1 b = 1000 mb
micro	μ	Millionth (1/1 000 000)	* 1 b = 10^6 μb
nano	n	Billionth (1/1 000 000 000)	* 1 b = 10^9 nb
pico	p	Trillionth (1/1 000 000 000 000)	1 b = 10^{12} pb

Subatomic Particles

Subatomic Particle	Charge	Mass	Location	Formula
Proton (defines the type of atom)	+1	1	Nucleus	= atomic number
Neutron	0	1	Nucleus	= atomic mass – atomic number
Electron	-1	0	Orbitals around the nucleus	= atomic number – charge

Average Atomic Mass	Σ (Abundance * mass)	The relative abundances of Carbon-12 and carbon-13 are 98.9% and 1.19 % respectively. The average atomic mass is: (12*0.989) + (13*0.0119) = 12.02 amu
% Error	Measure of accuracy (how correct your data is)	$\% \text{ Error} = \frac{\lvert \text{accepted} - \text{experimental} \rvert}{\text{accepted}} \times 100$
Molar mass	Formula mass/weight, molecular mass/weight	➢ Add up the atomic masses of each atom. Pay attention to subscripts and parentheses. EX: $Cu(C_2H_3O_2)_2$ = (63.546 g/mol) + (4 * 12.011 g/mol) + (6 * 1.0079 g/mol) + (4 * 15.999 g/mol) = 181.633 g/mol
Empirical Formula	* Calc. moles of each atom * Divide by smallest # moles to find a mole ratio	EX: 79.8% C, 20.2% H 79.8 g (1 mol/12.011 g) = 6.64 mol C /6.64 mol = 1 C 20.2 g (1 mol/ 1.008 g) = 20.0 mol H /6.64 mol = 3 H So $\boxed{CH_3}$
Molecular Formula	$\frac{\text{Molecular MM}}{\text{Empirical MM}}$ = factor to mult. empirical by	The molecular mass of the above compound is 45 g. 45 ÷ (12.011 + (3*1.008)) = 3 so 3 (CH_3) = C_3H_9
% Composition	$\frac{\text{Mass of part}}{\text{Mass of total}} \times 100$	EX: % composition of Na_3P $\%Na = \frac{3(22.98)}{99.91} \times 100 = 69.0\%$ $\% P = \frac{30.97}{99.91} \times 100 = 31.0\%$
Mole Conversions	grams ←→ moles	Use the compound's Molar Mass (from the periodic table)
	particles ←→ moles (atoms, molecules, or formula units)	Use Avogadro's Number: 1 mole = 6.022 X 10^{23} particles for ANY substance
	liters ←→ moles	GASES ONLY, Use 1 mole = 22.4 L for any gas at STP
	Examples	? molec = 2.5 mol $H_2O \left(\frac{1\text{ mol } H_2O}{18.015\text{ g } H_2O}\right)\left(\frac{6.02 \times 10^{23}\text{ molec H O}}{1\text{ mol } H_2O}\right)$ = 8.4 X 10^{22} molec ? molec = 2.5 L $H_2O \left(\frac{1\text{ mol } H_2O}{22.4\text{ L } H_2O}\right)\left(\frac{6.02 \times 10^{23}\text{ molec H O}}{1\text{ mol } H_2O}\right)$
Electron Config.	Read off the Periodic Table	**s block** = Groups 1 & 2 **d block** = transition metals **p block** = Groups 13 – 18 **f block** = bottom two rows(inner trans.metals)

Formula Writing	Ionic formulas (also called salts) ➢ Made of positive cations and negative anions	Ionic Compounds have no charge, so add ***subscripts*** so that the total + & - charges are equal. TRICK: Criss-cross the charges for subscripts and reduce (make sure the subscripts are not divisible by anything but 1) EX: **Tin (IV) oxide** = Sn^{4+} O^{2-} , so you have Sn_2O_4 which equals **SnO_2**
	Molecular formulas ➢ Made of two nonmetals	Use prefixes (mono, di, tri, tetra, penta, hexa, hepta, octa, nona, deca) to determine subscripts. EX: carbon tetrachloride = CF_4
Naming Compounds	Ionic Compounds (contains a metal and/or a polyatomic ion)	Name the cation. Name the anion (change ending to –ide if not –ite or –ate) (If the cation can have more than 1 charge, determine charge and include as a roman numeral)
	Molecular Compounds (two nonmetals)	Use prefixes to denote subscripts (mono,di,tri,tetra,penta,hexa,hepta, octa,nona,deca)(do not start the 1st word with mono-) EX: CO = carbon monoxide ; N_4O_8 = tetranitrogen octoxide
Naming Acids	(**H^+** with and anion)	Anion ending: -ide = hydro(stem)ic acid EX: HCl = hydrochloric acid -ite = (stem)ous acid EX: HNO_2 = nitrous acid -ate = (stem)ic acid EX: H_3PO_4 = phosphoric acid
Balancing Equations	$Sn(ClO_3)_4 \rightarrow SnCl_4 + O_2$ $Sn(ClO_3)_4 \rightarrow SnCl_4 + \mathbf{6}O_2$	Remember **you can't change the compounds at all.** Use *coefficients* to get the same number of atoms of **each** element on each side of the equation
Types of Equations	Synthesis (combination)	$A + B \rightarrow AB$ EX: $H_2 + O_2 \rightarrow H_2O$
	Decomposition	$AB \rightarrow A + B$ EX: $CaO \rightarrow Ca + O_2$
	Single Replacement	$A + BC \rightarrow B + AC$ EX: $Li + CaSO_4 \rightarrow Ca + Li_2SO_4$
	Double Replacement	$AB + CD \rightarrow CB + AD$ EX: $CaCl_2 + NH_4NO_3 \rightarrow Ca(NO_3)_2 + NH_4Cl$
	Combustion	Carbon compound (C_xH_y) + $O_2 \rightarrow CO_2 + H_2O$
	Neutralization	Acid + base → salt + water (type of double replacement) $H_2SO_4 + NaOH \rightarrow Na_2SO_4 + HOH$
Molarity concentration	$M = \frac{mol}{L}$ & $M_1V_1 = M_2V_2$	Molarity = $\frac{\text{Moles of solute}}{\text{Liters of solution}}$
Gas Laws	The Ideal Gas Law	**$PV=nRT$** (R=0.0821 L*atm/mol*K or 8.31 L* kPa/Mol*K or 62.4 L*mmHg/mol*K)
	The Combined Ideal Gas Law	$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$
	Charles' Law	$V_1/T_1 = V_2/T_2$ (direct proportion) V↑T↑
	Boyle's Law	$P_1V_1 = P_2V_2$ (indirect proportion) P↓V↑
	Avogadro's Law	$V_1/n_1 = V_2/n_2$ (direct proportion) V↑n↑
	Gay-Lussac's Law	$P_1/T_1 = P_2/T_2$ (direct proportion) P↑T↑
	Dalton's Law	$P_1 + P_2 + P_3 + P_4 + \ldots = P_{total}$ (add up the partial pressure of each gas in the mixture) ONLY USED when there is more than one gas in the container
Stoichiometry	1. You Need a Balanced Equation 2. Convert given amount into moles if not already in moles. Remember: 1 mole = molar mass = 22.4 L (only gases at STP) = 6.02 x 10^{23} particles 3. Use coefficients from balanced equation to convert from moles of given to moles of answer. 4. If solving for moles, STOP. If not, convert answer into desired units using conversions from step 2.	
Limiting Reactant/Reagent (the reactant that you run out of)	find out which reactant makes the smaller amount of product using stoichiometry	
% Yield	$\frac{\text{Experimental Yield}}{\text{Theoretical Yield}}$ X 100	*Given experimental yield (amount of product made) *Calculate theoretical (how much product can you make?) using stoichiometry (use mass of reactant to find mass of product you should get) * Plug it into the equation
pH	$pH = -\log[H^+]$	$pH + pOH = 14$ $pOH = -\log[OH^-]$
Equilibrium Constant K_{eq}	$N_2\ (g) + O_2\ (g) \rightarrow 2NO(g)$	$K_{eq} = \frac{[\text{product}]^{coef}}{[\text{reactants}]^{coef}}$ EXAMPLE : $K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$
Heat Calories or Joules	Specific heat of H_2O = 1 cal/g°C 4.18 J = 1 cal	$\Delta T = \Delta T_{final} - \Delta T_{initial}$ Within a phase: Heat = (mass) x (specific heat) x (ΔT) = $mc\Delta T$ Changing phase: Heat = $m\Delta H$
Oxidation & Reduction		Oxidized = loss e-/incr. in oxidation number **LEO says GER** Reduction = gain e- /decr. in oxidation number EX: $Al + NaNO_3 \rightarrow Al(NO_3)_3 + Na$ Al is oxidized (red agent) 0 +1+5-2 +3+5-2 0 Na is reduced (ox agent)
	½ Reactions	Ox ½ = $Al^0 \rightarrow Al^{+3} + 3e^-$ Red ½ = $Na^{+1} + 1e^- \rightarrow Na^0$