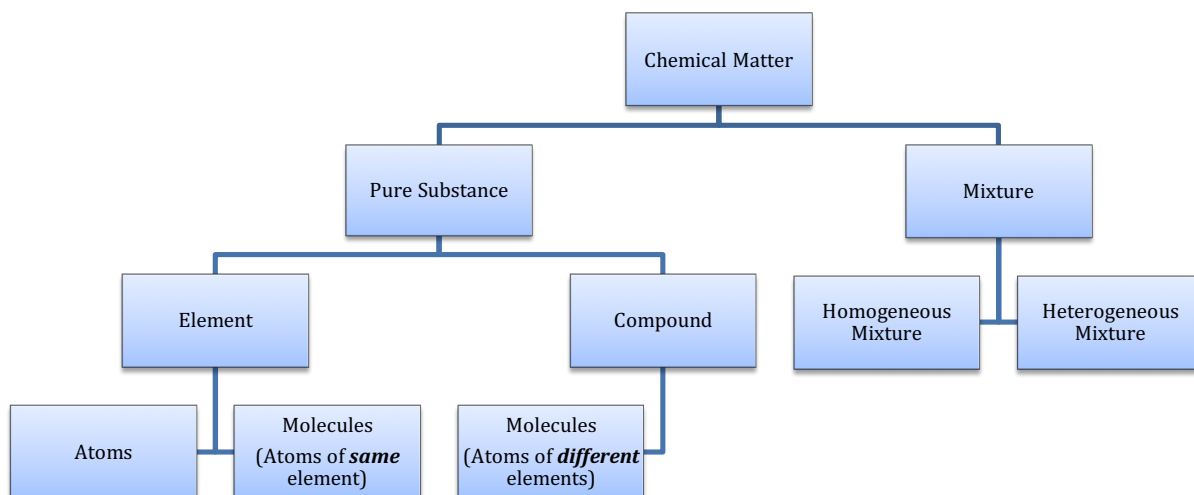


# College Chemistry First Semester Review Sheet

Fall, Dr. Wicks

## Chapter 1: Chemistry: The Study of Change

- I can explain how the subject of chemistry fits into science and into everyday life.
- I can explain the scientific method to someone not enrolled in AP Chemistry.
- I can understand the language used in the scientific method and I can distinguish between an experiment, data, an independent variable, and a dependent variable.
- I can distinguish between qualitative and quantitative results.
- I can distinguish between and give examples of a law, a hypothesis, and a theory.
- I can use a classification scheme for chemical matter.



- I can distinguish between and give examples of homogeneous mixtures and heterogeneous mixtures.
- I can distinguish between and give examples of elements, compounds, atoms, and molecules.
- I can explain the difference between chemical and physical changes and give examples of chemical and physical properties.
- I can distinguish between intensive and extensive properties.
- I can use metric-metric and English-metric conversion factors to solve problems.

Tera-	T	trillion	$10^{12} = 1,000,000,000,000$	1 inch (in.) = 2.54 cm
Giga-	G	billion	$10^9 = 1,000,000,000$	1 pound (lb.) = 454 g
Mega-	M	million	$10^6 = 1,000,000$	1 quart (qt.) = 0.946 L
Kilo-	k	thousand	$10^3 = 1,000$	
		one	$10^0 = 1$	1 mL = 1 cm <sup>3</sup> = 1 cc
Deci-	d	tenth	$10^{-1} = 0.1$	
Centi-	c	hundredth	$10^{-2} = 0.01$	pph = parts per hundred = %
Milli-	m	thousandth	$10^{-3} = 0.001$	ppm = parts per million
Micro-	μ	millionth	$10^{-6} = 0.000001$	ppb = parts per billion
Nano-	n	billionth	$10^{-9} = 0.000000001$	
Pico-	p	trillionth	$10^{-12} = 0.000000000001$	

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- I can convert between °F, °C, and K.

$$^{\circ}\text{C} = \left(\frac{5}{9}\right)(^{\circ}\text{F} - 32) \qquad ^{\circ}\text{F} = \left(\frac{9}{5}\right)^{\circ}\text{C} + 32 \qquad \text{K} = ^{\circ}\text{C} + 273.15$$

- I can explain the difference between precision and accuracy.
- I can apply the rules for using significant figures in calculations.
- I can use dimensional analysis for problem solving.
- I can use densities and percents as conversion factors in problem-solving.

### Chapter 2: Atoms, Molecules, and Ions

- I can describe how Dalton's atomic theory explained the law of conservation of mass, the law of definite proportions (law of constant composition), and the law of multiple proportions. See Table 1.

<b>Table 1: Laws Explained by Dalton's Atomic Theory</b>	
<i>Law</i>	<i>Meaning</i>
Law of Conservation of Mass	Matter can be neither created nor destroyed.
Law of Definite Proportions (Law of Constant Composition)	Different samples of the same compound always contain its constituent elements in the same proportions by mass.
Law of Multiple Proportions	If two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

- I can explain the historical development of atomic theory and identify some of the scientists who have made important contributions.
- I can explain the significance of Millikan's oil drop experiment and Rutherford's gold foil experiment.
- I can describe the structure of the atom using protons, neutrons, and electrons.
- Given atomic numbers and mass numbers, I can calculate the number of protons, neutrons, and electrons in atoms of given elements.
- I can explain what isotopes are and how isotopic abundance can be used to calculate the atomic mass of an element.
- Given a periodic table of the elements, I can identify the location of groups, periods, metals, nonmetals, metalloids (semimetals), alkali metals, alkaline-earth metals, halogens, noble gases, and transition metals.
- I can give examples of allotropes.
- I can use the periodic table to determine charges for ions of given elements.
- I know the names, chemical formulas, and charges for common polyatomic ions.

H <sub>3</sub> O <sup>+</sup>	Hydronium	HCO <sub>3</sub> <sup>-</sup>	Hydrogen Carbonate	ClO <sub>4</sub> <sup>-</sup>	Perchlorate
NH <sub>4</sub> <sup>+</sup>	Ammonium		(also called Bicarbonate)	ClO <sub>3</sub> <sup>-</sup>	Chlorate
OH <sup>-</sup>	Hydroxide	CO <sub>3</sub> <sup>2-</sup>	Carbonate	ClO <sub>2</sub> <sup>-</sup>	Chlorite
C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> <sup>-</sup>	Acetate	SO <sub>4</sub> <sup>2-</sup>	Sulfate	ClO <sup>-</sup>	Hypochlorite
CN <sup>-</sup>	Cyanide	PO <sub>4</sub> <sup>3-</sup>	Phosphate	MnO <sub>4</sub> <sup>-</sup>	Permanganate
NO <sub>3</sub> <sup>-</sup>	Nitrate	HPO <sub>4</sub> <sup>2-</sup>	Hydrogen Phosphate	CrO <sub>4</sub> <sup>2-</sup>	Chromate
NO <sub>2</sub> <sup>-</sup>	Nitrite	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	Dihydrogen Phosphate	Cr <sub>2</sub> O <sub>7</sub> <sup>2-</sup>	Dichromate

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- I can combine cations and anions to write formulas for ionic compounds.
- I can rapidly distinguish ionic compounds (metal and nonmetal elements) from molecular compounds (nonmetal elements only) for chemical nomenclature purposes.
- I can use the following prefixes to write the names for molecular compounds.

Mono- (1), di- (2), tri- (3), tetra- (4), penta- (5), hexa- (6), hepta- (7), octa- (8), nona- (9), deca- (10)

- I can write chemical names given chemical formulas and vice versa for ionic compounds, molecular compounds, acids, bases, and hydrates.

### Chapter 24: Organic Chemistry

- I can draw structures and name alkanes containing up to ten carbons. See Table 2.
- I can identify *cis*- and *trans*- isomers from the structures of simple alkenes.
- I can identify alkanes ( $C_nH_{2n+2}$ ), alkenes ( $C_nH_{2n}$ ), and alkynes ( $C_nH_{2n-2}$ ) from their structures and molecular formulas. (In a later chapter, we will compare and contrast the characteristics of single, double, and triple bonds in these compounds.)

<b>Table 2: Straight Chain Alkanes Containing 1 – 10 Carbons</b>		
<i>Molecular Formula</i>	<i>Expanded Molecular Formula</i>	<i>Name</i>
CH <sub>4</sub>	CH <sub>4</sub>	Methane
C <sub>2</sub> H <sub>6</sub>	CH <sub>3</sub> CH <sub>3</sub>	Ethane
C <sub>3</sub> H <sub>8</sub>	CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Propane
C <sub>4</sub> H <sub>10</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Butane
C <sub>5</sub> H <sub>12</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Pentane
C <sub>6</sub> H <sub>14</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Hexane
C <sub>7</sub> H <sub>16</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Heptane
C <sub>8</sub> H <sub>18</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Octane
C <sub>9</sub> H <sub>20</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Nonane
C <sub>10</sub> H <sub>22</sub>	CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>	Decane

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### Chapter 3: Mass Relationships in Chemical Reactions

- I can calculate an element's average atomic mass from the atomic mass and natural abundance data for multiple isotopes of the same element.
- I understand the chemical mole, and I can calculate the number of atoms or molecules present using Avogadro's number,  $6.022 \times 10^{23}$  particles/mole.
- I can calculate the molar mass for a chemical formula from the atomic masses on a periodic table of the elements.
- I can use Avogadro's number and molar masses as conversion factors to solve problems.
- I can express molecular composition in terms of percent composition. Remember it is helpful to assume you have one mole of a given compound during problem solving.

$$\% \text{ Composition of an Element in a Compound} = \left( \frac{\text{Mass of Element}}{\text{Mass of Compound}} \right) (100)$$

- I can solve empirical formula problems using the strategy outlined in Table 3.
- I can use percent composition to determine the empirical formula of a compound. Remember it is helpful to assume you have 100 g of a given compound during problem solving.
- I can use experimental data to calculate the number of water molecules in a hydrated compound. See Table 3.

<b>Table 3: Problem Solving Strategies for Empirical Formulas and Hydrate Formulas</b>	
<b><i>Empirical Formula Calculations</i></b>	<b><i>Formula of a Hydrate Calculations</i></b>
1. Get mass of each element	1. Get mass of water and anhydrous salt
2. Get moles	2. Get moles
3. Get mole ratio	3. Get mole ratio
4. Use whole number multiplier if needed	4. (Whole number multipliers are rarely needed)
5. Write the empirical formula	5. Write the formula of the hydrate

- I can obtain a molecular formula from an empirical formula using the molar masses of both. Recall that  $\left( \frac{MM_{\text{Molecular Formula}}}{MM_{\text{Empirical Formula}}} \right) = \text{Whole Number Multiplier}$  needed to obtain the molecular formula.
- I can balance simple chemical equations by inspection and complicated chemical equations by using the fraction method.
- I can balance chemical equations so that they have both material balance and charge balance.
- I can use reaction stoichiometry to interpret a chemical equation on a microscopic (molecular) level and a macroscopic (molar) level.
- I can calculate the mass (or moles) of one reactant or product from the mass (or moles) of another reactant or product in a balanced chemical equation.
- I can determine which reactant is the limiting reactant in a balanced chemical equation. I can also determine the amount of product formed and the amount of excess reactant leftover.
- I can distinguish between actual yield, theoretical yield, and percent yield.
- I can calculate theoretical yield and percent yield.  $\% \text{ Yield} = \left( \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \right) (100)$
- I can use stoichiometry principles to analyze a mixture or to find the empirical formula of an unknown compound.

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### Chapter 4: Reactions in Aqueous Solution

- I can explain the difference between a strong electrolyte, a weak electrolyte, and a nonelectrolyte.
- I can use the solubility rules to predict the solubility of ionic compounds in water. See Table 4.
- I can write the ions formed when an ionic compound dissolves in water.
- I can predict products for precipitation reactions (double replacement reactions).
- I can write molecular, ionic, and net ionic equations, and I can identify spectator ions.
- I can recognize common acids and bases.
- I can explain the difference between strong and weak acids and strong and weak bases.
- I can write equations for acid ionization and base dissociation.
- I can write molecular, ionic, and net ionic equations for acid-base neutralizations (double replacement reactions).
- I understand that when acid-base reactions form “salts,” this does NOT mean that they all form table salt, NaCl. The word “salt” in this context refers to an inorganic compound whose cation comes from a base and whose anion comes from an acid.

<b>Table 4: Solubility Rules for Common Ionic Compounds in Water at 25°C</b>	
<b><i>General Rule</i></b>	<b><i>Exceptions to the Rule</i></b>
Almost all compounds containing alkali metal ions ( $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , etc.) and $\text{NH}_4^+$ are soluble	
Almost all compounds containing nitrates ( $\text{NO}_3^-$ ), bicarbonates ( $\text{HCO}_3^-$ ), and chlorates ( $\text{ClO}_3^-$ ) are soluble	
Most compounds containing halides ( $\text{Cl}^-$ , $\text{Br}^-$ , and $\text{I}^-$ ) are soluble	Halides of $\text{Ag}^+$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$
Most sulfates ( $\text{SO}_4^{2-}$ ) are soluble	Sulfates of $\text{Ag}^+$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$ , $\text{Ba}^{2+}$ , $\text{Hg}_2^{2+}$ , and $\text{Pb}^{2+}$
Carbonates ( $\text{CO}_3^{2-}$ ), phosphates ( $\text{PO}_4^{3-}$ ), chromates ( $\text{CrO}_4^{2-}$ ), and sulfides ( $\text{S}^{2-}$ ) are usually insoluble	Compounds containing alkali metal ions ( $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , etc.) and $\text{NH}_4^+$
Hydroxides ( $\text{OH}^-$ ) are usually insoluble	Hydroxides containing alkali metal ions ( $\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , etc.) and $\text{Ba}^{2+}$

- I can identify Bronsted acids and Bronsted bases.
- I can predict the products of simple gas forming reactions.
- I can identify and write balanced chemical equations for four common types of reactions in aqueous solution:
  1. precipitation reactions - form a solid that does not dissolve in water (double replacement).
  2. acid-base reactions - form a “salt” and water (double replacement).
  3. gas-forming reactions - form a gas like  $\text{CO}_2$ .
  4. oxidation-reduction reactions - transfer electrons (combination, decomposition, combustion, and single replacement).
- I can use an activity series to predict whether a single replacement reaction will occur.

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- I can calculate oxidation numbers for elements in a compound.
- I can recognize “redox” reactions (oxidation-reduction reactions) and write simple half reactions.
- I can identify oxidizing agents and reducing agents. (Recall “OIL RIG,” which stands for “oxidation is loss” and “reduction is gain.”)
- I understand how molarity is used as a unit for concentration.
- I can calculate the molar concentration of ionic substances dissolved in water using

$$\text{Molarity} = \left( \frac{\text{Moles of Solute}}{\text{Liters of Solution}} \right)$$

- I can prepare solutions of various molar concentrations:
  1. from solids.
  2. by diluting liquids. (Recall  $C_1V_1 = C_2V_2$  where  $C$  = concentration and  $V$  = volume.)
- I can solve stoichiometry problems involving concentrations.
- I can perform a “titration” and a “standardization” experimentally and determine the unknown concentration for a reactant.

### Chapter 19: Electrochemistry

- I can balance redox equations using the ion-electron method.
- Here are the steps for balancing a redox equation occurring in acidic solution:
  1. Divide the equation into two half reactions
  2. Balance atoms other than H and “O”
  3. Balance “O” by adding water
  4. Balance H by adding hydrogen ion
  5. Balance the net charge by adding electrons
  6. Make electron gain = electron loss, and add the half-reactions
  7. Cancel anything the same on both sides of the final equation
- Here are the steps for balancing a redox equation occurring in basic solution:
  1. Divide the equation into two half reactions
  2. Balance atoms other than H and “O”
  3. Balance “O” and H by adding water and hydroxide ion
  4. Balance the net charge by adding electrons
  5. Make electron gain = electron loss, and add the half-reactions
  6. Cancel anything the same on both sides of the final equation

### Chapter 5: Gases

- I understand how pressure measurements are made.
- I can convert between various pressure units:  $1\text{ atm} = 760\text{ mm Hg} = 760\text{ torr} = 101.3\text{ kPa} = 14.7\text{ psi}$
- I know the standard conditions of temperature and pressure (STP) are  $0^\circ\text{C}$  (273 K) and 1 atm.
- I recall that all temperature-related gas laws use Kelvin rather than Celsius temperature.
- I understand that absolute zero on the Kelvin scale is the temperature at which all atomic and molecular motion stops. It can be determined by extrapolating a line to zero volume on a V-T diagram (Charles’ law) or by extrapolating a line to zero pressure on a P-T diagram (Gay-Lussac’s law).
- I can distinguish between Boyle’s law, Charles’ law, Gay-Lussac’s law, Avogadro’s law, the combined gas law, the ideal gas law, Dalton’s law, and Graham’s law, which are presented in Table 5 along with values for the gas law constant,  $R$ .

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- I can calculate gas densities and molar masses by substituting  $n = \frac{m}{MM}$  into the ideal gas law,  $PV = nRT$ , to get  $d = \frac{m}{V} = \frac{PMM}{RT}$  and  $MM = \frac{mRT}{PV} = \frac{dRT}{P}$ .
- I understand that the volume occupied by one mole of an ideal gas at standard conditions of temperature and pressure (STP, 0°C and 1 atm) is 22.4 L/mole, and I can use this conversion factor in problem solving.
- I can apply the gas laws when solving stoichiometry problems involving gaseous materials.
- I understand that the partial pressure of a gas in a mixture,  $P_A$ , is given by its mole fraction,  $X_A$ , multiplied by the total pressure of the mixture,  $P_{total}$ :  $P_A = X_A P_{total}$ . (The mole fraction of a component  $A$  in a mixture is defined as the number of moles of  $A$  divided by the total moles of all components.)
- I can apply the kinetic-molecular theory of gas behavior at the molecular level.
  - $KE = \frac{1}{2}mu^2$  where  $KE$  = average kinetic energy of the gas molecules,  $m$  = mass of the gas, and  $u^2$  = the average of the squares of the molecular speeds.
  - Since  $KE$  is determined by temperature, heavier molecules move with a slower average speed than lighter molecules at a given temperature.
  - For a given gas, the only way to change its average  $KE$  is to change its temperature.
- I can distinguish between ideal gases and real gases. Ideal gas molecules are assumed to have no volume and to not interact with one another by intermolecular forces. Deviations from ideal behavior are greatest at high pressure and low temperature.
- Ideal gas characteristics:
  - Gases are composed of very tiny molecules with lots of empty space between them.
  - Gas molecules move rapidly, move in straight lines, and travel in random directions.
  - Gas molecules do not attract each other.
  - Gas molecules have elastic collisions. In other words, the total kinetic energy of the molecules before and after a collision is the same.
  - The average kinetic energy of gas molecules is proportional to the absolute (Kelvin) temperature.
- In kinetic theory, the molecules of a gas are imagined to be a large number of points bouncing off the walls of a container. I understand that gas pressure is related to the number of collisions that occur with the walls per unit time.
- I can calculate the root mean square (rms) speed of molecules in a gas at Kelvin temperature  $T$  using  $u_{rms} = \sqrt{\frac{3RT}{MM}}$  where  $MM$  is the molar mass in kg/mole.
- I can compare and contrast diffusion and effusion and I know how to use Graham's Law. Graham's Law states that the rates of diffusion (or effusion) of two gases are inversely proportional to the square roots of their molar masses at the same temperature and pressure:  $\frac{r_1}{r_2} = \sqrt{\frac{MM_2}{MM_1}}$
- When studying the behavior of real gases, I can use the van der Waals equation:  $\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$  where  $a$  and  $b$  are van der Waals constants which correct for intermolecular forces and molecular volume.

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<b>Table 5: Gas Laws</b>		
<i>Gas Law Name</i>	<i>Equation</i>	<i>Quantities Held Constant</i>
<b>Boyle's Law:</b>	$P_1V_1 = P_2V_2$	Temperature and moles
<b>Charles' Law:</b> (Absolute zero, 0 K, can be determined by extrapolating a line to zero volume on a V-T diagram.)	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	Pressure and moles
<b>Gay-Lussac's Law:</b> (Absolute zero, 0 K, can be determined by extrapolating a line to zero pressure on a P-T diagram.)	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	Volume and moles
<b>Avogadro's Law:</b> (Stoichiometric coefficients can be used to compare volumes of different gases in a chemical equation,)	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$	Temperature and pressure
<b>Combined Gas Law:</b> (Combines Boyle's, Charles', and Gay-Lussac's Laws)	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	Moles
<b>Ideal Gas Law:</b> (Combines Boyle's, Charles', and Avogadro's Laws)	$PV = nRT$ where $R = 0.0821 \text{ L}\cdot\text{atm}/(\text{mole}\cdot\text{K})$ $= 8.314 \text{ J}/(\text{mole}\cdot\text{K})$ $= \text{Gas Law Constant}$	None
<b>Dalton's Law:</b> Total pressure is equal to the sum of the partial pressures.	$P_{Total} = \sum_{i=1}^n P_i$	Not applicable
<b>Graham's Law:</b>	$\frac{r_1}{r_2} = \sqrt{\frac{MM_2}{MM_1}}$	Not applicable