Unit 2 Summary Sheet

Types of Reactions

1. **Synthesis reaction**: occurs when two or more elements combine to form a compound.
   e.g. $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
   $A + B \rightarrow AB$
2. **Decomposition reaction**: occurs when a compound breaks down into two or more simpler substances.
   e.g. $2AgCl(s) \rightarrow 2Ag(s) + Cl_2(g)$
   $AB \rightarrow A + B$
3. **Single Displacement reaction**: occurs when an element replaces another element in a compound.
   e.g. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
   $AB + C \rightarrow AC + B$
4. **Double Displacement reaction**: occurs when the cations and anions displace each other to form two new products.
   e.g. $AgNO_3(aq) + NaCl(s) \rightarrow AgCl(s) + NaNO_3(aq)$
   $AB + CD \rightarrow AD + CB$
5. **Complete combustion reaction**: occurs when a hydrocarbon reacts with a sufficient amount of oxygen gas to produce carbon dioxide and water (blue flame).
   $C_xH_y(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
6. **Incomplete combustion reaction**: occurs when a hydrocarbon reacts with an insufficient amount of oxygen gas to produce carbon dioxide, water and carbon monoxide (orange flame).
   $C_xH_y(g) + O_2(g) \rightarrow CO(g) + CO_2(g) + H_2O(g)$

Balancing Chemical Equations

Law of Conservation of Mass: mass in an isolated system is neither created nor destroyed by chemical reactions or physical transformations.
Skeleton Equation: chemical equation in which the number of atoms is not equal on both sides.
Balanced Equation: number of atoms in the *reactants* must equal the number of atoms in the *products*.
e.g. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Activity Series

The Activity Series: used for predicting single displacement reactions.

- The element on the reactant side must be more "active" than the one it could replace in order for a single displacement reaction to occur.
  e.g. $Al(s) + CuCl_2(aq) \rightarrow Cu(s) + AlCl_3(aq)$
- If the element is not more "active" than the one it could replace, then the single displacement reaction will not occur.
  e.g. $Ni(s) + NaCl(aq) \rightarrow$ NR (no reaction)
- The reactivity of H_2O depends on the metal and water temperature.
- Near the top, the metals are more reactive because their valence electrons are easier to be removed.
- There is an activity series just for non-metals, this is much shorter.

Solubility Rules

Solubility Rules: used for predicting what products will be soluble or insoluble.

A **double displacement reaction** has occurred when there is:

- a formation of a precipitate (s),
- a formation of a gas (g),
- a formation of water (neutralization reaction).

Molecular Equation:
$KBr(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgBr(s)$

Complete Ionic Equation: all reactants and products separated into ions, contains spectator ions.
$K^+ + Br^- + Ag^+ + NO_3{}^- \rightarrow K^+ + NO_3{}^- + Ag^+ + Br^-$

Net Ionic Equation: eliminates the ions not directly involved in making the reaction happen.
$Br^- + Ag^+ \rightarrow AgBr(s)$

The Mole

Avogadro's constant (N_A): 6.022×10^{23}, this is how many particles are in one mole. The same number of atoms in 12 grams of carbon-12.
Mole: used to count atoms, ions, molecules, or formula units in groups of 6.022×10^{23}. A mole of pure carbon-12 is exactly 12 grams.
Molecule: basic unit of a covalent substance.
Formula Unit: basic unit of an ionic substance.

- n = number of moles
- N_A = Avogadro's constant
- N = number of particles
- $n = \dfrac{N}{N_A}$
- $N = n \times N_A$

e.g. How many gold atoms in 4.70×10^{-4} mol of gold?
$N = 4.70 \times 10^{-4}\, mol \times 6.022 \times 10^{23}\, atoms/mol$
$N \approx 2.83 \times 10^{20}\, Au\ Atoms$

Molar Mass

Molar mass (M): the mass in grams of one mole of a substance. It is numerically the same as the **atomic mass unit (amu)** of an element, except the units are in grams per mole (g/mol).
e.g. Atomic mass of Calcium = 40.078 amu
e.g. Molar mass of Calcium = 40.078 g/mol
Representation: A mole (6.022×10^{23} Ca atoms) of calcium atoms weigh 40.078 grams.

e.g. What is the molar mass of H_2O?
$M_{H_2O} = 2 \times 1.0079\, g/mol + 15.9994\, g/mol$
$M_{H_2O} \approx 18.02\, g/mol$

- n = number of moles (mol)
- M = molar mass (g/mol)
- m = mass (g)
- $M = \dfrac{m(g)}{n(mol)}$, $n = \dfrac{m}{M}$, $m = n \times M$

Percent Composition

Law of Definite Proportions: the proportions of each element in a chemical compound, regardless of quantity, are the same.
Percent Composition: the percent by mass of each element in a compound.

$$\%\, mass\, of\, element = \frac{mass\, of\, element}{total\, mass\, of\, compound} \times 100\%$$

e.g. What is the percent composition of CO_2?
$M_{CO_2} = 44.01\, g/mol$
$\%C = \dfrac{12.011 g/mol}{44.01\, g/mol} \times 100\%$
$\%C = 27.3\,\%$

$\%O = \dfrac{32.00 g/mol}{44.01\, g/mol} \times 100\%$
$\%O = 72.7\,\%$

Empirical vs. Molecular Formula

Empirical Formula: shows the lowest whole number ratio of elements in a compound (simplest formula of a compound). Some compounds can have the same empirical formula.
Molecular Formula: shows the exact number of atoms of each element (actual formula of a compound). Each molecular formula is unique.

Molecular Formula: $C_6H_{12}O_6$
Empirical Formula: CH_2O

Determining the Empirical Formula:

1. Assume 100.0 g of the compound, change percentages of each of the elements to grams.
2. Convert the grams of each elements to moles.
3. Divide by the smallest number of moles.
4. Multiply by a constant to get a whole number.

Determining the Molecular Formula:

$$ratio = \frac{mass\, of\, molecular\, formula}{mass\, of\, empirical\, formula}$$

Multiply each element by this ratio in order to determine the molecular formula.

Molecular Formula of a Hydrate

Hydrate: ionic compound that contains water molecules in its structure.
Anhydrate: the substance that remains after the water has been removed from the hydrate.
e.g. What is the formula of $MgSO_4 \cdot XH_2O$?

	$MgSO_4$ Hydrate	13.52 g
-	Anhydrate	6.60 g
	Water	6.92 g

$n_{MgSO_4} = \dfrac{6.60\, g}{120.37\, g/mol}$
$n_{MgSO_4} = 0.054831\, mol$

$n_{H_2O} = \dfrac{6.92\, g}{18.02\, g/mol}$
$n_{H_2O} = 0.38402\, mol$

$\therefore$ dividing n_{H_2O} by n_{MgSO_4} to determine the coefficient of H_2O, the hydrate is $MgSO_4 \cdot 7H_2O$.

Stoichiometry

Mole-to-Mole Calculations: mol A → mol B

1. Write the balanced chemical equation.
2. Write the mole ratios for the given substance(s).
3. Use the mole ratio to determine what amount of desired substance is produced or needed.

Mass-to-Mass Calculations: grams A → grams B

1. Write the balanced chemical equation.
2. Use molar mass to convert mass to moles.
3. Use mole ratio to find the amount in moles.
4. Convert moles of desired substance to mass.

Mass-to-Particle Calculations: grams A → mol B

1. Write the balanced chemical equation.
2. Use molar mass to convert the given mass to moles.
3. Apply the mole ratio to determine what amount of desired substance is produced or needed.
4. Multiply by avogradro's constant to determine the number of particles.

Limiting Reactant

Limiting Reagent: the first reactant to be used up in a reaction; determines when the reaction stops.
Excess Reagent: the reactant(s) that are not used up when the reaction is finished; it is left over.
Stoichiometry involving Limiting Reactant: identifying which reactant will run out first.

1. Write the balanced chemical equation.
2. Convert each reactant to moles.
3. Determine the amount of desired product that can be produced using the mole ratio.
4. The limiting reactant is the one that produces the least amount of product.
5. Using the amount of product based on the limiting reactant, convert the moles of the desired substance into mass.

Percent Yield

Theoretical Yield: calculated amount of product based on the stoichiometry of the reaction.
Actual Yield: amount of product collected during the experiment.
Usually, the theoretical yield > actual yield.
Percent Yield: the percent ratio of the actual yield to the theoretical yield.

$$percent\, yield = \frac{actual\, yield}{theoretical\, yield} \times 100\%$$

Factors affecting percentage yield:

- Reversible reactions.
- Not enough time to complete the reaction.
- Reactants contain impurities.

Too high of a temperature does not affect percent yield.

Percent Purity

Percent Purity: the percent ratio of the pure product to the impure product obtained by mass.

$$\%\, purity = \frac{mass\, of\, pure\, product}{mass\, of\, impure\, product\, obtained} \times 100\%$$