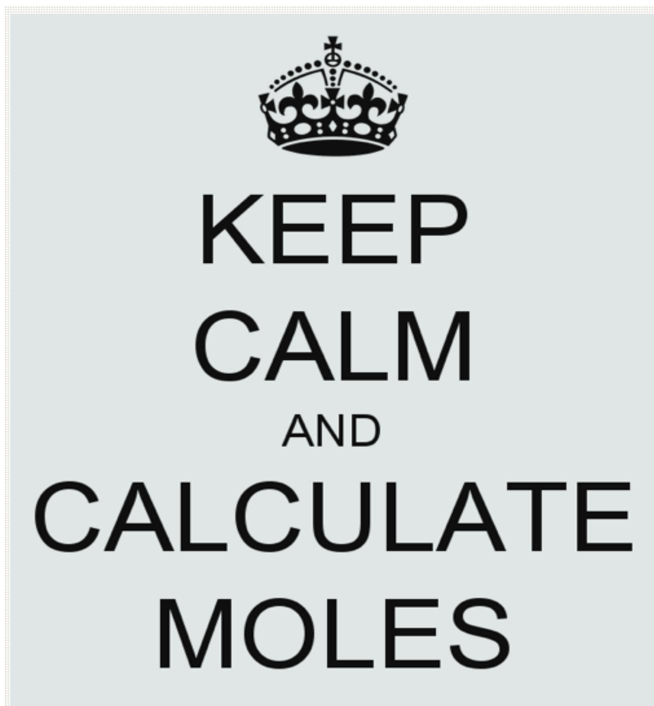
Regents Chemistry:

Notes: Unit 7 Moles & Stoichiometry



KEY IDEAS

- In all chemical reactions there is a conservation of mass, energy, and charge. (3.3a)
- A balanced chemical equation represents conservation of atoms. The coefficients in a balanced chemical equation can be used to determine mole ratios in the reaction. (3.3c)
- The formula mass of a substance is the sum of the atomic masses of its atoms. The molar mass (gram formula mass) of a substance equals one mole of that substance. (3.3e)
- The empirical formula of a compound is the simplest whole-number ratio of atoms of the elements in a compound. It may be different from the molecular formula, which is the actual ratio of atoms in a molecule of that compound. (3.3d)
- The percent composition by mass of each element in a compound can be calculated mathematically. (3.3f)
- Types of chemical reactions include synthesis, decomposition, single replacement, and double replacement. (3.2b)

PROCESS SKILLS

- Calculate the formula mass and gram-formula mass of a substance (3.3viii)
- Determine the mass of a given number of moles of a substance (3.3 vii)
- Determine the empirical formula from a molecular formula (3.3v)
- Determine the molecular formula, given the empirical formula and the molecular mass (3.3vii)
- Identify types of chemical reactions (3.2ii)
- Balance equations, given the formulas of reactants and products (3.2v, 3.3i)
- Calculate simple mole-mole stoichiometry problems, given a balanced equation (3.3iv)
- Determine a missing reactant or product in a balanced equation (3.2iii)
- Interpret balanced chemical equations in terms of conservation of matter and energy (3.3ii)
- Create and use models of particles to demonstrate balanced equations (3.3iii)
- Calculate simple mole-mole stoichiometry problems, given a balanced equation (3.3iv)

Word	Definition
Balanced Equation	An expression of a chemical reaction where mass and charge are conserved; the number and type of atoms in the reactants being equal to the number and type of atoms in the products and the total charge of the reactants equal to the total charge of the products.
Coefficient	A number placed in front of a formula to balance a chemical reaction.
Decomposition	A redox reaction in which a compound breaks up to form two elements.
Double replacement	A reaction in which the positive ion of one compound combines with the negative ion of the other compound; usually occurs in solution where a precipitate is formed with other ions remaining dissolved in solution.
Empirical Formula	Formula for a compound which provides the simplest ratio of the elements present.
Formula Mass	The sum of the atomic masses of a substance in a.m.u.
Gram Formula Mass	The sum of the atomic masses of a substance in grams; also called molar mass.
Law of conservation of charge	Charge cannot be created or destroyed by physical or chemical change. This is the basis for writing chemical formulas and half-reactions, and balancing redox ionic reactions.
Law of conservation of energy	Energy cannot be created or destroyed by physical or chemical change. This is the basis for calculating the heat of reaction.
Law of conservation of mass	Matter cannot be created or destroyed by physical or chemical change. This is the basis for balancing chemical reactions.
Mole	A quantity of 6.02×10^{23} units of a substance; the amount of a substance equal to the sum of the atomic masses in grams; Avogadro's number.
Mole ratio	The whole-number ratio between components of a balanced chemical reaction.
Molecular Formula	Formula for a compound which provides the number and identity of the atoms of each element present.

Vocabulary

Oxidation	The loss of electron(s), causing the oxidation number of a species to become more positive.
Precipitate	An insoluble solid that is formed either in a double-replacement reaction or as excess solute added to a saturated solution.
Product	The substances that are formed by a chemical reaction, designated as the right side of a chemical equation.
Reactant	The substances that are reacted together, designated as the left side of a chemical equation.
Reaction	A chemical change where reactants are turned into products.
Redox reaction	A reaction in which one element is oxidized and another element is reduced.
Reduction	The gain of electron(s), causing the oxidation number of a species to become more negative.
Single replacement	A redox reaction in which an element replaces an ion in a compound.
Species	The individual products and reactants in a chemical reaction.
Spectator ion	An ion that does not participate in the chemical reaction. In a redox reaction, it is the ion whose charge does not change. In a double replacement reaction, they are the ions that remain dissolved in solution.
Stoichiometry	The mathematics of mole relationships.
Synthesis	A redox reaction in which two elements combine to form a compound.

Lesson 1: Moles and Molar Mass

Objective:
- ***Determine the number of atoms in a molecule (formula)***
- ***Calculate Molar Mass (gram formula mass)***

Stoichiometry: An accounting of atoms in a chemical reaction.

REVIEW OF COUNTING ATOMS

Subscripts: number of atoms in a formula
Coefficients: TOTAL number of molecules or compounds

Ex. $Ca_3(PO_4)_2$

Ex. $3Ca_3(PO_4)_2$

PRACTICE: How many oxygen atoms in each?
1. NH_4NO_3
2. $Al_2(SO_4)_3$
3. $3Al_2(SO_4)_3$

The MOLE (mol):
- Unit that measures the **amount** (number of particles) of a substance
- 1 mol = 6.02×10^{23} particles *(things)*
- 1 mol = gram formula mass

Lesson 1: Moles and Molar Mass

Why we use it: atoms are sooooooo tiny that we have LOTS of them in a given sample.
1 dozen eggs = 12 eggs
1 mol of eggs = 6.02×10^{23} eggs

Example:
3.5 dozen roses = 3.5 (12) = 42 roses
3.5 mol roses = 3.5 (6.02×10^{23}) = 2.1×10^{24} roses

<u>MOLAR MASS</u>: (aka **Gram Formula Mass** or **Molecular Mass)**

**Element's molar masses are equal to the atomic mass as reported on the periodic table, only in grams/mole as opposed to amu/atom

<u>Calculating Molar Mass (Gram Formula Mass GFM)</u>

1. Count the atoms in the compound
2. Look up mass of each atom on periodic table
3. Multiply mass by the # of atoms of each element
4. Add up the total mass

Molar Mass Examples: Elements

1. What is the molar mass of iron?

2. What is the molar mass of copper?

Molar Mass Examples: Compounds

1. What is the molar mass of water?

H_2O

H: 2 x =

O: 1 x = + ____________

GFM for H2O: ______________

2. What is the gram-formula-mass of calcium chloride?

IN CLASS EXAMPLE:

$Al_2(SO_4)_3$

Lesson 2: Calculating Percent Composition

Objective:

- ***Calculate Percent Composition of an element in a compound***
- ***Calculate Percent composition of a hydrate***

PERCENT COMPOSITION:

The percentage by mass of each element in a compound

Percent Composition	$\% \text{ composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100$

***Formula located on Table T

<u>CALCULATING PERCENT COMPOSITION</u>

1. Calculate GFM of compound (mass of whole)
2. Plug into % composition formula

EXAMPLE: What is the percent composition of Calcium in $CaCl_2$?

GFM $CaCl_2$: Ca = 40.08g x 1 = 40.08g
Cl = 35.453g x 2 = <u>70.906g</u>
= 110.99 g/mol

$$\boxed{\% \text{ composition} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100}$$

$$= \frac{40.08\text{g}}{110.99\text{g/mol}} \times 100 = 36.11 \% \text{ Ca}$$

EXAMPLE: What is the percentage by mass of carbon in CO_2?

EXAMPLE: What is the percent by mass of nitrogen in NH_4NO_3?

HYDRATES

Hydrate- Ionic solids with water trapped in the crystal lattice.

It is written like this:

Ionic Compound's Formula • n H_2O

(n) is a whole number

Ex. $CuSO_4 \cdot 5H_2O$

The dot shows that **5 H_2O molecules are attached to 1 $CuSO_4$ molecule**.

CALCULATING PERCENT COMPOSITION OF A HYDRATE

1. Calculate GFM of HYDRATE including the water (mass of whole)
2. Plug into % composition formula T

EXAMPLE: What is the percentage by mass of water in sodium carbonate crystals ($Na_2CO_3 \cdot 10H_2O$)?

Step 1- Calculate GFM of Hydrate

Na =	2 x 22.98977	= 45.97954
C=	1 x 12.0111	= 12.0111
O=	3 x 15.9994	= 47.9982
H_2O =	10 x 18.01528	= 180.1528
	Gram Formula Mass of Hydrate	= 286.1 g/mol

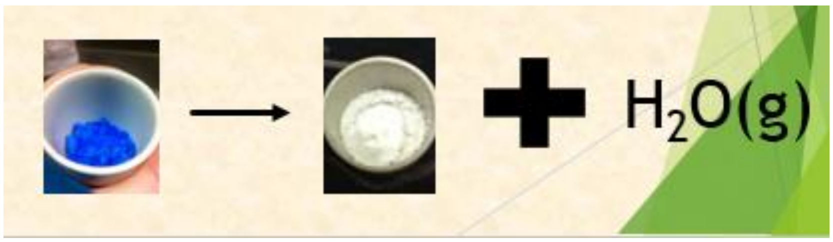
Step 2- Plug values into Formula

$$\% H_2O_{\text{by mass}} = \frac{\overbrace{180.2g}^{\text{Mass of } H_2O}}{\underbrace{286.1g}_{\text{Mass of hydrated compound}}} \times 100 = 63.0\ \%$$

EXAMPLE: What is the percent by mass of water in $BaCl_2 \cdot 2H_2O$?

Calculating Percent Composition of Water in a Hydrate using Experimental Data:

- **If you don't know the formula, you can still determine the percent of water in a hydrate**
- **Heat the hydrate to evaporate the water**
  - **Starting mass is mass of whole**
  - **Subtract ending mass (part without water) to get mass of water part.**



EXAMPLE:

A 12.2g sample of a hydrate was heated to a constant mass of 10.2 grams. What is the percent by mass of the water in the hydrate?

Water is heated (evaporated) off so the decrease in mass is the mass of the water

12.2g – 10.2g = $\dfrac{\text{2.0g (mass of water)}}{\text{12.2g (mass of whole)}}$ x 100 = 16 % water in the hydrate

Objective:

- ***Calculate the number of moles given the grams***
- ***Calculate the number of grams given the moles***

Use MOLE FORMULA from Table T

Mole Calculations	$\text{number of moles} = \dfrac{\text{given mass}}{\text{gram-formula mass}}$

***Given mass is the <u>mass in grams</u>

How to Convert from Grams to Moles

1. Calculate the GFM for the compound. (round to nearest tenth)
2. Plug the given value and the GFM into the "mole calculations" formula and solve for the number of moles.

EXAMPLE 1: How many moles are in 39.0 grams of LiF?

$$\begin{array}{r} Li = 6.9g \\ \underline{+\ F = 19.0g} \\ = 25.9g/mol \end{array}$$

$$mol = \frac{39.0\cancel{g}}{25.9\cancel{g}/mol} = 1.51\ mol$$

EXAMPLE 2: What is the number of moles of potassium chloride present in 148 g?

How to Convert from Moles to Grams:

Use the same mole formula to solve for grams by setting mole over 1 and cross multiplying

CHECK:

- If more than 1 mol, mass will be **GREATER** than GFM
- If less than 1 mol, mass will be **SMALLER** than GFM

EXAMPLE: What is the mass of 4.5 moles of KOH?

K= 39.1g
O= 16.0 g
+ H= 1.0 g
= 56.1 g/mol

$$\frac{4.5\,\text{mol}}{1} = \frac{x}{56.1\ \text{g/mol}}$$

$= 252.5\ \text{g}$

$\boxed{= 250\ \text{g}}$

EXAMPLE: What is the mass of 0.50 mol of $CuSO_4$?

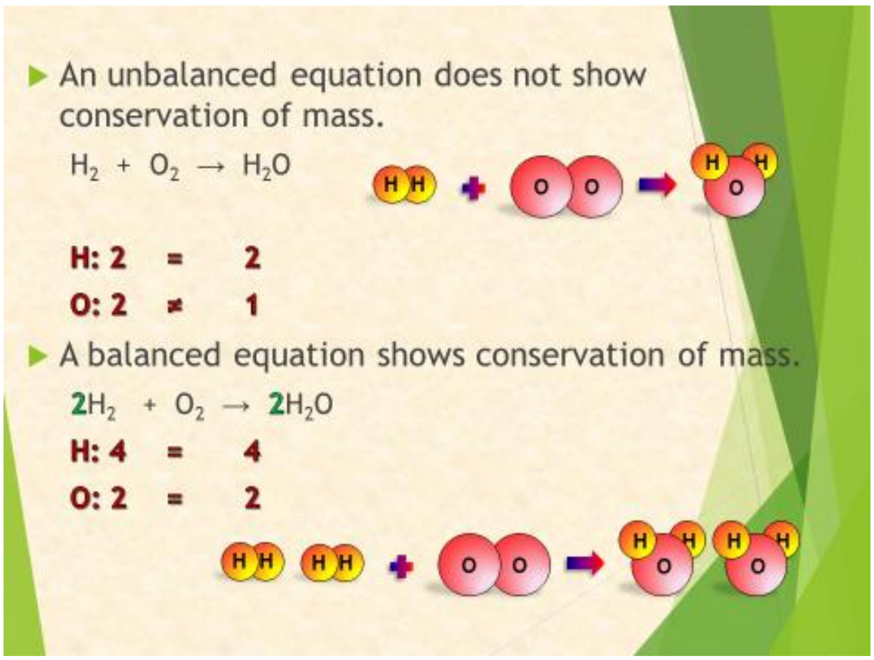
Objective:

- ***Assess and Balance chemical reactions using coefficients***

CONSERVATION OF MASS

- In all chemical reactions there is a CONSERVATION of mass, energy, and charge "what goes in must come out"
- The number of ATOMS of each element must be EQUAL on BOTH SIDES of the equation.

***Matter and energy cannot be created nor destroyed, only changed from one form to another*



$H_2 + O_2$ = 2g + 32g = 34g

≠

H_2O = 18g

BUT

$2H_2 + O_2$ = 4g + 32g = 36g

=

$2H_2O$ = 2x18g = 36g

BALANCING EQUATIONS

Balance with **COEFFICIENTS**
**NOTE: WE NEVER CHANGE THE SUBSCRIPTS IN A FORMULA!

STEPS FOR BALANCING EQUATIONS

1. Draw a line
2. Count the number of each type of atom in the reactants and the products
3. Keep polyatomic ions **(TABLE E)** together. Count as a unit if not broken up.
4. Start with the element that is only found once on both sides.
5. Keep changing coefficients until all atoms match
6. Coefficients must be smallest possible whole number (reduce if not)
7. Finally check each atom to see if its balanced

EXAMPLE: ___Al + ___$CuCl_2 \rightarrow$ ___Cu + ___ $AlCl_3$

EXAMPLE: ___N_2 + ___$H_2 \rightarrow$ ___NH_3

EXAMPLE: ___H_2 + ___$O_2 \rightarrow$ ___ H_2O

EXAMPLE: **__$BaCl_2$ + __$AgNO_3 \rightarrow$ __$Ba(NO_3)_2$ + __AgCl**

Lesson 5: Types of Chemical Reactions

Objective:

- **_Identify Types of Reactions_**

TYPES OF REACTIONS

1. Combustion: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

   Burning of Hydrocarbons in presence of Oxygen

Table I
Heats of Reaction at 101.3 kPa and 298 K

Reaction	ΔH (kJ)*
$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$	–890.4
$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(\ell)$	–2219.2
$2C_8H_{18}(\ell) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(\ell)$	–10943
$2CH_3OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(\ell)$	–1452
$C_2H_5OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(\ell)$	–1367
$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(\ell)$	–2804
$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$	–566.0
$C(s) + O_2(g) \longrightarrow CO_2(g)$	–393.5
$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$	–3351
$N_2(g) + O_2(g) \longrightarrow 2NO(g)$	+182.6
$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g)$	+66.4
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$	–483.6
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(\ell)$	–571.6
$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$	–91.8
$2C(s) + 3H_2(g) \longrightarrow C_2H_6(g)$	–84.0
$2C(s) + 2H_2(g) \longrightarrow C_2H_4(g)$	+52.4
$2C(s) + H_2(g) \longrightarrow C_2H_2(g)$	+227.4
$H_2(g) + I_2(g) \longrightarrow 2HI(g)$	+53.0
$KNO_3(s) \xrightarrow{H_2O} K^+(aq) + NO_3^-(aq)$	+34.89

2. Synthesis: $A + 2B \rightarrow AB_2$

   Making one compound from two substances…

3. Decomposition: $AB_2 \rightarrow A + 2B$

   Chemically separating a compound:

   The REVERSE of Synthesis

4. Single Replacement: $AB + C \rightarrow CB + A$

   When an element replaces an ion in an ionic compound

   Examples:

5. Double Replacement: $AB + CD \rightarrow AD + CB$

   When two ionic compounds "switch" partners. When done in solution, result is a PRECIPITATE:

   Examples:

PRACTICE: *Identify the type of reaction*

1.) $4Li + O_2 \rightarrow 2\ LI_2O$

2.) $Al + CuCl_2 \rightarrow AlCl_3 + Cu$

3.) $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow 2NaCl_{(aq)} + BaSO_{4\,(s)}$

4.) $2AuBr_3 \rightarrow 2Au + 3Br_2$

Lesson 6: Mole Ratios

Objective:

- ***Calculate mole ratios in a chemical formula***

Mix + 2 Eggs + 1 cup water = Cake

Reactants Products

Just like a recipe in a chemical Rxn you can double, triple, halve etc. the amounts of the ingredients to change how much product you make

Example:
2 Mix = 2 Cakes
6 Eggs = 3 Cakes

COEFFICIENTS: How many MOLES of the substance are needed in a reaction.

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

reactants products

CALCULATING MOLE RATIOS:

1. Circle the substances involved (one will be given, the other the unknown)
2. Set up a ratio of moles (proportion) of SUBSTANCES in the balanced equation to the ACTUAL MOLE values and solve for the unknown – do this using the equation directly.

EXAMPLE: How many moles of oxygen are consumed when 0.6 moles of hydrogen burns to produce water?

known ***unknown***
0.6 ***x***

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O$$

$$\frac{0.6}{2} = \frac{x}{1} \qquad \mathbf{x = 0.3\ mol}$$

Remember when no coefficient is written, the coefficient is 1!

EXAMPLE: How many moles of nitrogen gas (N_2) would be needed to produce 10 moles of ammonia (NH_3) in the following reaction?

$$\underline{x} \qquad\qquad \underline{10}$$

$$N_2 + 3H_2 \rightarrow 2NH_3$$

$$\frac{x}{1} = \frac{10}{2}$$

$$2x = 10$$

$$X = 5 \text{ mol}$$

EXAMPLE: If 12 moles of C_3H_8 react completely, how many moles of H_2O are formed in the reaction below?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

PRACTICE: If 20 moles of CO_2 are formed, how many moles of O_2 reacted in the reaction below?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

Lesson 7: Determining empirical and molecular formulas

Objective:

- ***Determine the empirical formula from the molecular formula***
- ***Determine the molecular formula from the empirical formula***

EMPIRICAL FORMULA: Smallest whole number ratio of atoms in a compound (cannot be reduced any further)

Example: $B_{10}G_{20}$
B_1G_2 (empirical formula)

Example: C_2H_6
(empirical formula)

MOLECULAR FORMULA: "Actual Formula" - the actual number of atoms in a compound

- Whole # multiple of empirical
- If the empirical formula is CH_4 a molecular formula could be CH_4, C_2H_8, C_3H_{12} etc.

Example: B_1G_2 (empirical formula)
$B_{10}G_{20}$ (molecular formula)

Example: CH_3 (empirical formula)
(molecular formula)

More Examples:

Molecular Formula	Empirical Formula
N_2O_4	NO_2
C_3H_9	CH_3
$C_6H_{12}O_6$	CH_2O
B_4H_{10}	B_2H_5
C_5H_{12}	C_5H_{12}

DETERMINE THE EMPIRICAL FORMULA

Divide subscripts by the greatest common factor

Example: molecular formula = C_4H_{10}
Divide by 2 *(greatest common factor)*
Answer: C_2H_5

EXAMPLE: Determine empirical formula from molecular formula.

1. $C_6H_{12}O_6$
2. N_2O_4
3. $BaCl_2$
4. C_2H_6
5. CH_3

CALCULATING MOLECULAR FORMULA FROM EMPIRICAL FORMULA

1. Calculate Gram Formula Mass of the **EMPIRICAL FORMULA**
2. Divide molecular mass given by the empirical formula mass
3. Multiply all of the subscripts in the empirical formula by the number (multiple) you calculated in step 2

EXAMPLE:

The empirical formula for ethylene is CH_2. Find the molecular formula if the molecular mass is 28.1 g/mol?

empirical mass CH_2 =

(1 C X 12.0 g/mol) + (2 H X 1.0 g/mol) = 14.0 g/mol

$$\frac{28.1 \text{ g/mol}}{14.0 \text{ g/mol}} = 2.00$$

$$(CH_2) \times 2 \Rightarrow C_2H_4$$

EXAMPLE: What is the molecular formula of a compound that has an empirical formula of NO_2 and molecular mass of 92.0 g?

N= 14.0 x 1 = 14.0g

O= 16.0 x 2 = 32.0g

\+ = 46.0g/mol

$$\frac{92.0 \text{ g/mol}}{46.0 \text{ g/mol}} = 2.00$$

$$(NO_2) \times 2 \Rightarrow N_2O_4$$

EXAMPLE: A compound has an empirical formula of HCO_2 and a molecular mass of 90 grams per mole. What is the molecular formula of this compound?

H= 1.0 x 1 = 1.0g

C= 12.0 x 1 = 12.0g

O= 16.0 x 2 = 32.0g

\+ = 45.0g/mol

$$\frac{90.0 \text{ g/mol}}{45.0 \text{ g/mol}} = 2.00$$

$$(HCO_2) \times 2 \Rightarrow H_2C_2O_4$$

Practice Space