Ms. Randall
Regents Chemistry
Unit 6: Moles and Chemical Reactions Notes

Unit Objectives:

- Understand the mole concept
- Calculate gram formula mass
- Differentiate between formula mass and gram formula mass
- Convert between grams and moles
- Balance a chemical reaction by adjusting only the coefficients
- State the Law of Conservation of Mass and Energy and relate it to balanced chemical equations
- Create and use models of particles to demonstrate balanced equations
- Identify various types of reactions: synthesis, decomposition, single replacement, & double replacement
- Solve mole-mole Stoichiometry problems given a balanced reaction
- Calculate the empirical formula from percent mass
- Differentiate between empirical and molecular formulas
- Determine the molecular formula from the empirical formula and molecular mass

Define the following vocabulary:

Mole
Formula mass (FM)
Gram formula mass (GFM)
Coefficient
Subscript Species
Law of conservation of mass and energy
Balanced equation
Synthesis reaction
Decomposition reaction
Single-replacement reaction
Double-replacement reaction
Molecular formula
Empirical formula
Percent mass

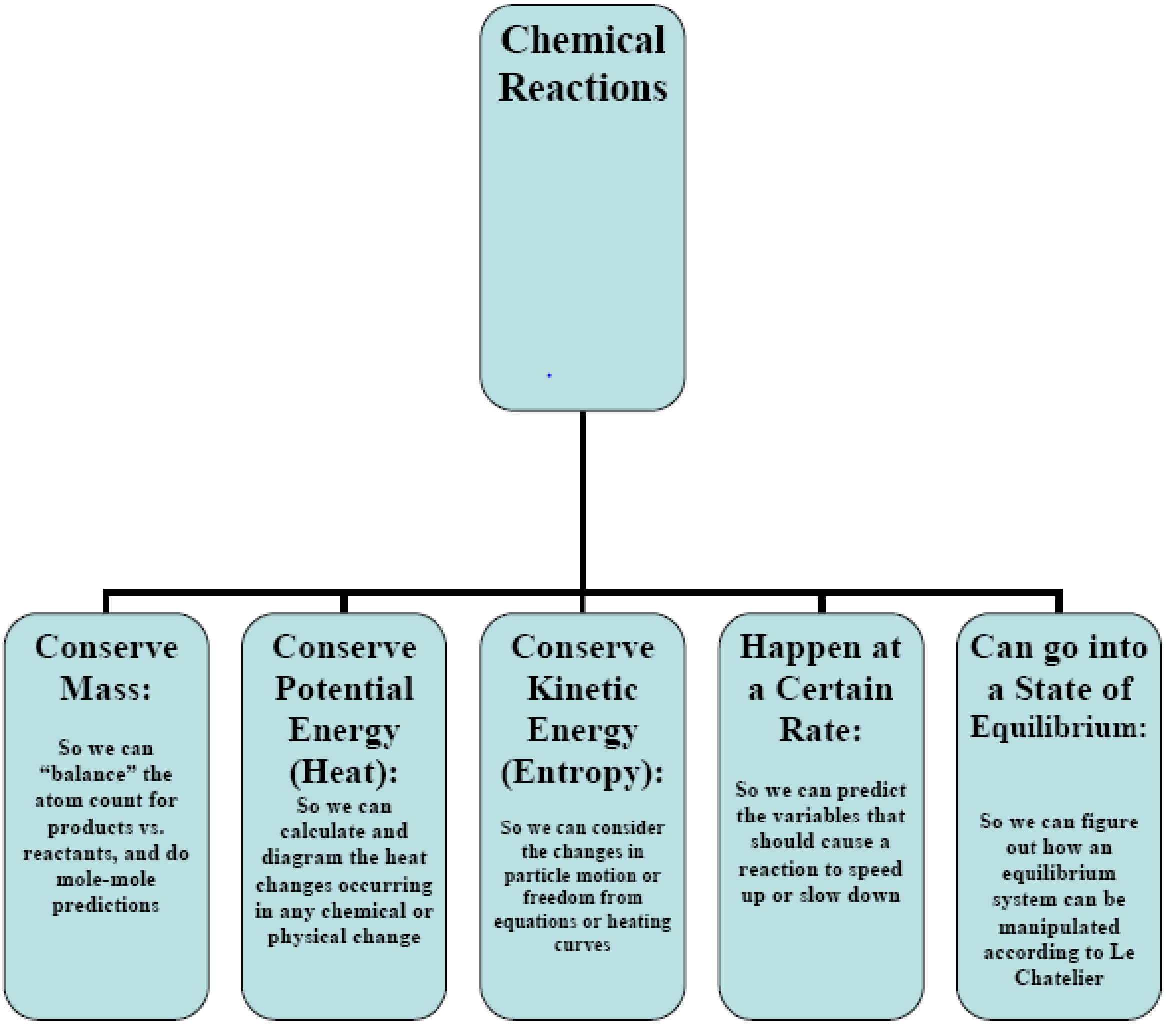
Lesson 1: Chapter Diary 7

Objective: To summarize the concepts related to Moles and Stoichiometry.

Directions: After reading the Chapter diary 7 answer the questions in your unit workbook.

Lesson 2: Chemistry and the Mole

Objective: To define and calculate molar mass. To apply the formula relating mass in grams to moles



Chemistry is a basic science whose central concerns are -

- the structure and behavior of atoms (elements)
- the composition and properties of compounds
- the reactions between substances with their accompanying energy exchange
- The laws that unite these phenomena into a comprehensive system.

We have already studied the structure and behavior of atoms (elements) as well as the composition and properties of compounds. We will now be moving on to learn about the reactions that occur between substances. In order to do this we must be able to quantify the reactants and products to see the changes that are occurring.

<u>What is a Mole?</u> = 6.02×10^{23} things

- Symbol: **mol**
- Like **a dozen, only waaaaaaaaaaay bigger**
- Why we use it: **atoms are soooooo tiny that we have LOTS of them in a given sample**

1 dozen eggs = **12 eggs**
1 mol eggs = **6.02×10^{23} eggs**

Both subscripts and coefficients represent MOLES

Ex: in *1 mole* of NaCl, there are 1 mol of Na^+ ions
1 mol of Cl^- ions

in *1 mole* of H_2O, there are 2 mol of H atoms
1 mol of O atoms

in *1 mole* of Na_2SO_4, there are 2 mol of Na^+ *ions*
1 mol of SO_4^{2-} *ions*

How many moles of <u>atoms</u> in total? **<u>7</u>**

in *1 mole* of $(NH_4)_2CO_3$, there are 2 mol of NH_4^+ *ions*
1 mol of CO_3^{2-} *ions*

How many moles of <u>atoms</u> in total? **<u>14</u>**

<u>Hydrates:</u> ionic compounds that have water molecules attached to the ions and written into its chemical formula

Ex: $Na_2CO_3 \bullet 7\ H_2O$ in *1 mole* of this substance, there are

2 mol of Na atoms
1 mol of C atoms
10 mol of O atoms
14 mol of H atoms

<u>Molar Mass:</u> the mass in grams of 1 mole of a substance

a. **molar mass is equal to the atomic weight in grams of an element taken from the periodic table**.
Example: the molar mass of carbon is 12.011 g/mol

b. molar mass unit is : grams or g/mol

c. molar mass is also called : "**molecular mass**" for molecular compounds or "**formula mass"** for ionic compounds

d. The molar mass of a compound is the sum of all the atomic weights of the atoms in the compound
Example: H2O = 2 (1.0079g) + 1 (15.999g) = 18.0 g/mol

Summary: Gram-formula mass (also called "molar mass" or GFM)

Definition: the mass of one mole of a substance

Calculate: add together all the atomic masses of the atoms in the formula

Unit: **grams/mol**

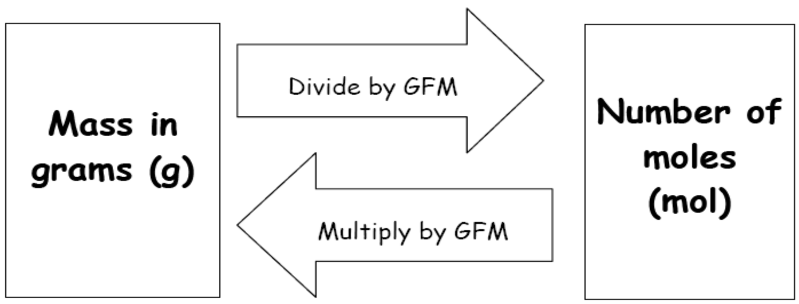
Ex: Calculate the gram-formula mass of NaCl

Atom	x	Mass		
Na	1 x	23g =	23 g	
Cl	1 x	35g =	35 g	
			$\boxed{58\ g}$	

Calculate the gram-formula mass of $Ba(NO_3)_2$

Atom	x	Mass		
Ba	1 x	137g	=	137 g
N	2 x	14g	=	28 g
O	6 x	16g	=	96 g

We will need to convert from grams to moles and vice versa for this class. The diagram below summarizes these processes:



Converting Grams to Moles:

- **From Table T**, you would use the **MOLE FORMULA:**

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

Example: How many moles are in 4.75 g of sodium hydroxide? (NaOH)

- **Step 1**: Calculate the GFM for the compound**.**

Na = 1 x =	22.9807	=	22.98977
O = 1 x =	15.9994	=	15.9994
H = 1 x =	1.00794	= +	1.00794
		=	39.99711 g/mol = 40.0g/mol

- **Step 2:** Plug the given value and the GFM into the "mole calculations" formula and solve for the number of moles.

$$\# \text{ of moles} = \frac{\text{given mass (g)}}{\text{GFM (g/mol)}} = \frac{4.75g}{40.0g/mol} = 0.119g/mol$$

Converting Moles to Grams:

The same formula can be used to convert moles back to grams

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

Example: You have a 2.50 mole sample of sulfuric acid (H_2SO_4). What is the mass of your sample in grams?

- **Step 1:** Calculate the GFM for the compound.

  H = 2 x = 2.01588
  S = 1 x = 32.06
  O= 4 x = 63.9976
  = 98.1 g/mol

- **Step 2:** Plug the given value of moles and the GFM into the "mole calculation" formula and solve for the mass of the sample.

  $$2.50\text{mol} = \frac{\text{grams}}{98.1\text{ g/mol}}$$

  grams =(2.50 mol) x (98.1 g/mol)
  =245 g H_2SO_4

Lesson 3: Types of Formulas

Objective: To compare and contrast Empirical and molecular formulas. To calculate the molecular formula from the empirical formula and molecular mass.

In order to make it easier to describe elements and molecules, **chemical formulas** are used. Chemical formulas are used to describe the types of atoms and their numbers in an element or compound. The atoms of each element are represented by one or two different letters. When more than one atom of a specific element is found in a molecule, a subscript is used to indicate this in the chemical formula.

The **Empirical formula** is the simplest WHOLE NUMBER ratio of atoms in a compound or molecule. The subscripts CANNOT be reduced any further

***Ionic compounds are already empirical formulas.

The **Molecular formula** is the ACTUAL FORMULA for a compound. It is always whole number multiples of empirical formula. The Subscripts can be reduced same as when you reduce a fraction. Remember…subscripts are just ratios of how the atoms are coming together in a compound!

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}

Determining 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

Determining MOLECULAR Formula from the Empirical formula and molecular mass

- **Step 1:** Calculate Gram Formula Mass of the **EMPIRICAL FORMULA**
- **Step 2:** Divide the MASS of the **molecular formula** by the MASS of the **empirical formula**.
- **Step 3:** MULTIPLY the subscripts of the empirical formula by the answer in step 2.

<u>Example</u>: The empirical formula of a compound is C_2H_3, and the molecular mass is 54.0 grams/mole. What is the molecular formula?

Step 1. Determine the GFM of the empirical formula.

C_2H_3 = (2 C X 12.0 g/mol) + (3 H X 1.0 g/mol) = 27.0 g/mole

Step 2. Divide the molecular mass by the empirical mass. This will give you a whole-number multiple that tells you how many times larger the molecular formula is than the empirical formula.

(54.0 g/mol) / (27.0 g/mol) = 2

Step 3. Multiply the whole number by the empirical formula. This will give the molecular formula.

2 X C_2H_3 = C_4H_6

Lesson 4 Types of Chemical Reactions

Objective: Identify various types of reactions: synthesis, decomposition, single replacement, & double replacement

1. **Synthesis**
   During a **SYNTHESIS** or **COMBINATION** reaction, substances COMBINE to form a new compound with new chemical and physical properties.

   General pattern: **A + B → AB**

   Example:

   $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

2. **Decomposition**
   During a **DECOMPOSITION** reaction one reactant BREAKS APART into TWO or more elements or compounds.

   General pattern: **AB → A + B**

   Example:
   $2HgO(s) \rightarrow 2Hg(\ell) + O_2(g)$

3. **Single Replacement**
   During a **SINGLE REPLACEMENT** reaction a more active METAL replaces a less active METAL from its compound

   General Pattern: **AB + C → CB + A**
   Use **Table J** for the activity of metals and nonmetals

   Example:
   $Zn_{(s)} + CuSO_{4(aq)} \quad Cu_{(s)} + ZnSO_{4(aq)}$

   Zinc replaces copper because zinc is more active than copper (HIGHER on chart)

   $Cu_{(s)} + ZnSO_{4(aq)}$ NO REACTION
   Copper cannot replace zinc because it is LOWER than ZINC on chart)

Table J
Activity Series**

	Metals	Nonmetals	
Most	Li	F_2	Most
	Rb	Cl_2	
	K	Br_2	
	Cs	I_2	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	**H_2		
	Cu		
	Ag		
Least	Au		Least

**Activity Series based on hydrogen standard

4. Double Replacement

During a **DOUBLE REPLACEMENT** reaction METALS in two aqueous compounds SWITCH places.

General Pattern: **AB + CD → CB + AD**

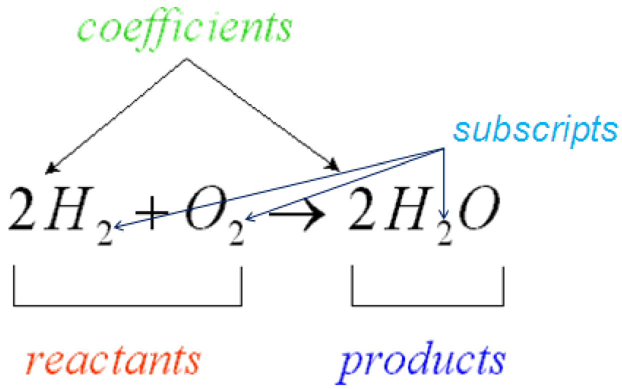
Example: $AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$

A **PRECIPITATE** is a solid that forms in solution.

Check **Table F** to see if there is a PRECIPITATE during a double replacement reaction.

Lesson 5: Balancing Equations

Objective: Balance a chemical reaction by adjusting only the coefficients



Review:

- **REACTANTS** = the STARTING substances in a chemical reaction (found to the LEFT of the arrow)
- **PRODUCTS** = a substance **PRODUCED** by a chemical reaction (found to the RIGHT of the arrow)
- **COEFFICIENT** = the integer in front of an element or compound indicates the number of moles present
- **SUBSCRIPT** = the integer to the lower right of an element indicates the number of atoms present

*__COEFFICIENTS__ and **SUBSCRIPTS** tell us how many moles we have for each element

Chemical Symbols (states of matter)

(s) solid
(g) gas
(l) liquid
(aq) dissolved in water (aqueous)

I

Example: **: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$**

- **CATALYST=** A substance that SPEEDS UP a reaction written ABOVE the arrow

Example: $2\ H_2O_2(aq)$ ---- KI -> $2\ H_2O(l) + O_2(g)$
The catalyst in this example is KI

You Must Watch this!!!!!
[Balancing chemical reactions]

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*

An equation is balanced when:

- The number of ATOMS of each type present are the SAME on both sides of the equation.
- **COEFFICIENTS** are **ONLY** used to BALANCE equations

****NOTE: WE <u>NEVER</u> CHANGE THE SUBSCRIPTS IN A FORMULA!**

<u>Steps for Balancing Equations</u>

Step 1: Find the most complex compound in the equation. Balance the elements found in that compound on the opposite side of the arrow by changing the coefficients for those atoms.

Step 2: Continue balancing until all atoms are balanced (save pure elements for last)

Step 3: Go back and check each atom to see if it is balanced on both sides of the equation.

Step 4: POLYATOMIC IONS may be balanced as a **SINGLE UNIT** rather than as separate elements as long as they stay intact during the reaction.

<u>Lesson 6 : Calculating Mole Ratios</u>

<u>Objective:</u> Solve mole-mole Stoichiometry problems given a balanced reaction

Stoichiometry (pronounced as: stow – ik – ee – **om**' – etree) is the calculation of quantities in chemical equations. The question involved with Stoichiometry is "How much…?
For example: based on the balanced chemical equation **$3\ A_2(gs) + 2\ B(aq) \rightarrow 2\ A_3B(s)$**
How much:

a) of reactant A is needed to completely react with 5.00 g of reactant B?
b) solid product will form if 0.25 g of reactant A is used?

One of the important concepts in Stoichiometry is that of **mole ratios. Mole ratios are the ratios of the number of moles of each reactant and product to each other.**

When we balance chemical equations with the coefficients, we set up the whole number ratios between reactants, between reactants and each product, and between products.

For example, consider the following balanced chemical equation:

$$N_2\,(g) + 3\ H_2\,(g) \rightarrow 2\ NH_3\,(g)$$

Using the coefficients to represent moles of a substance, **1 mole of nitrogen gas** reacts with **3 moles of hydrogen gas** to form **2 moles of ammonia gas.**

Based on the correctly balanced chemical equation, the **mole ratios** for this reaction are:

ratios between reactants: $\dfrac{\text{1 mole } N_2}{\text{3 moles } H_2}$ $\dfrac{\text{3 moles } H_2}{\text{1 mole } N_2}$

ratios between each reactant and the product:

$\dfrac{\text{1 mole } N_2}{\text{2 moles } NH_3}$ $\dfrac{\text{2 moles } NH_3}{\text{1 mole } N_2}$ $\dfrac{\text{3 moles } H_2}{\text{2 moles } NH_3}$ $\dfrac{\text{2 moles } NH_3}{\text{3 moles } H_2}$

*******Mole ratios for a given chemical equation are always based on the <u>balanced</u> chemical equation and <u>NEVER</u> change. ********
For example: **$N_2 + 3\ H_2 \rightarrow 2\ NH_3$**

1 mole	3 moles	2 moles
2 moles	6 moles	4 moles
3 moles	9 moles	6 moles
0.5 moles	1.5 moles	1.0 moles

Remember: the balanced chemical equation represents the <u>SMALLEST</u> whole number ratios of reactants and products.

Each of the mole ratios from the <u>original balanced equation</u> may be used as a conversion factor to change from moles of one substance to moles of another substance in the <u>same</u> chemical equation.
You MUST use the <u>ORIGINAL balanced chemical equation</u> with the <u>correct</u> coefficients as your reference for mole ratios.

Using the reaction of nitrogen gas and hydrogen gas to form ammonia gas;

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

Suppose you had only 0.25 mole $N_2(g)$ available. How much hydrogen gas would you need to react completely with the available nitrogen gas? How much ammonia would be formed?

(1) Start with the amount you are given in the problem

(2) multiply the starting amount by the appropriate mole ratio to convert from the unit with which you start to the substance desired

(3) cancel units

(4) do the arithmetic and write the answer with appropriate significant figures, correct unit, and correct chemical formula

$$0.25\ \cancel{\text{mole } N_2} \times \frac{\textbf{3 moles } H_2}{\textbf{1 } \cancel{\text{mole } N_2}} = 0.75 \text{ moles } H_2 \text{ are needed}$$

$$0.25\ \cancel{\text{mole } N_2} \times \frac{\textbf{2 moles } NH_3}{\textbf{1 } \cancel{\text{mole } N_2}} = 0.50 \text{ mole } NH_3 \text{ would be formed}$$

NOTE: In these calculations, the numbers in the mole ratios do NOT have significant figures. USE THE NUMERICAL VALUES AS PRINTED IN THE QUESTION AS THE BASIS FOR YOUR SIGNIFICANT FIGURES. In this calculation, the given value of
0.25 mole $N_2(g)$ has 2 significant figures, therefore the answers must each have 2 sig. figs.

More examples:
Consider the reaction between sodium metal and chlorine gas to form solid sodium chloride.

$$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$$

If 25.0 moles of sodium chloride are formed, how many moles of sodium metal were used? How many moles of chlorine gas were used?

$$25.0\ \cancel{\text{moles NaCl}} \times \frac{2 \text{ moles Na}}{2\ \cancel{\text{moles NaCl}}} = 25.0 \text{ moles Na were used}$$

$$25.0\ \cancel{\text{moles NaCl}} \times \frac{1 \text{ mole } Cl_2}{2\ \cancel{\text{moles NaCl}}} = 12.5 \text{ moles } Cl_2 \text{ were used}$$

Solid potassium chlorate decomposes when heated to produce solid potassium chloride and oxygen gas according to the following equation:

$$2\ KClO_3(s) \xrightarrow{\text{heat}} 2\ KCl(s) + 3\ O_2(g)$$

If 4 moles of potassium chloride are formed, how many moles of oxygen gas are formed? How many moles of potassium chlorate decomposed?

$$4\ \cancel{\text{moles KCl}} \times \frac{3\ \text{moles } O_2}{2\ \cancel{\text{moles KCl}}} = 6\ \text{moles } O_2 \text{ are formed}$$

$$4\ \cancel{\text{moles KCl}} \times \frac{2\ \text{moles } KClO_3}{2\ \cancel{\text{moles KCl}}} = 4\ \text{moles } KClO_3 \text{ decomposed}$$

Lesson 7: Percent Composition

Objective: To apply a formula to calculate % composition

A common practice in the chemical laboratory is to determine the **percent composition** of each element in a compound and use that information to determine the formula that shows the simplest whole-number ratio of elements present in the compound, the **empirical formula** of the compound.

Percent composition is defined as how much of an element or compound is in a mixture and can be calculated with a simple formula.

Calculating Percent Composition : Formula located on Table T

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

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

- **Step 1:** Calculate the GFM for the compound.

Ca = 1 × 40.08 = 40.1 g (this is the "part" Ca)
Cl = 2 × 35.453 = 70.9 g (this is the "part" Cl)

= **111.0 g/mol**

- **Step 2:** Use the formula to find the % composition of each element or "part" in our compound (to the nearest tenth of a %).

$$\%Ca \text{ in } CaCl_2 = \frac{40.08 \text{ g}}{111.0 \text{ g}} = 36.11\ \%$$

% Composition of Hydrates

Water, the most common chemical on earth, can be found in the atmosphere as water vapor. Some chemicals, when exposed to water in the atmosphere, will reversibly either adsorb it onto their surface or include it in their structure forming a complex in which water generally bonds with the cation in ionic substances. The water present in the latter case is called *water of hydration* or *water of crystallization*. Common examples of minerals that exist as hydrates are gypsum ($CaSO_4 \bullet 2H_2O$), Borax ($Na_3B_4O_7 \bullet 10H_2O$) and Epsom salts ($MgSO_4 \bullet 7H_2O$). Hydrates generally contain water in specific amounts. The hydrate's formula are represented using the formula of the ***anhydrous salt*** (non-water) component of the complex followed by a dot then the water (H_2O) preceded by a number corresponding to the ratio of H_2O moles per mole of the anhydrous component present. They are typically named by stating the name of the anhydrous component followed by the Greek prefix specifying the number of moles of water present then the word ***hydrate*** (example: $MgSO_4 \bullet 7H_2O$: magnesium sulfate heptahydrate).

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

Notice how **WATER** molecules are BUILT INTO the chemical formula

Anhydrous salt **-** A crystal with NO WATER trapped inside its lattice

Calculate % Composition of water in a Hydrate

Step 1: Calculate GFM of the HYDRATE
Step 2: Plug into % composition formula from Table T

$$\text{\% composition by mass} = \frac{\text{mass water}}{\text{mass of whole hydrate}} \times 100$$

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

Step 1- Calculate GFM of Hydrate

Na =	2 × 22.98977	= 45.97954	
C =	1 × 12.0111	= 12.0111	
O =	3 × 15.9994	= 47.9982	
H_2O =	10 × 18.01528	= 180.1528	=part
	GFM	= 286.1 g/mol	=whole

Step 2- Plug values into Formula

$$\text{\% } H_2O \text{ by mass} = \frac{180.2\text{g}}{286.1\text{g}} \times 100 = 63.0\ \%$$

****Pay attention because in some problems you will have to calculate the amount of water lost by subtracting the anhydrate mass from the hydrate mass.