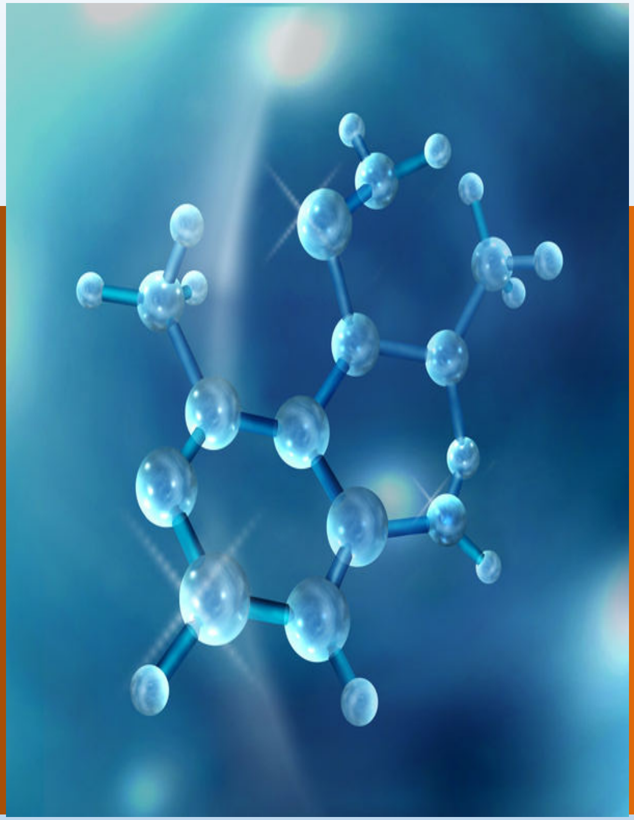


Unit 7

Percent Composition
Empirical Formulas
Molecular Formulas

Percent Composition of M & M's

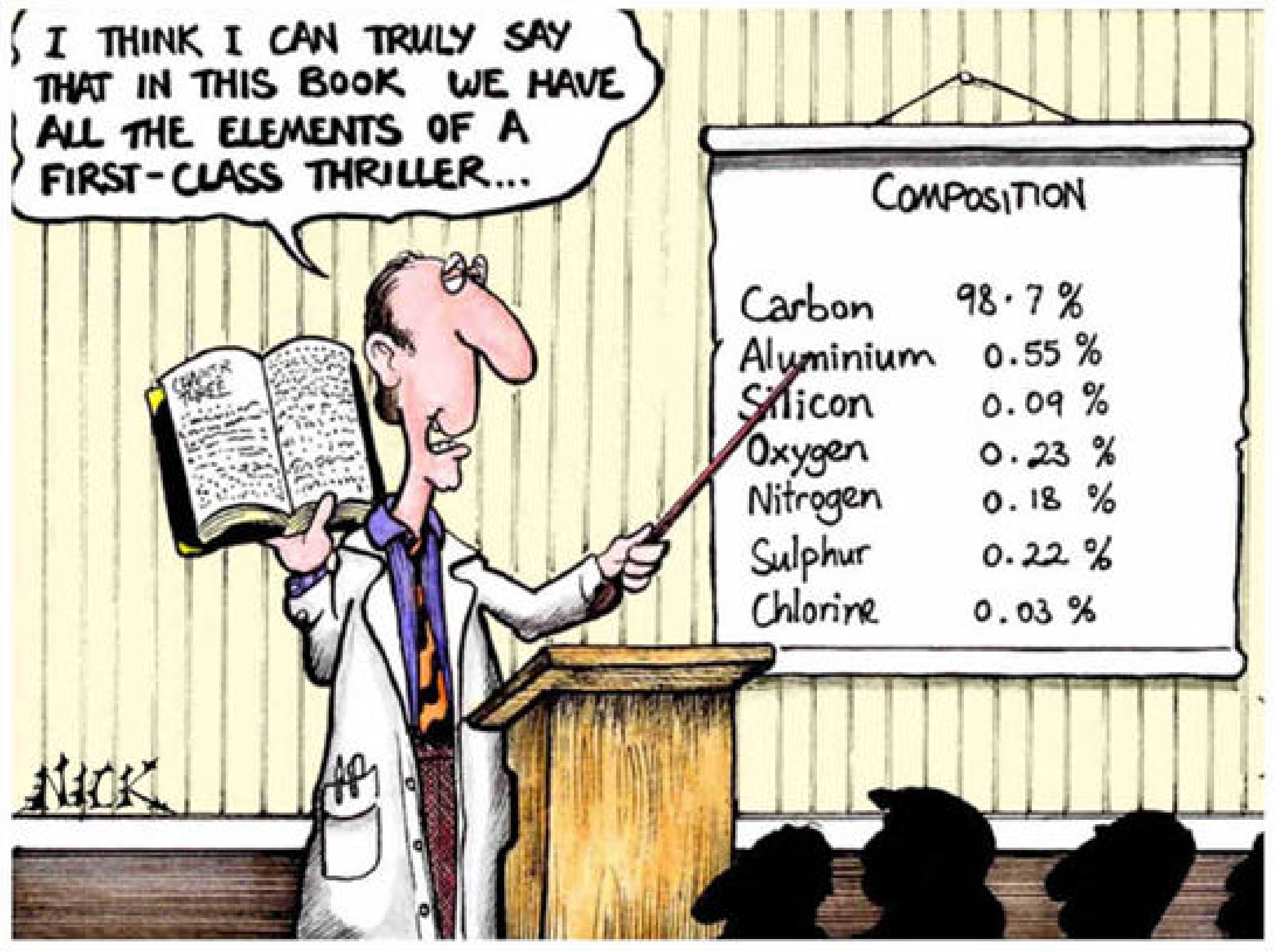
1. Open your bag of M & M's.
2. Separate the colors onto a plate or napkin.
3. Use your white board to answer the following questions:
   - **How many of each color do you have?**
   - **How many total M & M's do you have?**
   - **What is the percent of each color from your bag?**

Percent Composition of M & M's

According to Mars, the maker of M&Ms, the % Composition should be:

Red:	20%
Orange:	10%
Brown:	30%
Green:	10%
Yellow:	20%
Blue:	10%

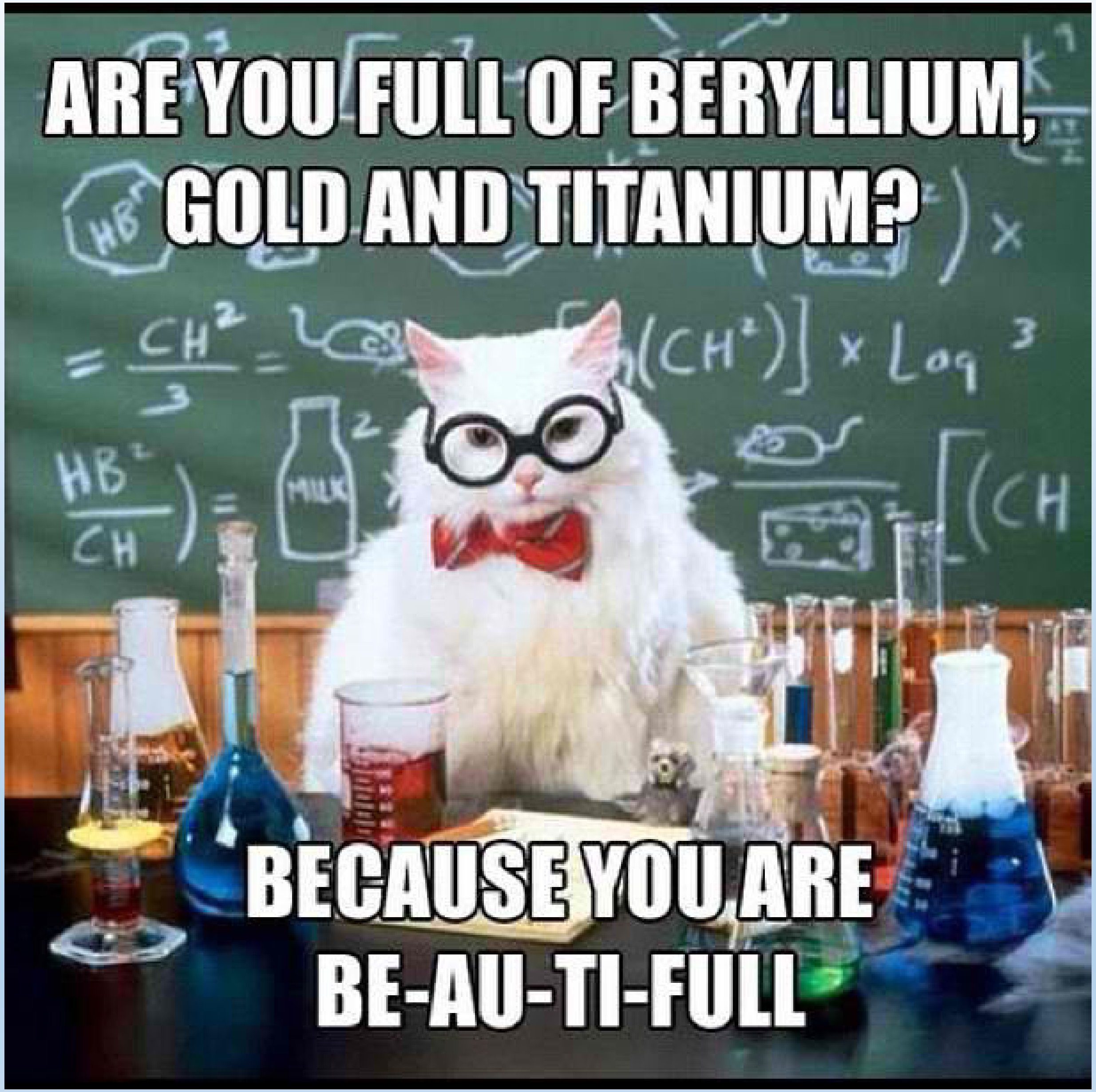
Percent Composition



Calculating Percent Composition

1. Separate your atoms.
2. Multiply the atomic mass of each atom by the # of atoms to get the mass of each element.
3. Add up the masses from all the elements to get the mass of the compound.
4. A percent is always calculated by part over whole times 100.

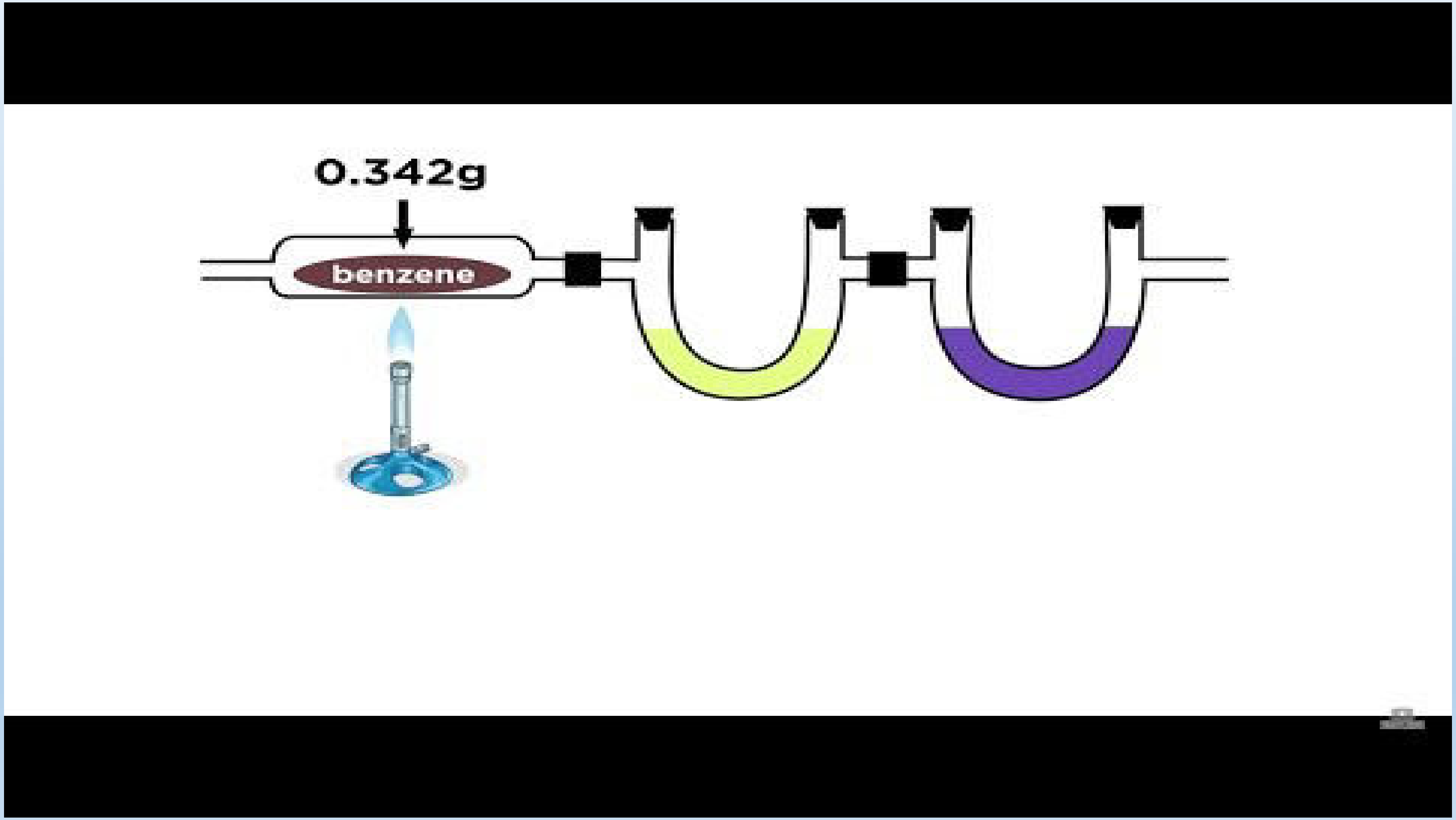
$$\text{Percent Composition} = \frac{\text{Mass of Element}}{\text{Mass of Compound}} \times 100$$


ARE YOU FULL OF BERYLLIUM,
GOLD AND TITANIUM?
BECAUSE YOU ARE
BE-AU-TI-FULL

What is the chemical formula for a banana?

$Ba(na)_2$

Empirical Formulas



Empirical and Molecular Formulas

Chemical Name	Molecular Formula	Empirical Formula	Scale Factor
glyceraldehyde	$C_3H_6O_3$	CH_2O	
erythrose	$C_4H_8O_4$	CH_2O	
arabinose	$C_5H_{10}O_5$	CH_2O	
glucose	$C_6H_{12}O_6$	CH_2O	

Calculating the Empirical Formula

1. Convert % of each element into grams.
2. Divide by molar mass to get moles of each element.
3. Divide each # of moles by the smallest amount.
4. Assign subscripts to the Empirical Formula.

For example, in the previous handout, we determined that Aluminum Sulfate, $Al_2(SO_4)_3$, was composed of 15.77% Aluminum, 28.12% Sulfur, and 56.11% oxygen. Now we need to drop the % signs and replace it with grams for each element.

Step 1: Change % to grams and determine number of moles.

$$15.77\ grams\ Al \times \left(\frac{1\ mole}{26.98}\right) = 0.5845\ moles\ Al$$

$$28.17\ grams\ S \times \left(\frac{1\ mole}{32.07}\right) = 0.8784\ moles\ S$$

$$56.11\ grams\ O \times \left(\frac{1\ mole}{16.00}\right) = 3.5069\ moles\ O$$

Step 2: Divide each by element with fewest moles.

$$0.5845\ moles\ Al \div 0.5845\ moles = 1.0\ Aluminum$$

$$0.8784\ moles\ S \div 0.5845\ moles = 1.5\ Sulfur$$

$$3.5069\ moles\ O \div 0.5845\ moles = 6.0\ Oxygen$$

Step 3: Convert ratio to whole numbers and write as subscripts.

$1.0\ Aluminum$ $\times 2$	$1.5\ Sulfur$ $\times 2$	$6.0\ Oxygen$ $\times 2$
Al_2	S_3	O_{12}

Step 4: If the compound is ionic (contains metal), the anions may need to be rearranged to create a polyatomic ion.

$$Al_2(SO_4)_3$$

Calculating the Molecular Formula

1. Calculate the Empirical mass based on the Empirical Formula.

2. Divide the Molecular mass by the Empirical mass to determine the scale factor.

3. Multiply each subscript in the Empirical Formula by the scale factor in order to find the Molecular Formula.

Example: A substance has an empirical formula of C_4H_4S, and its molecular mass is 168 g/mole. What is the molecular formula of the compound?

Step 1: Determine the mass of the empirical formula, C_4H_4S.

Carbon	12 amu x 4 atoms =	48 g/mol
Hydrogen	1 amu x 4 atoms =	4 g/mol
Sulfur	32 amu x 1 atom =	+ 32 g/mol

$\boxed{\text{Empirical Formula Mass = 84 g/mol}}$

Step 2: Enter data into squares of the box, and determine the scale factor.

Empirical Formula: C_4H_4S	Empirical Mass: 84 g/mol
Molecular Formula: $C_?\ H_?\ S_?$	Molecular Mass: 168 g/mol

Scale up = x 2

Step 3: Since the scale up of the masses is doubled, the scale up of the formula is also doubled. Therefore the molecular formula will equal...

$C_{(4x2)}\ H_{(4x2)}\ S_{(1x2)}$ or $C_8H_8S_2$