

#### **Molecular and Formula Mass**

The molecular mass is the mass in atomic mass units (amu) of an individual molecule.

To calculate molecular mass, multiply the atomic mass of each element in a molecule by the number of atoms of that element and then total the masses

Molecular mass of H<sub>2</sub>O:

2(atomic mass units of H) +(1)(atomic mass units of O)

2(1.008 amu) + (1)(16.00 amu) = 18.02 amu

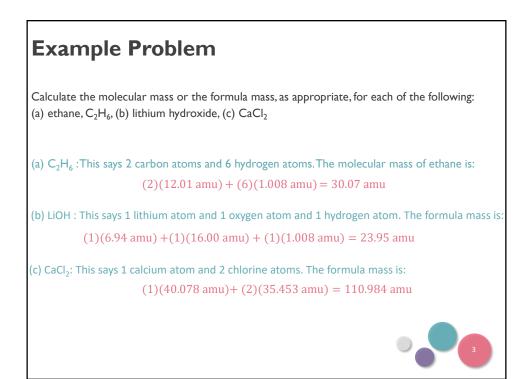
Because the atomic masses on the periodic table are average atomic masses, the result of such a determination is an average molecular mass, sometimes referred to as the **molecular weight**.

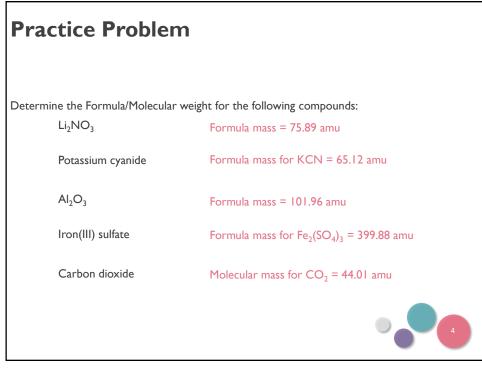
Although an ionic compound does not have a molecular mass, we can use its empirical formula to determine its **formula mass** (the mass of a "formula unit"), sometimes called the **formula weight**.

The process is the same:

To calculate formula mass, multiply the atomic mass for each element in a formula unit by the number of atoms of that element and then total the masses







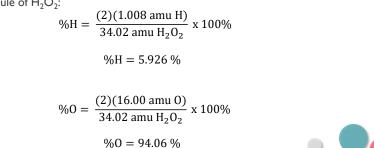
## **Percent Composition of Elements**

A list of the percent by mass of each element in a compound is known as the compound's **percent composition by mass**.

mass percent of an element =  $\frac{n \times atomic mass of element}{molecular or formula mass of compound} \times 100\%$ 

where n is the number of atoms of the element in a molecule or formula unit of the compound

For a molecule of  $H_2O_2$ :



5

## **Percent Composition of Elements**

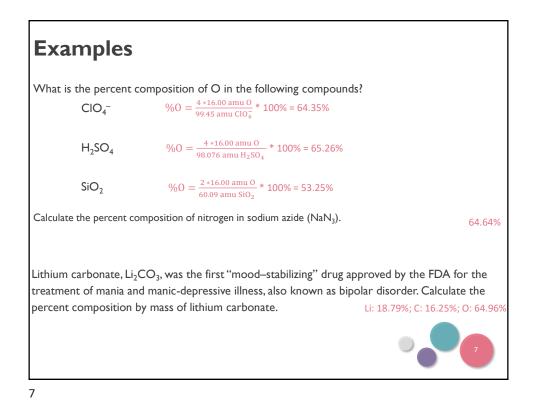
We could also have used the empirical formula of hydrogen peroxide (HO) for the calculation.

In this case, we could have used the **empirical formula mass**, the mass in amu of one empirical formula, in place of the molecular formula.

The empirical formula mass of  $H_2O_2$  (the mass of HO) is 17.01 amu.

%H = 
$$\frac{(1)(1.008 \text{ amu H})}{17.01 \text{ amu HO}} x 100\%$$

$$\%0 = \frac{(1)(16.00 \text{ amu } 0)}{17.01 \text{ amu } \text{HO}} \ x \ 100\%$$
$$\%0 = 94.06\%$$

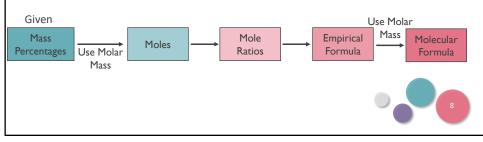


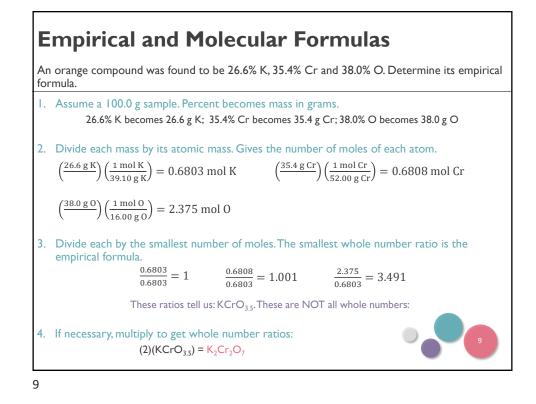
#### Determination of Empirical Formula and Molecular Formula from Percent Comp.

Using the concepts of the mole and molar mass, we can now use an experimentally determined percent composition to determine the empirical and/or molecular formula.

The empirical formula gives only the ratio of atoms in a molecule (many compounds can have the same empirical formula).

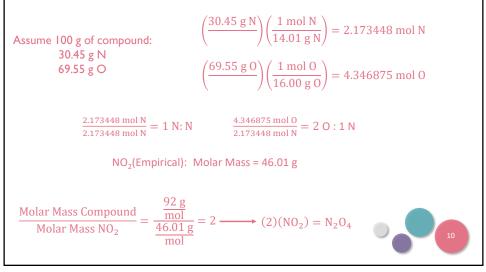
A compound's empirical formula can be determined from its percent composition. A compound's molecular formula is determined from the molar mass and empirical formula.





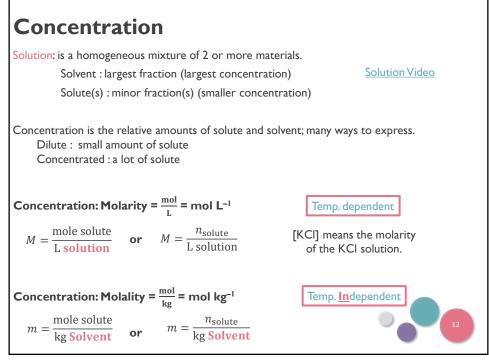
#### **Practice**

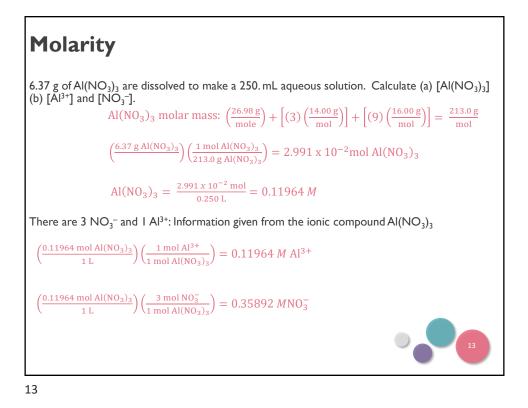
Determine the empirical formula of a compound that is 30.45 percent nitrogen and 69.55 percent oxygen by mass. Given that the molar mass of the compound is approximately 92 g mol<sup>-1</sup>, determine the molecular formula of the compound.

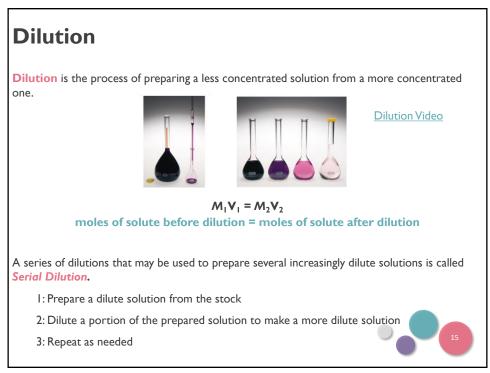


#### **Poll Question: Campusknot**

(Ch. 6-Slide 11) How many carbon atoms are found in the empirical formula for a compound that contains 26.1% carbon, 4.3% hydrogen and 69.6 % oxygen?







## Dilution

In an experiment, a student needs 1.00 L of a 0.400 M  $KMnO_4$  solution. A stock solution of 1.00 M  $KMnO_4$  is available. How much of the stock solution is needed?

**Solution:** Use the relationship that moles of solute before dilution = moles of solute after dilution. (1.00 M/KM = 0.001) = (0.400 M/KM = 0.011)

 $(1.00 \text{ M KMnO}_4)(V_1) = (0.400 \text{ M KMnO}_4)(1.00 \text{ L})$ 

V<sub>1</sub> = 0.400 L or 400 mL

To make the solution:

I. Pipet 400 mL of stock solution into a 1.00 L volumetric flask.

2. Carefully dilute to the calibration mark.

Because most volumes measured in the laboratory are in milliliters rather than liters, it is worth pointing out that the equation can be written as

$$M_1 \times \mathbf{mL}_1 = M_2 \times \mathbf{mL}_2$$

16

## **Practice**

Commercial concentrated sulfuric acid is 17.8 M. If 75.0 mL of this acid is diluted to 1.00 L, what is the final concentration of the acid?

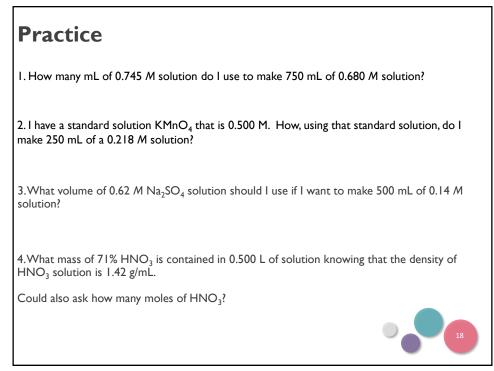
$$M_{\text{conc}} = 17.8 \text{ M} \qquad V_{\text{conc}} = 75.0 \text{ mL}$$

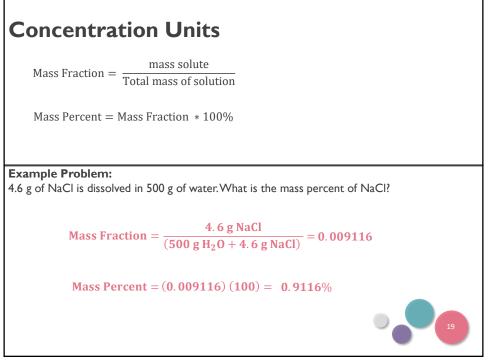
$$M_{\text{dil}} = ? \qquad V_{\text{dil}} = 1000. \text{ mL}$$

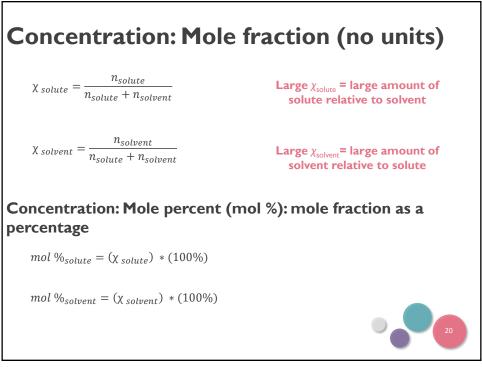
$$M_{\text{dil}} = \frac{M_{\text{con}}V_{\text{con}}}{V_{\text{dil}}} = \frac{(17.8 \text{ M})(75.0 \text{ mL})}{(1000 \text{ mL})} = 1.34 \text{ M}$$

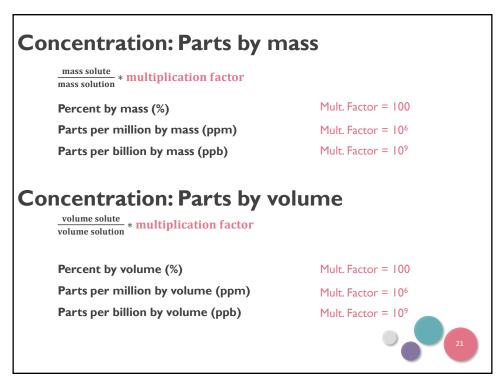
Prepare a 0.5000 M solution of potassium permanganate in a 250.0 mL volumetric flask.

Mass of KMnO<sub>4</sub> required: to find mass we have to find moles. We can use the volume and concentration for this  $\frac{\text{Remember: } M = \frac{\text{mol}}{\text{L}}}{\left(\frac{0.2500 L}{0}\right)\left(\frac{0.5000 \text{ mol } KMnO_4}{L}\right)} = 0.1250 \text{ mol } KMnO_4$   $\left(\frac{0.1250 \text{ mol}}{1 \text{ mol}}\right)\left(\frac{158.03 \text{ g } KMnO_4}{1 \text{ mol}}\right) = 19.75 \text{ g } KMnO_4$ 17



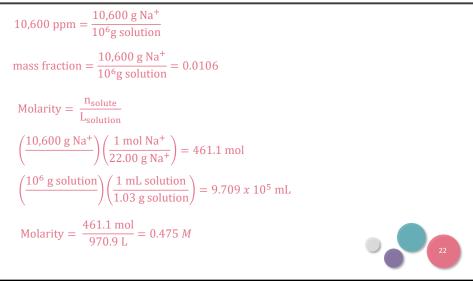






# **Converting Units**

Sea water is 10,600 ppm Na<sup>+</sup>. Calculate the mass fraction and molarity of sodium ions in sea water. The density of sea water is 1.03 g/mL.

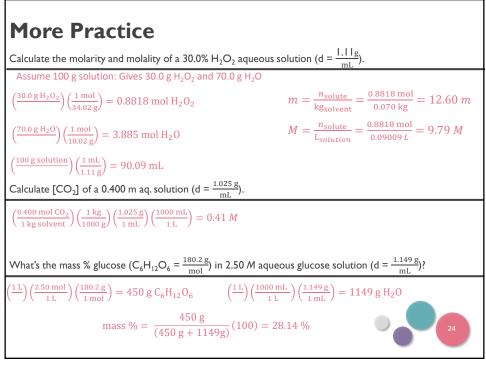


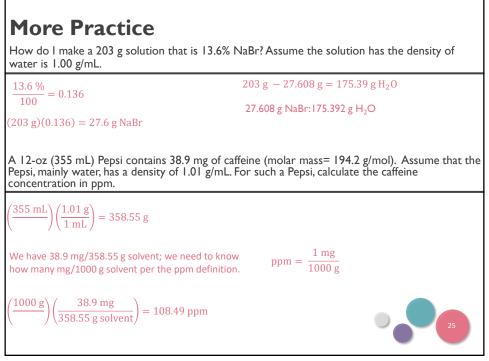
22

# **Converting Units**

A solution is prepared by dissolving 17.2 g ethylene glycol ( $C_2H_6O_2$ ) in 0.500 kg of water (at 25°C). The final solution volume is 515 mL. Calculate the following concentrations: a.) M b.) m c.) % by mass d.) mol fraction solute e.) mol % solute

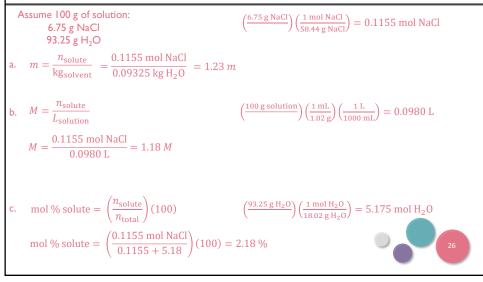
$$\left(\frac{17.2 \text{ g } \text{C}_2 \text{H}_6 \text{O}_2}{62.07 \text{ g}}\right) \left(\frac{1 \text{ mol}}{62.07 \text{ g}}\right) = 0.2771 \text{ mol } \text{C}_2 \text{H}_6 \text{O}_2$$
  
a.) Molarity  $= \frac{n_{\text{solute}}}{\text{L}_{\text{solution}}} = \frac{0.2771 \text{ mol}}{0.515 \text{ L}} = 0.538 M$   
b.) Molality  $= \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{0.2771 \text{ mol}}{0.500 \text{ kg}} = 0.554 m$   
c.) Mass %  $= \left(\frac{\text{mass}_{\text{solute}}}{\text{mass}_{\text{total}}}\right) (100) = \left(\frac{17.2 \text{ g}}{(17.2+500 \text{ g})}\right) (100) = 3.33 \%$   
d.)  $\left(\frac{500 \text{ g } \text{H}_2 \text{O}}{18.02 \text{ g } \text{H}_2 \text{O}}\right) = 27.75 \text{ mol}$   $X_{\text{solute}} = \frac{0.2771 \text{ mol}}{(0.2771+27.75 \text{ mol})} = 9.89 \text{ x}10^{-3}$   
e.) mol % solute =  $(9.89 \text{ x}10^{-3})(100) = 0.989 \%$ 





## **More Practice**

An aqueous solution is 6.75 % by mass NaCl and has a density of 1.02 g/mL.What is the a. molality, b. Molarity, and c. Mole percent of NaCl?



26

## **More Practice**

In the lab, you are making 0.829 m manganese(II) sulfate solution using 350 grams of water. How many grams of MnSO<sub>4</sub> should you add?

$$\left(\frac{350 \text{ g H}_2\text{O}}{1000 \text{ g H}_2\text{O}}\right) \left(\frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}}\right) \left(\frac{0.829 \text{ mol MgSO}_4}{1 \text{ kg H}_2\text{O}}\right) \left(\frac{120.364 \text{ g MgSO}_4}{1 \text{ mol MgSO}_4}\right) = 34.924 \text{ g MgSO}_4$$

Calculate the molality of a.) 2.50 M CaCl<sub>2</sub> solution (density of solution= 1.12 g/mL), b.) 48.2 % by mass KBr solution, and c.) 64.8 grams of ethylene glycol ( $C_2H_6O_2$ ) mixed with water to make 3,500 grams of solution.

a. 
$$\left(\frac{1 \text{ L}}{1 \text{ L}}\right) \left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) \left(\frac{1.12 \text{ g}}{1 \text{ mL}}\right) = 1120 \text{ g}$$
  
 $m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{2.50 \text{ mol CaCl}_2}{1.120 \text{ kg}} = 2.23 m$   
b. Assume a 100 g solution; 48.2 g KBr and 51.8 g H<sub>2</sub>O  
 $\left(\frac{48.2 \text{ g KBr}}{119.002 \text{ g KBr}}\right) \left(\frac{1 \text{ mol KBr}}{119.002 \text{ g KBr}}\right) = 0.405 \text{ mol KBr}$   
 $m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{0.405 \text{ mol KBr}}{0.0518 \text{ kg H}_2O} = 7.82 m$   
c.  $\left(\frac{64.8 \text{ g C}_2 \text{ H}_6 \text{ O}_2}{62.08 \text{ g C}_2 \text{ H}_6 \text{ O}_2}\right) = 1.0438 \text{ mol C}_2 \text{ H}_6 \text{ O}_2$   
 $m = \frac{n_{\text{solute}}}{\text{kg}_{\text{solvent}}} = \frac{1.0438 \text{ mol C}_2 \text{ H}_6 \text{ O}_2}{3.435 \text{ kg H}_2O} = 0.304 m$ 

