RELATING MASS AND MOLES

- Use DIMENSIONAL ANALYSIS (a.k.a "drag and drop")

- Need CONVERSION FACTORS - where do they come from?

- We use ATOMIC WEIGHT as a conversion factor.

Mg : 24.31 | 24.31 g Mg = 1 mol Mg

"mol" is the abbreviation for "mole"

Example: How many moles of atoms are there in 250. g of magnesium metal?

24.31 g Mg = mol Mg

$$250.\,\cancel{g\,Mg} \times \frac{mol\ Mg}{24.31\,\cancel{g\,Mg}} = \boxed{10.3\ mol\ Mg}$$

Example: You need 1.75 moles of iron. What mass of iron do you need to weigh out on the balance?

$$55.85\,g\ Fe = mol\ Fe$$

$$1.75\ \cancel{mol\ Fe} \times \frac{55.85\,g\ Fe}{\cancel{mol\ Fe}} = \boxed{97.7\,g\ Fe}$$

WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

Example: 25.0 g of WATER contain how many MOLES of water molecules?

(H_2O)

H_2O

$H: 2 \times 1.008 = 2.016$

$O: 1 \times 16.00 = 16.00$

18.016 — FORMULA WEIGHT of water

$\boxed{\text{Formula weight = mass of one mole of either an element OR a compound!}}$

$18.016\,g\,H_2O = mol\,H_2O$

$$25.0\,\cancel{g\,H_2O} \times \frac{mol\,H_2O}{18.016\,\cancel{g\,H_2O}} = \boxed{1.39\,mol\,H_2O}$$

Formula weight goes by several names:

- For atoms, it's the same thing as ATOMIC WEIGHT
- For molecules, it;s called MOLECULAR WEIGHT
- Also called "MOLAR MASS"

Example: How many grams of ammonium carbonate do we need to weigh out to get 3.65 moles of ammonium carbonate?

First, we need to find out the formula of ammonium carbonate!

NH_4^+ CO_3^{2-}

NH_4^+

$(NH_4)_2CO_3$

N: 2 x 14.01

H: 8 x 1.008

C: 1 x 12.01

O: 3 x 16.00

96.094] Formula weight of ammonium carbonate

$96.094\ g\ (NH_4)_2CO_3 = mol\ (NH_4)_2CO_3$

$$3.65\ mol\ (NH_4)_2CO_3 \times \frac{96.094\ g\ (NH_4)_2CO_3}{mol\ (NH_4)_2CO_3} = \boxed{351\ g\ (NH_4)_2CO_3}$$

PERCENTAGE COMPOSITION

- sometimes called "percent composition" or "percent composition by mass"

- the percentage of each element in a compound, expressed in terms of mass

Example: Find the percentage composition of ammonium nitrate.

NH_4^+ NO_3^-

NH_4NO_3:

N: 2 x 14.01 = 28.02

H: 4 x 1.008 = 4.032

O: 3 x 16.00 = 48.00

80.052 g NH_4NO_3 = mol NH_4NO_3

$$\%N: \frac{28.02\text{ g N}}{80.052\text{ g } NH_4NO_3} \times 100\% = 35.0\%\ N$$

$$\%H: \frac{4.032\text{ g H}}{80.052\text{ g } NH_4NO_3} \times 100\% = 5.0\%\ H$$

$$\%O: \frac{48.00\text{ g O}}{80.052\text{ g } NH_4NO_3} \times 100\% = 60.0\%\ O$$

These percentages should sum to 100%, but you may see some roundoff error - depending on which decimal place you round to!

So far, we have

ch8

- looked at how to determine the composition by mass of a compound from a formula
- converted from MASS to MOLES (related to the number of atoms/molecules)
- converted from MOLES to MASS

Are we missing anything?

Sec 15.4 p483-488

- What about SOLUTIONS, where the desired chemical is not PURE, but found DISSOLVED IN WATER?
- How do we deal with finding the moles of a desired chemical when it's in solution?

MOLAR CONCENTRATION

- unit: MOLARITY (M): moles of dissolved substance per LITER of solution

$$M = \text{MOLARITY} = \frac{\text{moles of solute}}{\text{L solution}}$$

6.0 M HCl solution: $\frac{6.0 \text{ mol HCl}}{\text{L}}$

There are 6.0 moles of hydrochloric acid in each liter of this solution, so you can write this relationship another way:

$$6.0 \text{ mol HCl} = 1 \text{ L}$$

If you have 0.250 L (250 mL) of 6.0 M HCl, how many moles of HCl do you have?

$$6.0 \text{ mol HCl} = \text{L}$$

$$0.250 \text{ L} \times \frac{6.0 \text{ mol HCl}}{\text{L}} = \boxed{1.5 \text{ mol HCl}}$$

If you need 0.657 moles of hydrochloric acid, how many milliliters of 0.0555 M HCl do you need to measure out?

$0.0555 \text{ mol HCl} = \text{L}$ $\text{mL} = 10^{-3}\text{L}$

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{0.0555 \text{ mol HCl}} \times \frac{\text{mL}}{10^{-3}\text{L}} = \boxed{11800 \text{ mL of } 0.0555 \text{ M HCl}}$$

This is an extremely large volume for lab-scale work. We should use a more concentrated HCl solution to get the required moles of HCl!

What if we used 6.00 M HCl?

$6.00 \text{ mol HCl} = \text{L}$ $\text{mL} = 10^{-3}\text{L}$

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} \times \frac{\text{mL}}{10^{-3}\text{L}} = \boxed{11\bar{0} \text{ mL of } 6.00 \text{ M HCl}}$$

This volume is more practical for lab work. We can measure it with a 250 mL cylinder or we can use a 100 mL cylinder twice.

If you're preparing a solution by dissolving a solid in water, you can easily calculate the molarity of the solution. How?

Just find the number of moles of solid you dissolved, then divide by the volume of the solution (expressed in liters!)

What is the molarity of a solution made by dissolving 3.50 g of NaCl in enough water to make 250. mL of solution?

$$M = \frac{\text{mol NaCl}}{\text{L solution}}$$

1 - Find moles of NaCl dissolved using the FORMULA WEIGHT of NaCl
2 - Divide moles of NaCl / LITERS of solution (Convert 250. mL to L)

1 NaCl: Na: 1 x 22.99
Cl: 1 x 35.45

Find formula weight of NaCl

$58.44\text{ g NaCl} = \text{mol NaCl}$

$$3.50\text{ g NaCl} \times \frac{\text{mol NaCl}}{58.44\text{ g NaCl}} = 0.059890\text{ mol NaCl}$$

2 $\text{mL} = 10^{-3}\text{ L}$

$$250.\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} = 0.250\text{ L}$$

$$M = \frac{\text{mol NaCl}}{\text{L solution}} = \frac{0.059890\text{ mol NaCl}}{0.250\text{ L}} = \boxed{0.240\ \underline{M}\text{ NaCl}}$$

A few more examples...

Use FORMULA WEIGHT when relating mass and moles

You have a 250.g bottle of silver(I) chloride (AgCl). How many moles of AgCl do you have?

AgCl: 1 x 107.9
1 x 35.45

$143.35\ g\ AgCl = mol\ AgCl$

$$250.\ g\ AgCl \times \frac{mol\ AgCl}{143.35\ g\ AgCl} = \boxed{1.74\ mol\ AgCl}$$

How many moles of NaOH are present in 155 mL of 1.50 M NaOH?

When relating moles and VOLUME, we need to use CONCENTRATION (usually MOLARITY - M)

$mL = 10^{-3}\ L$ $\quad 1.50\ mol\ NaOH = L$

$$155\ mL \times \frac{10^{-3}\ L}{mL} \times \frac{1.50\ mol\ NaOH}{L} = \boxed{0.233\ mol\ NaOH}$$

End of material for test 3

Summer 2010 Test 3 is
7/20/2010 (T)