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

Use the atomic weight of iron (55.85 g/mol) to relate moles of iron and mass.

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

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

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

Formula weight = mass of one mole of either an element OR a compound!

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

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

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, find the chemical formula. Ammlonium carbonate is IONIC.

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 nitrate

Use the formula weight as a conversion factor.

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

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

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:

$N: 2 \times 14.01 = 28.02$

$H: 4 \times 1.008 = 4.032$

$O: 3 \times 16.00 = 48.00$

$80.052\ g\ NH_4NO_3 = mol\ NH_4NO_3$

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

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

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

These percentages should sum to 100%, but you may have a little roundoff error!

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?

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- 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

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

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

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

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

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

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

$$0.250\ L \times \frac{6.0\ \text{mol}\ HCl}{1\ L} = \boxed{1.5\ \text{mol}\ HCl}$$

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

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

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{0.0555 \text{ mol HCl}} \times \frac{\text{mL}}{10^{-3}\text{ L}} = \boxed{11\,800 \text{ mL of } 0.0555 \text{ M HCl}}$$

THis is an extremely large volume for lab-scale work. We should use a more concentrated solution to get our 0.657 moles of HC

What if we used 6.00 M HCl?

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

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

This is a more practical lab volume. You may measure this quantity in a 250 mL graduated cylinder!

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?

1 - Find moles of sodium chloride dissolved. Use formula weight of NaCl
2 - Divide moles sodium chloride / LITERS solution. We'll need to change 250 mL to L.

$NaCl$: $Na: 1 \times 22.99$
$Cl: 1 \times 35.45$
$\overline{58.44 \text{ g NaCl} = \text{mol NaCl}}$

Find formula weight!

① $$3.50 \text{ g NaCl} \times \frac{\text{mol NaCl}}{58.44 \text{ g NaCl}} = 0.059890 \text{ mol NaCl}$$

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

$$250.\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} = 0.250 \text{ L}$$ Find volume in liters!

② $$M = \frac{\text{moles NaCl}}{\text{L}} = \frac{0.059890 \text{ mol NaCl}}{0.250 \text{ L}} = \boxed{0.240 \text{ } \underline{M} \text{ NaCl}}$$

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 : Ag : 1 x 107.9

Cl : 1 x 35.45

143.35 g AgCl = mol AgCl

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

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)

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

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

End of material for test 3