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

$55.85\ g\ Fe = mol\ Fe$

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

WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

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

H_2O: H : 2 x 1.008 = 2.016
O : 1 x 16.00 = 16.00

18.016 ← FORMULA WEIGHT of water

$18.016 \text{ g } H_2O = \text{mol } H_2O$

FORMULA WEIGHT is the mass of one mole of either an element OR a compound.

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

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?

Find the fomula of ammonium carbonate!

$NH_4^+ \quad CO_3^{2-}$

NH_4^+

$(NH_4)_2CO_3$

Once you've found the formula, you can find the formula weight:

N : 2 x 14.01

H : 8 x 1.008

C : 1 x 12.01

O : 3 x 16.00

96.094

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

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

$$= \boxed{351\,g\ (NH_4)_2CO_3}$$

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_4NO_3: N: $2 \times 14.01 = 28.02$

H: $4 \times 1.008 = 4.032$

O: $3 \times 16.00 = 48.00$

$80.052\ g\ NH_4NO_3 = 1\ mol\ NH_4NO_3$

These numbers are the masses of each element in a mole of the compound!

$\%N: \dfrac{28.02\ g\ N}{80.052\ g\ total} \times 100\% = 35.0\%\ N$

$\%H: \dfrac{4.032\ g\ H}{80.052\ g\ total} \times 100\% = 5.0\%\ H$

$\%O: \dfrac{48.00\ g\ O}{80.052\ g\ total} \times 100\% = 60.0\%\ O$

Check: Make sure all the percentages sum to 100% - within roundoff error...

NOTE: CHM 110-03 TEST 2 COVERS THROUGH THIS MATERIAL

So far, we have

- 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?

- 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

dissolved substance

$$M = \text{molarity} = \frac{\text{moles of SOLUTE}}{\text{L SOLUTION}}$$

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

(read aloud as "six molar")

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

6.0 mol HCl = L

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

*See SECTIONS 4.7 - 4.10 for more information about MOLARITY and solution calculations (p 154 - 162 - 9th edition) (p 156-164 - 10th edition)

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

$$\frac{0.0555 \text{ mol HCl}}{\text{L}}$$

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{0.0555 \text{ mol HCl}} = \boxed{11.8 \text{ L}}$$

11 800 mL

This is too large a volume for lab-scale work. To get a more reasonable volume, we should use a more concentrated solution!

What if we used 6.00 M HCl?

$$\frac{6.00 \text{ mol HCl}}{\text{L}}$$

$$0.657 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} = \boxed{0.110 \text{ L}}$$

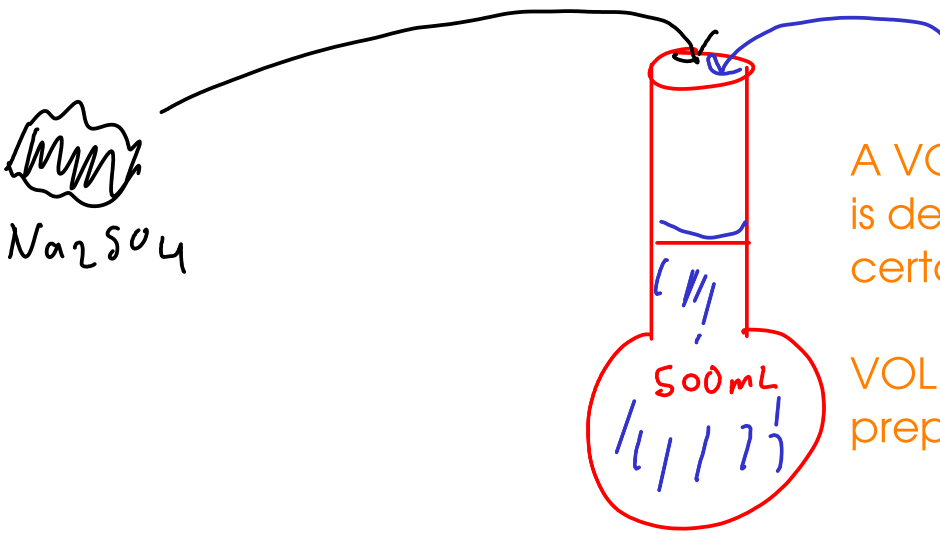
$11\bar{0}$ mL

110. mL is a more reasonable volume for lab-scale work. Easily measured with our lab equipment.

Example: How would we prepare 500. mL of 0.500 M sodium sulfate in water?

Na_2SO_4 : 142.05 g/mol

Dissolve the appropriate amount of sodium sulfate into enough water to make 500. mL of solution.



volumetric flask

A VOLUMETRIC FLASK is a flask that is designed to precisely contain a certain volume of liquid.

VOLUMETRIC FLASKS are used to prepare solutions.

We know that we need 500 mL of solution, and we also know that the concentration should be 0.500 moles of sodium sulfate per liter. How many moles of sodium sulfate should there be in 500 mL? Then, change to mass.

$0.500 \text{ mol } Na_2SO_4 = L$ | $mL = 10^{-3} L$ | $142.05 \text{ g } Na_2SO_4 = \text{mol } Na_2SO_4$

$$500.\text{ mL} \times \frac{10^{-3} L}{mL} \times \frac{0.500 \text{ mol } Na_2SO_4}{L} \times \frac{142.05 \text{ g } Na_2SO_4}{\text{mol } Na_2SO_4} = \boxed{35.5 \text{ g } Na_2SO_4}$$

So, to prepare the solution, put 35.5 grams sodium sulfate into a 500 mL volumetric flask, then fill the flask to the line with distilled (or deinonzed) water.

More on MOLARITY

To prepare a solution of a given molarity, you generally have two options:

1. Weigh out the appropriate amount of solute, then dilute to the desired volume with solvent (usually water)

"stock solution"

2. Take a previously prepared solution of known concentration and DILUTE it with solvent to form a new solution

- Use DILUTION EQUATION

The dilution equation is easy to derive with simple algebra.

$$M \times V$$

$$\frac{mol}{L} \times L = \text{moles solute}$$

... but when you dilute a solution, the number of moles of solute REMAINS CONSTANT. (After all, you're adding only SOLVENT)

$$M_1V_1 = M_2V_2$$

Since the number of moles of solute stays the same, this equality must be true!

before diution

after dilution

$$M_1V_1 = M_2V_2$$ ... the "DILUTION EQUATION"

M_1 = molarity of concentrated solution

V_1 = volume of concentrated solution

M_2 = molarity of dilute solution

V_2 = volume of dilute solution (total volume, not volume of added solvent!)

The volumes don't HAVE to be in liters, as long as you use the same volume UNIT for both volumes!

Example: Take the 0.500 M sodium sulfate we discussed in the previous example and dilute it to make 150. mL of 0.333 M solution. How many mL of the original solution will we need to dilute?

$$M_1V_1 = M_2V_2$$

$$(0.500\ M)V_1 = (0.333\ M)(150.\ mL)$$

$$V_1 = \boxed{99.9\ mL \text{ of } 0.500\ M\ Na_2SO_4}$$

To make the solution, measure out 99.9 mL of the 0.500 M stock solution into a 150. mL volumetric flask, then dilute to the mark with distilled water. (If no flask is available, you can do the same thing with a large graduated cylinder.)

CHEMICAL CALCULATIONS CONTINUED: REACTIONS

- Chemical reactions proceed on an ATOMIC basis, NOT a mass basis!

- To calculate with chemical reactions (i.e. use chemical equations), we need everything in terms of ATOMS ... which means MOLES of atoms

$$2\,Al(s) + 3\,Br_2(l) \longrightarrow 2\,AlBr_3(s)$$

coefficients are in terms of atoms and molecules!

$$\frac{2 \text{ atoms } Al = 3 \text{ molecules } Br_2 = 2 \text{ formula units } AlBr_3}{2 \text{ mol } Al = 3 \text{ mol } Br_2 = 2 \text{ mol } AlBr_3}$$

- To do chemical calculations, we need to:
  - Relate the amount of substance we know (mass or volume) to a number of moles
  - Relate the moles of one substance to the moles of another using the equation
  - Convert the moles of the new substance to mass or volume as desired

$$2\,Al(s) + 3\,Br_2(l) \longrightarrow 2\,AlBr_3(s)$$

* Given that we have 25.0 g of liquid bromine, how many grams of aluminum would we need to react away all of the bromine? How many grams of aluminum bromide would be produced?

1. Convert grams of bromine to moles: Need formula weight Br_2: $\frac{2 \times 79.90}{159.80}$

$159.80\,g\,Br_2 = 1\,mol\,Br_2$

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} = 0.15645\,mol\,Br_2$$

2. Use the chemical equation to relate moles of bromine to moles of aluminum

$2\,mol\,Al = 3\,mol\,Br_2$

$$0.15645\,mol\,Br_2 \times \frac{2\,mol\,Al}{3\,mol\,Br_2} = 0.10430\,mol\,Al$$

3. Convert moles aluminum to mass: Need formula weight Al: 26.98

$26.98\,g\,Al = 1\,mol\,Al$

$$0.10430\,mol\,Al \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$