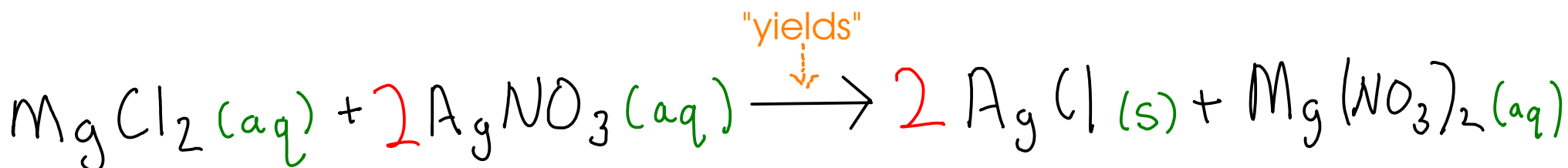


## CHEMICAL EQUATIONS

- are the "recipes" in chemistry
- show the substances going into a reaction, substances coming out of the reaction, and give other information about the process



REACTANTS - materials that are needed for a reaction

PRODUCTS - materials that are formed in a reaction

COEFFICIENTS - give the ratio of molecules/atoms of one substance to the others

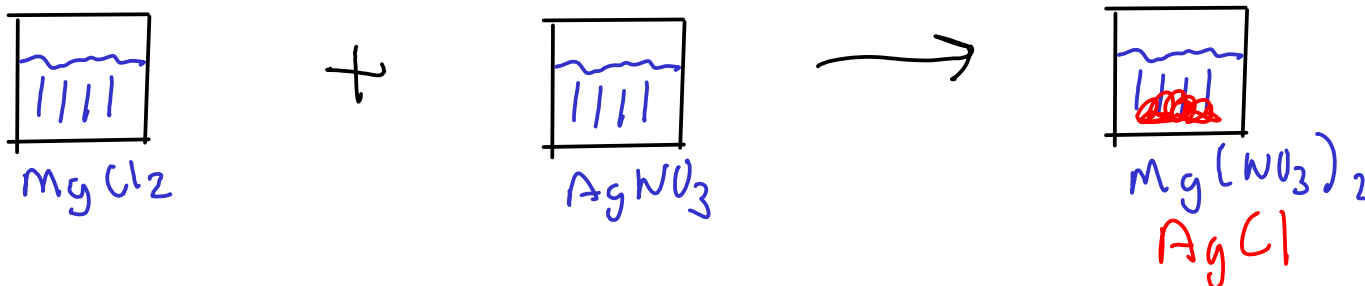
PHASE LABELS - give the physical state of a substance:

(s) - solid

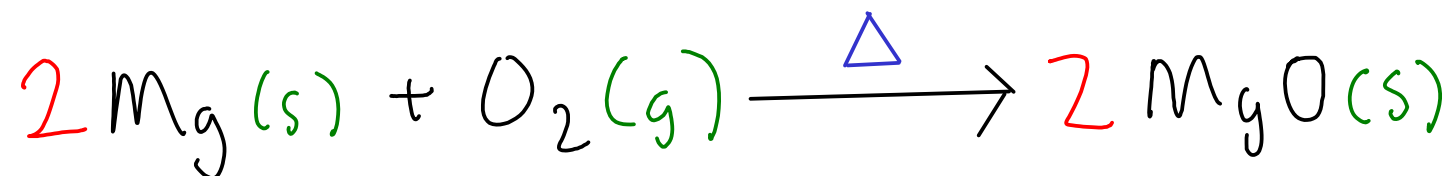
(l) - liquid

(g) - gas

(aq) - aqueous. In other words, dissolved in water



## CHEMICAL EQUATIONS



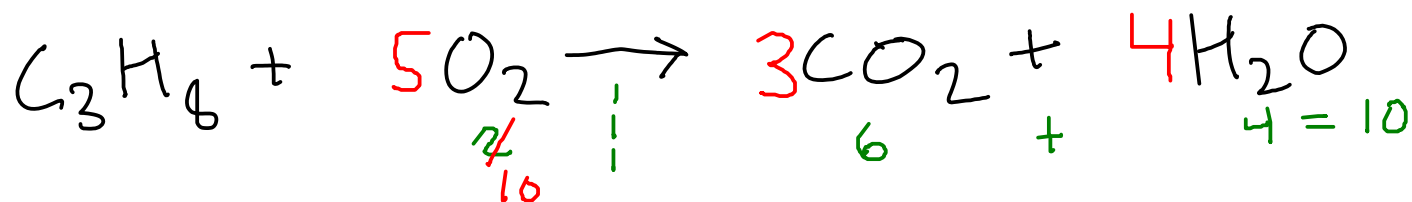
REACTION CONDITIONS - give conditions necessary for chemical reaction to occur. May be:

- $\Delta$  apply heat
  - catalysts - substances that will help reaction proceed faster
  - other conditions, such as required temperatures
- Reaction conditions are usually written above the arrow, but may also be written below if the reaction requires several steps or several different conditions

## COEFFICIENTS

- Experimentally, we can usually determine the reactants and products of a reaction
- We can determine the proper ratios of reactants and products WITHOUT further experiments, using a process called BALANCING
- BALANCING a chemical equation is making sure the same number of atoms of each element go into a reaction as come out of it.
- A properly balanced chemical equation has the smallest whole number ratio of reactants and products.
- There are several ways to do this, but we will use a modified trial-and-error procedure.

## BALANCING



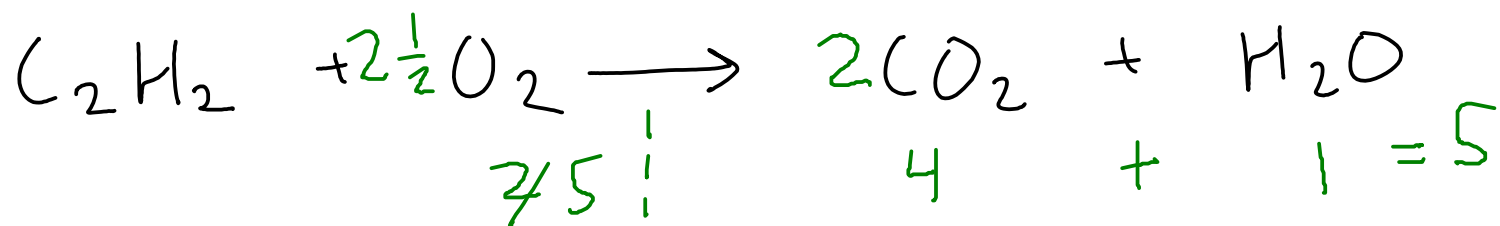
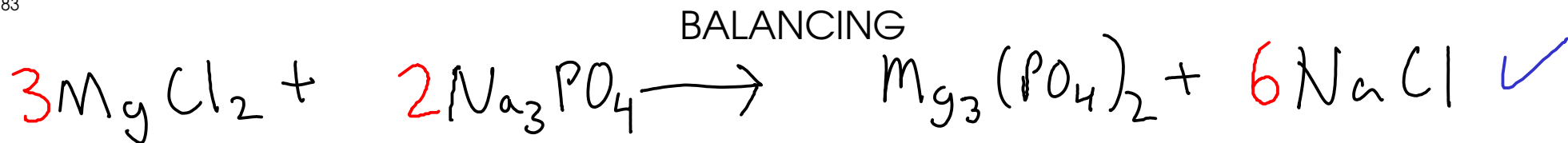
① Pick an element. Avoid (if possible) elements that appear in more than one substance on each side of the equation.

② Change the coefficients on substances containing this element so that the same number of atoms of the element are present on each side. **CHANGE AS LITTLE AS POSSIBLE!**

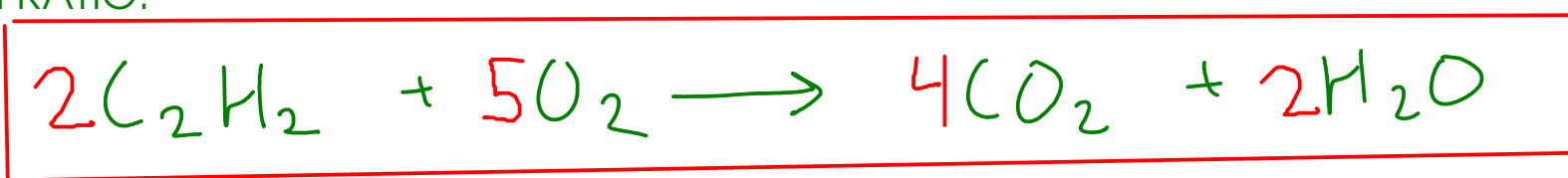
③ Repeat 1-2 until all elements are done.

④ Go back and quickly VERIFY that you have the same number of atoms of each element on each side. If you used any fractional coefficients, multiply each coefficient by the DENOMINATOR of your fraction.

**Use SMALLEST WHOLE NUMBER RATIOS!**

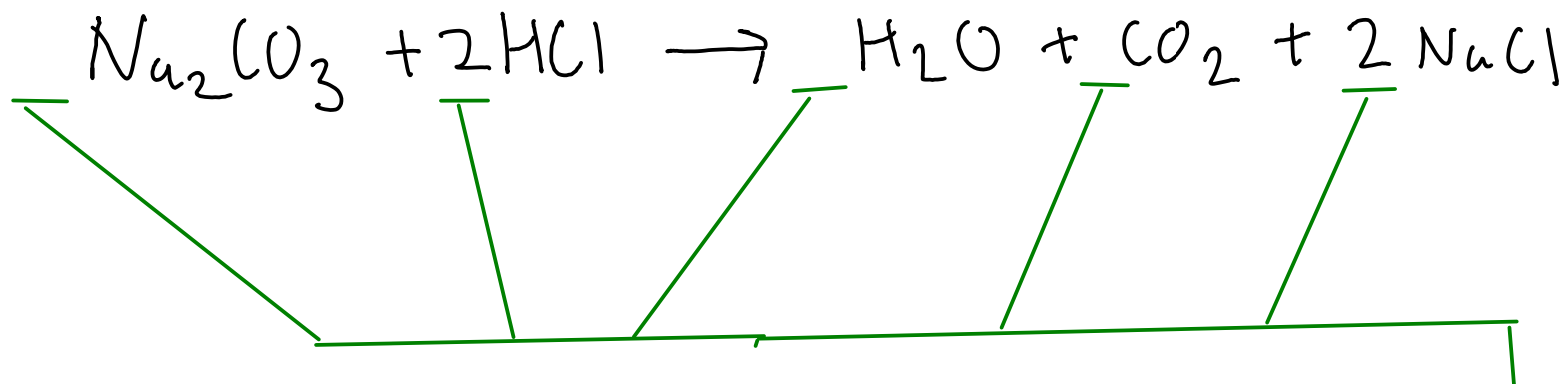


To make 5 oxygen atoms on the reactant side, we needed to use a coefficient of 2 1/2 for oxygen gas. To get rid of the fraction, we multiply ALL the coefficients by the denominator of the fraction (here, 2). We can do this because the coefficients form a RATIO.



- 1 - Avoid H, balance S. (H shows up twice on left)
- 2 - Avoid O, balance Na (O shows up in all four substances)
- 3 - Balance H (shows up less than O)
- 4 - Balance O (already balanced by previous changes!)

## CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS



Chemical equations are written and balanced in terms of ATOMS and MOLECULES

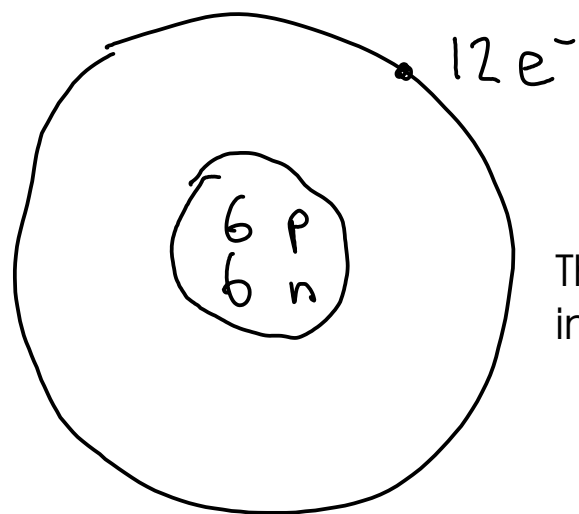
- While chemical equations are written in terms of ATOMS and MOLECULES, that's NOT how we often measure substances in lab!
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

## THE MOLE CONCEPT

- A "mole" of atoms is  $6.022 \times 10^{23}$  atoms

Why so big? Because atoms are so small!

- Why - in the metric dominated world of science - do we use such a strange number for quantity of atoms?



carbon-12

The mole is also defined as the number of carbon-12 atoms in exactly 12 g of carbon-12

## THE MOLE CONCEPT

- Why define the mole based on an experimentally-measured number?
- The atomic weight of an element (if you put the number in front of the unit GRAMS) is equal to the mass of ONE MOLE of atoms of that element!

Carbon (C): Atomic mass 12.01 ~~amu~~ → 12.01 g

↓

the mass of ONE MOLE of naturally-occurring carbon atoms

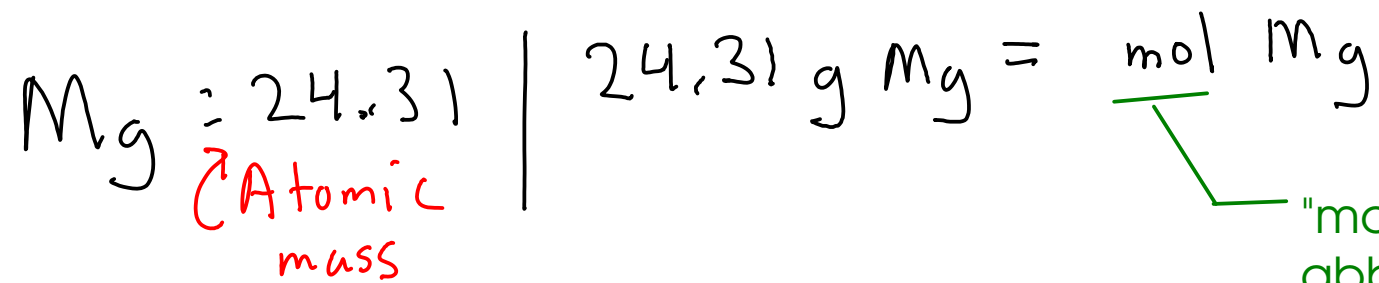
Magnesium (Mg): 24.31 g = the mass of ONE MOLE OF MAGNESIUM ATOMS

- So, using the MOLE, we can directly relate a mass and a certain number of atoms!



## 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.



"mol" is the abbreviation for "mole"

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

$$24.31 \text{ g Mg} = \text{mol Mg}$$

$$250. \text{ g Mg} \times \frac{\text{mol Mg}}{24.31 \text{ g Mg}} = 10.3 \text{ mol Mg}$$

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

$$\text{Fe: } 55.85$$

$$55.85 \text{ g Fe} = \text{mol Fe}$$

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

## WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

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

$$\text{H}_2\text{O}: \quad \text{H}: 2 \times 1.008 = 2.016$$

$$\text{O}: 1 \times 16.00 = \underline{16.00}$$

18.016 ← FORMULA WEIGHT of water

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

$$18.016 \text{ g H}_2\text{O} = 1 \text{ mol H}_2\text{O}$$

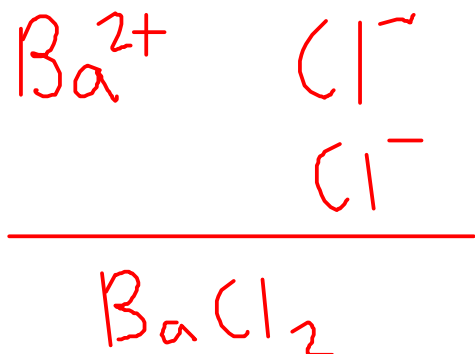
$$25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 1.39 \text{ mol H}_2\text{O}$$

Formula weight goes by several names:

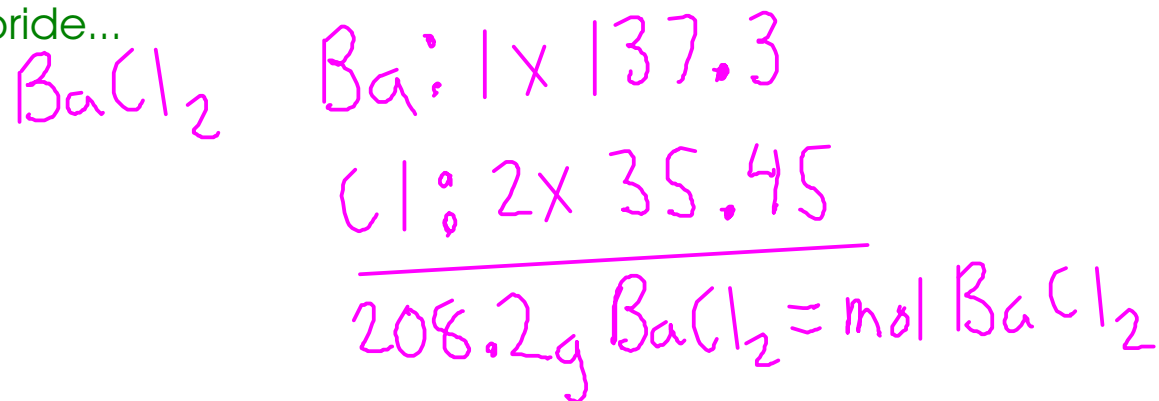
- For atoms, it's the same thing as ATOMIC WEIGHT
- For molecules, it's called MOLECULAR WEIGHT
- Also called "MOLAR MASS"

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

First, figure out the formula of barium chloride. It's an ionic compound (barium is a metal!)



Second, calculate the formula weight of barium chloride...



Finally, find the mass barium chloride...

$$3.65 \text{ mol BaCl}_2 \times \frac{208.2 \text{ g BaCl}_2}{\text{mol BaCl}_2} = \boxed{760. \text{ g BaCl}_2}$$

(760 g)

## 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 barium chloride.

$$\text{BaCl}_2 : \text{Ba} : 1 \times 137.3 = 137.3$$

$$\text{Cl} : 2 \times 35.45 = 70.90$$

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

$$208.2 \text{ g BaCl}_2 = \text{mol BaCl}_2$$

$$\% \text{Ba} : \frac{137.3 \text{ g Ba}}{208.2 \text{ g BaCl}_2} \times 100 = 65.95 \% \text{ Ba}$$

$$\% \text{Cl} : \frac{70.90 \text{ g Cl}}{208.2 \text{ g BaCl}_2} \times 100 = 34.05 \% \text{ Cl}$$

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

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

↖ dissolved substance

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

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

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

$$0.250 \text{ L} \times \frac{6.0 \text{ mol HCl}}{\text{L}} = \boxed{1.50 \text{ mol HCl}}$$

\*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?

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

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

(11800 mL)

What if we used 6.00 M HCl?

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

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

(110 mL)

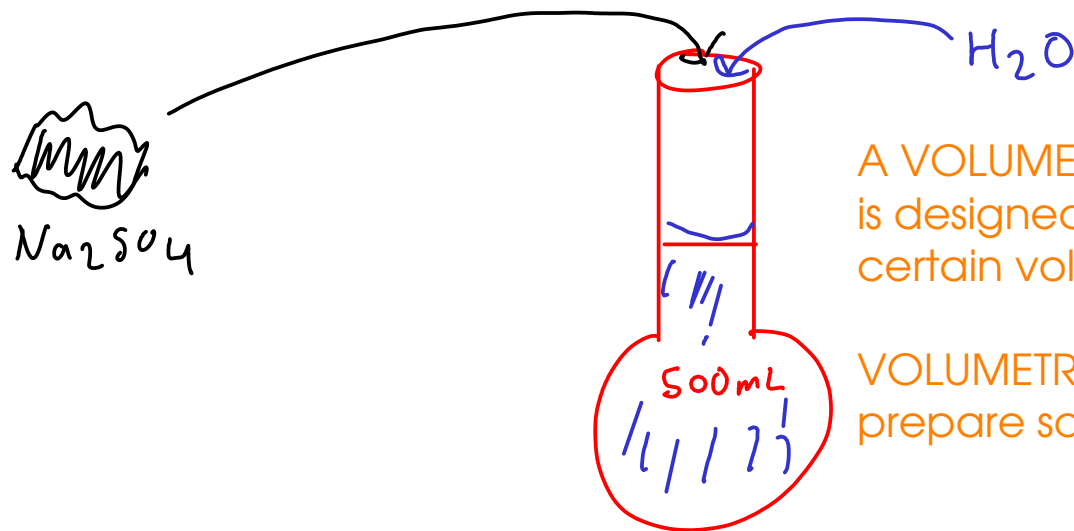
In this case, we'd probably use the solution on the bottom (the 6.00 M HCl) in the lab if we needed 0.657 moles HCl, as it's not likely we even HAVE 11800 mL of the first solution!



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



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



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

VOLUMETRIC FLASKS are used to prepare solutions.

volumetric flask

From the volume and concentration, calculate the moles sodium sulfate. (Use MOLARITY.)  
Then, convert the moles sodium sulfate to mass. (Use FORMULA WEIGHT.)

$$\textcircled{1} 0.500 \text{ mol Na}_2\text{SO}_4 = L \quad \text{mL} = 10^{-3} L$$

$$500 \text{ mL} \times \frac{10^{-3} L}{\text{mL}} \times \frac{0.500 \text{ mol Na}_2\text{SO}_4}{L} = 0.250 \text{ mol Na}_2\text{SO}_4$$

$$\textcircled{2} 142.05 \text{ g Na}_2\text{SO}_4 = \text{mol Na}_2\text{SO}_4$$

$$0.250 \text{ mol Na}_2\text{SO}_4 \times \frac{142.05 \text{ g Na}_2\text{SO}_4}{\text{mol Na}_2\text{SO}_4} = 35.5 \text{ g Na}_2\text{SO}_4$$

Put 35.5 grams sodium sulfate into the flask, then add enough water to get a volume of 500. mL ...

## 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)

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

"stock solution"

- Use DILUTION EQUATION

The dilution equation is easy to derive with simple algebra.

$$M \times V$$

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

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

$$M_1 V_1 = M_2 V_2$$

before  
dilution

after  
dilution

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

$$M_1 V_1 = M_2 V_2 \quad \dots \text{the "DILUTION EQUATION"}$$

$M_1$  = 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?

$$M_1 V_1 = M_2 V_2 \quad M_1 = 0.500 \text{ M} \quad M_2 = 0.333 \text{ M}$$

$$V_1 = ? \quad V_2 = 150. \text{ mL}$$

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

$$V_1 = 99.9 \text{ mL of } 0.500 \text{ M } \text{Na}_2\text{SO}_4$$

To make the solution, take 99.9 mL of 0.500 M sodium sulfate, then add water until the total volume is 150. mL ...