BALANCING

$$3MgCl_2 + 2Na_3PO_4 \longrightarrow Mg_3(PO_4)_2 + 6NaCl$$

$$C_2H_2 + \frac{5}{2}O_2 \longrightarrow 2CO_2 + H_2O$$

$\cancel{2}$ 5   4   1 $\overline{5}$

... we used 5/2 as the coefficient of oxygen to balance the equation, BUT the equation needs WHOLE NUMBER coefficients, not fractions. Multiply all coefficients by the DENOMINATOR of the fraction (2).

$$\boxed{2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O}$$

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

* Start with sulfur, since hydrogen appears in two compounds on the left.
* Then balance sodium, since oxygen appears in AL:L of these compounds.
* Next, balance hydrogen. It's easier than oxygen (it shows up in fewer compounds)
* Finally, balance oxygen. (It's already balanced)

CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS

$$\_\, Na_2CO_3 + \underline{2}\,HCl \rightarrow \_\, H_2O + \_\, CO_2 + \underline{2}\,NaCl$$

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!

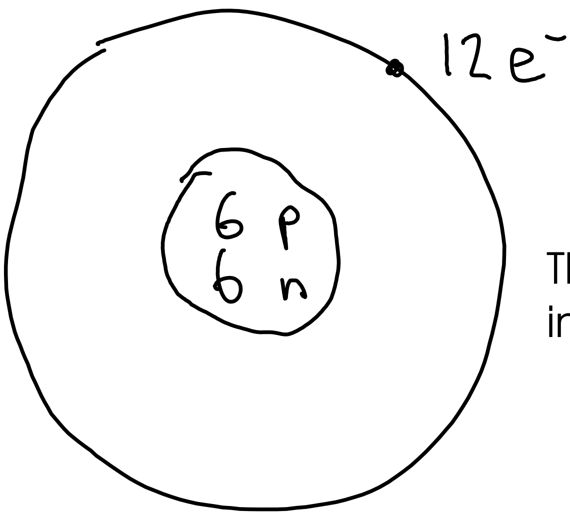
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

THE MOLE CONCEPT

- A "mole" of atoms is 6.022×10^{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~~ → $\boxed{12.01\text{ g}}$

↓

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.

Mg : 24.31 | 24.31 g Mg = 1 mol Mg

Atomic mass

"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

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

* Note: Atomic weights DO have significant figures; they are measured numbers, not exact numbers!

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

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

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

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

FORMULA WEIGHT is the 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 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

After we find the FORMULA, we can find the FORMULA WEIGHT

$$96.094 \text{ g } (NH_4)_2CO_3 = \text{mol } (NH_4)_2CO_3$$

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

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 : $3 \times 16.00 = 48.00$

$80.052 \text{ g } NH_4NO_3 = 1 \text{ mol } NH_4NO_3$

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

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

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

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

These should sum to approximately 100% (within roundoff error)

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

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)

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

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

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

11800 mL

This is too large of a volume for typical lab-scale work, so we should pick a more concentrated solution for this experiment.

What if we used 6.00 M HCl?

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

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

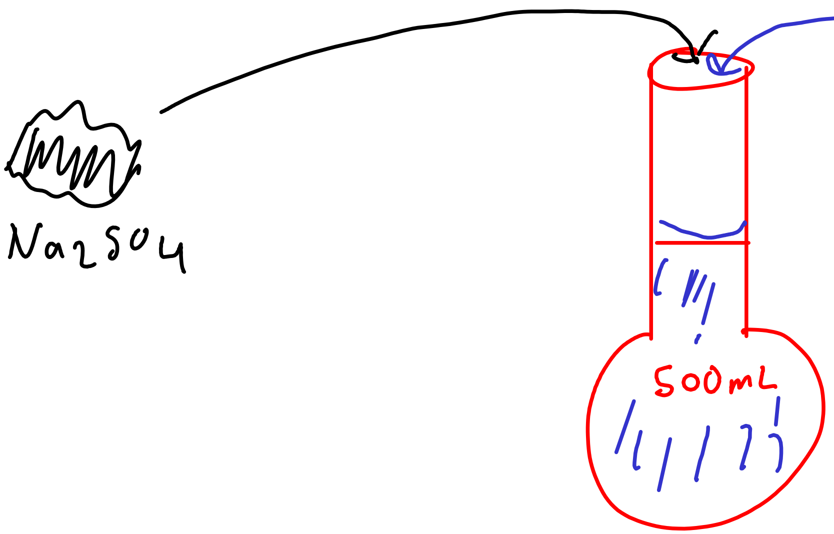
110 mL

110 mL is a reasonable lab volume
(Use a 250 mL cylinder)

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.

Want 0.500 M, calculate moles of Na_2SO_4 needed, then convert to g.

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

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

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.

$$M \times V$$

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

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

$$M_1V_1 = M_2V_2$$

before diution

after dilution

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

$$M_1V_1 = M_2V_2$$ ... 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

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?

$M_1 = 0.500\ M \qquad M_2 = 0.333\ M$

$V_1 = ?\ mL \qquad V_2 = 150.\ mL$

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

$$V_1 = 99.9\ mL$$

So, to make the new solution, put 99.9 mL of 0.500 M sodium sulfate solution into a 150 mL volumetric flask, then fill to 150 mL with water.

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

$$\underline{2}Al(s) + \underline{3}Br_2(l) \longrightarrow \underline{2}AlBr_3(s)$$

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

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

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

You can combine all three steps on one line if you like!

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} \times \frac{2\,mol\,Al}{3\,mol\,Br_2} \times \frac{26.98\,g\,Al}{1\,mol\,Al} = \boxed{2.81\,g\,Al}$$

1 2 3

You can solve the second part of the question using CONSERVATION OF MASS - since there's only a single product and you already know the mass of all reactants.

$$\begin{array}{r} 25.0\,g\,Br_2 \\ +\ 2.81\,g\,Al \\ \hline 27.8\,g\,AlBr_3 \end{array}$$

But ...

...what would you have done to calculate the mass of aluminum bromide IF you had NOT been asked to calculate the mass of aluminum FIRST?

$$25.0\,g\,Br_2 \times \frac{1\,mol\,Br_2}{159.80\,g\,Br_2} \times \frac{2\,mol\,AlBr_3}{3\,mol\,Br_2} \times \frac{266.694\,g\,AlBr_3}{1\,mol\,AlBr_3} = 27.8\,g\,AlBr_3$$

1 convert mass bromine to moles

2 convert moles bromine to moles aluminum bromide

3 convert moles aluminum bromide to mass