BALANCING

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

$$C_2H_2 + 2\tfrac{1}{2}O_2 \longrightarrow 2CO_2 + H_2O$$

$\cancel{2}$ 5   4   1

$\underbrace{4 \qquad\qquad 1}_{5}$

We used a coefficient of 2 1/2 (or 5/2) for our coefficient in front of O2. That gives us the right number of oxygen atoms going in, but we're supposed to use WHOLE NUMBERS for coefficients! We'll get whole numbers by multiplying EVERY coefficient by the denominator of the fraction (2). We can do that since the coefficients are the RATIO of one substance to another...

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

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

1) Avoid H, start with S, since H shows up twice on the left.
2) Avoid O, balance Na next - since O shows up in all four compounds!
3) Balance H, since it shows up less than O.
4) Balance O. (It's already done!)

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!

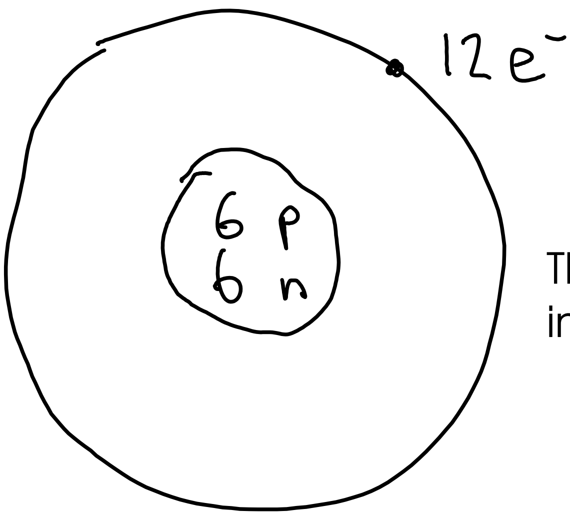
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

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

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

ATOMIC WEIGHT is a MEASURED number - in other words, it has significant figures. Usually we can find atomic weights with more significant figures if necessary.

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

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!

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

NH_4^+

$(NH_4)_2CO_3$

Now, calculate 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 g $(NH_4)_2CO_3$ = 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 g NH_4NO_3 = 1 mol NH_4NO_3

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

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

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

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

Check: All these should sum to 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 - 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?

$0.0555 \text{ mol HCl} = L$

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

11800 mL

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

What if we used 6.00 M HCl?

$6.00 \text{ mol HCl} = L$

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

$11\bar{0}$ mL

110 mL is a reasonable volume for us to ,measure out using common lab equipment (like our graduated cylinders)