BALANCING

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

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

~~4~~ 5

$$\frac{4 \qquad\qquad 1}{5}$$

We used 2.5 (2 1/2) for the coefficient of the oxygen, BUT we need WHOLE NUMBERS to have a properly balanced equation. We've accounted for all atoms, but we can't do the reaction with half of an oxygen molecule. To get rid of the fraction, we multiply ALL COEFFICENTS by the denominator of the fraction.

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

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

1) Avoid H, balance S (no changes necessary for S). H appears twice on left.
2) Avoid O, balance Na. O appears in ALL substances here.
3) Balance H, since it shows up fewer times 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!

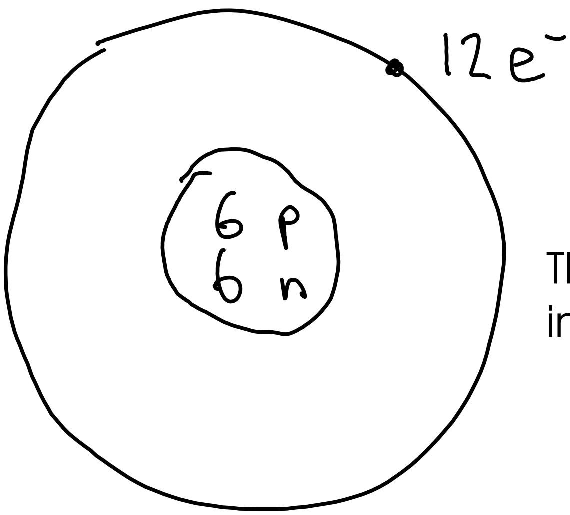
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

THE MOLE CONCEPT

- A "mole" of atoms is 6.022 $\times 10^{23}$ 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~~ → $\boxed{12.01\ g}$

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.

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

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

Atomic weight is a measured number - in other words, it has significant figures. Usually, though, it's not the atomic weight that limits us - you can often look up atomic weights that are known to a higher precision 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^+

$(NH_4)_2CO_3$

Next, calculte 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

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

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

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 x 14.01 = 28.02

H: 4 x 1.008 = 4.032

O: 3 x 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!

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

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

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

Check: All these percentages 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?

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

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

11,800 mL

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

What if we used 6.00 M HCl?

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

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

110. mL

110. mL is a much more reasonable volume to measure in the lab. (Use a 250 mL cylinder, or a 100 mL cylinder twice)