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

$$C_3H_8 + 5\,O_2 \rightarrow 3CO_2 + 4H_2O$$

$$\underbrace{6 \qquad 4}_{10}$$

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 DENOMIMATOR of your fraction.

Use SMALLEST WHOLE NUMBER RATIOS!

BALANCING

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

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

2

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

To get rid of the fractional coefficient, multiply ALL coefficients by the denominator (in this case, 2) of the fraction:

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

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

* Start with S, since H appears in two compounds on the left
* Next, do Na (O appears in ALL compounds!)
* Then, balance H (easier than O)
* Finally, balance O (it's already done!)

CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS

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

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!

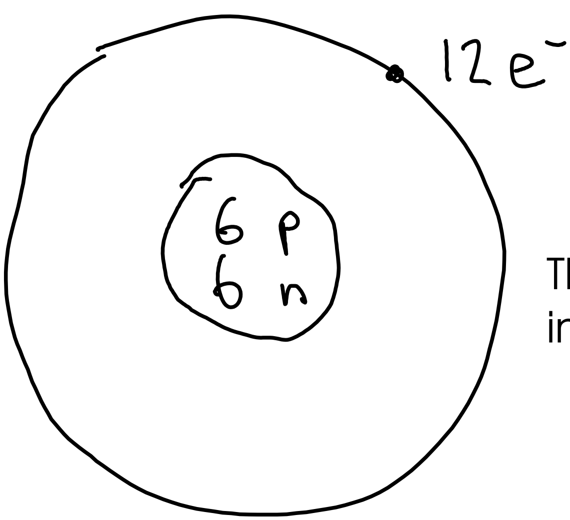
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

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 a~~m~~u → $\boxed{12.01\text{ g}}$

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

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

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

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

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

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?

Translate "ammonium carbonate" to a chemical formula

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

NH_4^+

$(NH_4)_2CO_3$

N: 2 × 14.01

H: 8 × 1.008

O: 3 × 16.00

C: 1 × 12.01

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 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!

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

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

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

These should sum to approximately 100%

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?

$0.0555\ mol\ HCl = L$

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

(11,800 mL)

This is not a reasonable volume for lab scale work. We should use a more concentrated solution if we need to dispense 0.657 moles!

What if we used 6.00 M HCl?

$6.00\ mol\ HCl = L$

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

($11\bar{0}$ mL)

This amount is more reasonable. We probably have at least 110 mL on hand, and we can easily measure this volume with a standard 250 mL cylinder.