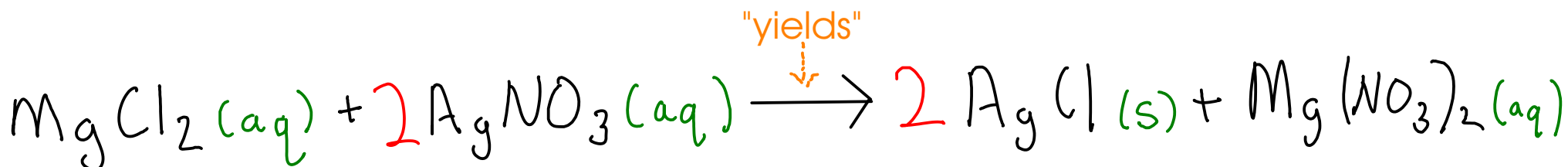


## CHEMICAL EQUATIONS

- 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

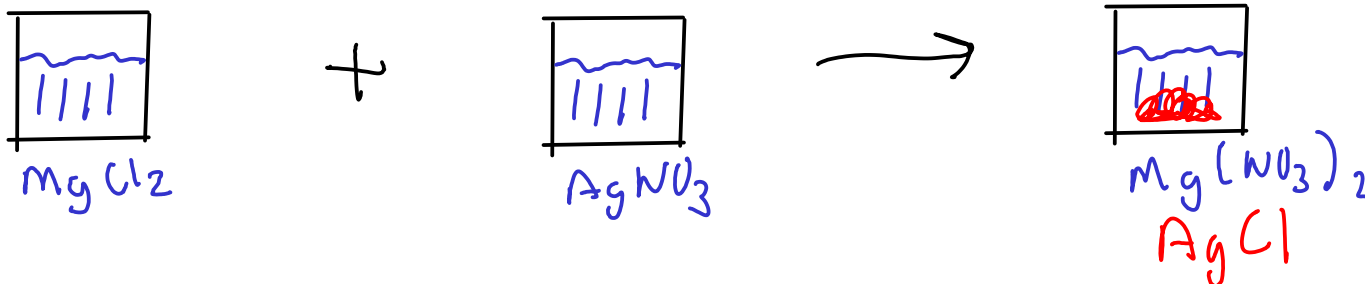
PHASE LABELS - give the physical state of a substance:

(s) - solid

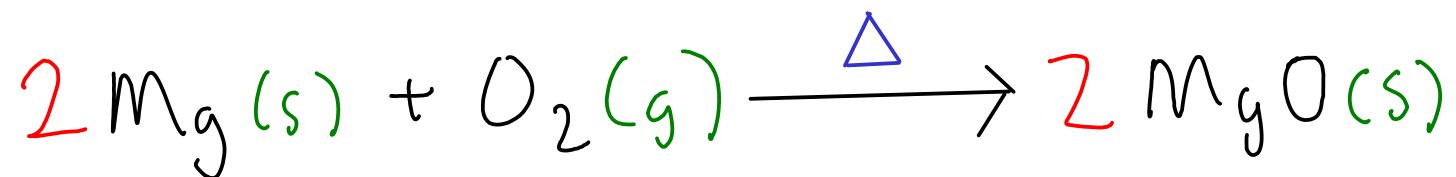
(l) - liquid

(g) - gas

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



## CHEMICAL EQUATIONS



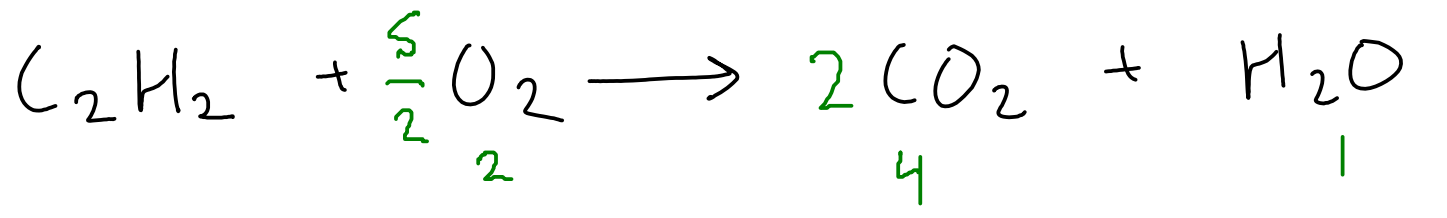
REACTION CONDITIONS - give conditions necessary for chemical reaction to occur. May be:

- $\Delta$  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

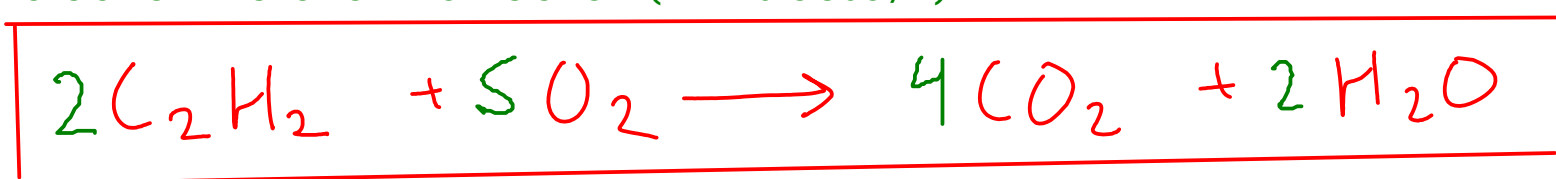
## COEFFICIENTS

- 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.



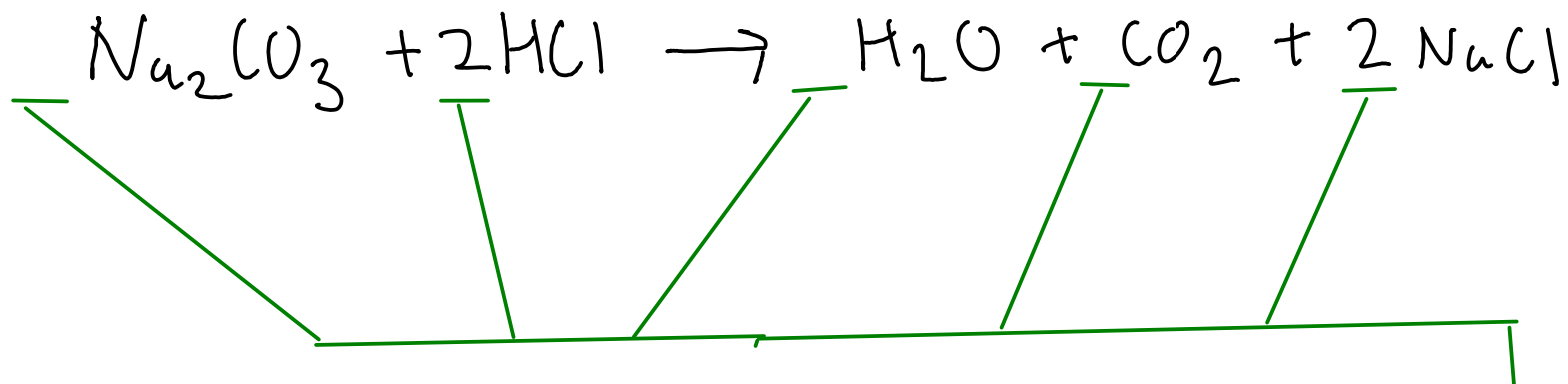


... we used  $5/2$  ( $2 \frac{1}{2}$ ) as our coefficient for oxygen on the left, but we need to use WHOLE NUMBER coefficients to balance properly. Multiply ALL the coefficients by the denominator of the fraction (in this case, 2).



- 1) Avoid H, and balance S first - since H appears twice on the left.
- 2) Avoid O and balance Na - since O appears in all four substances
- 3) Balance H, since it shows up fewer times than O.
- 4) Balance O - and it's already done!

## CHEMICAL CALCULATIONS - RELATING MASS AND ATOMS



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

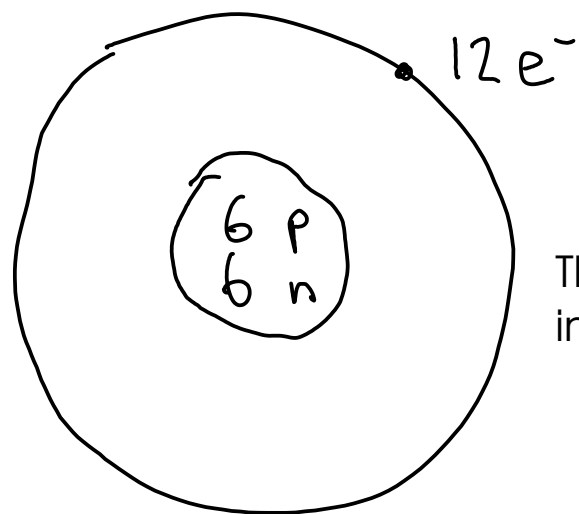
- While chemical equations are written in terms of ATOMS and MOLECULES, that's NOT how we often measure substances in lab!
- measurements are usually MASS (and sometimes VOLUME), NOT number of atoms or molecules!

## 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 ~~amu~~ → 12.01 g

↓  
the mass of ONE MOLE of  
naturally-occurring carbon atoms

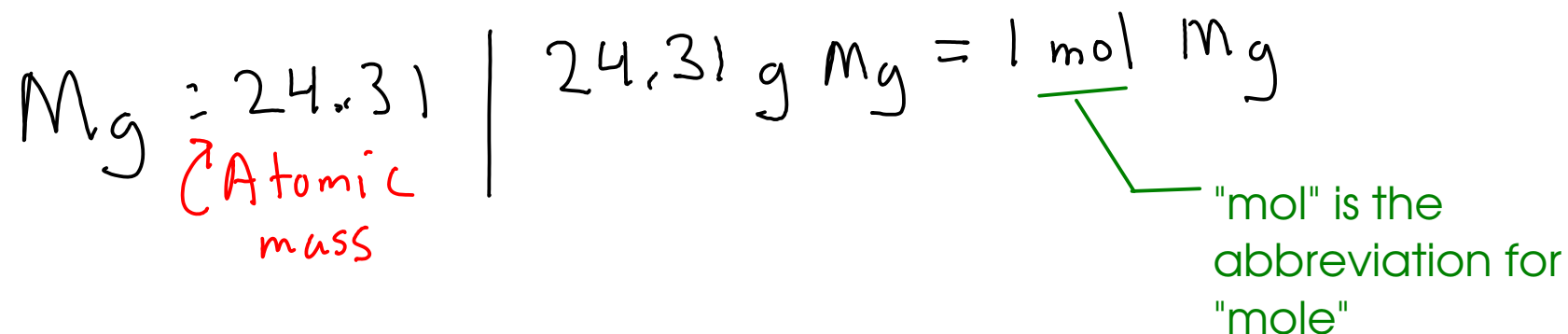
Magnesium (Mg): 24.31 g = the mass of ONE MOLE OF MAGNESIUM ATOMS

- 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.



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

$$24.31 \text{ g Mg} = 1 \text{ mol Mg}$$

$$250. \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 10.3 \text{ mol Mg}$$

Atomic weight is a MEASURED number. In other words, atomic weight has significant figures. Usually, we can find more precise atomic weights if necessary, so it doesn't usually limit us in calculations..

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

$$55.85 \text{ g Fe} = 1 \text{ mol Fe}$$

$$1.75 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = \boxed{97.7 \text{ g Fe}}$$

## WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

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

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

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

18.016 ← FORMULA WEIGHT of water

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

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

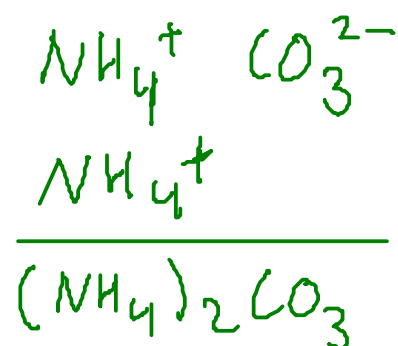
$$25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 1.39 \text{ mol H}_2\text{O}$$

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 ammonium carbonate do we need to weigh out to get 3.65 moles of ammonium carbonate?

First, we need the FORMULA!



Now, find FORMULA WEIGHT

$$\text{N} : 2 \times 14.01$$

$$\text{H} : 8 \times 1.008$$

$$\text{C} : 1 \times 12.01$$

$$\text{O} : 3 \times 16.00$$

$$\underline{96.094}$$

$$96.094 \text{ g } (\text{NH}_4)_2\text{CO}_3 = 1 \text{ mol } (\text{NH}_4)_2\text{CO}_3$$

$$3.65 \text{ mol } (\text{NH}_4)_2\text{CO}_3 \times \frac{96.094 \text{ g } (\text{NH}_4)_2\text{CO}_3}{1 \text{ mol } (\text{NH}_4)_2\text{CO}_3} = \boxed{351 \text{ g } (\text{NH}_4)_2\text{CO}_3}$$

## 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.

$$\text{NH}_4\text{NO}_3 : \text{N} : 2 \times 14.01 = 28.02$$

$$\text{H} : 4 \times 1.008 = 4.032$$

$$\text{O} : 3 \times 16.00 = 48.00$$

$$\underline{80.052 \text{ g NH}_4\text{NO}_3 = 1 \text{ mol NH}_4\text{NO}_3}$$

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

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

$$35.0\% \text{ N}$$

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

$$5.0\% \text{ H}$$

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

$$60.0\% \text{ O}$$

CHECK: All percentages should sum up 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?