

## 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~~ → 12.01 g

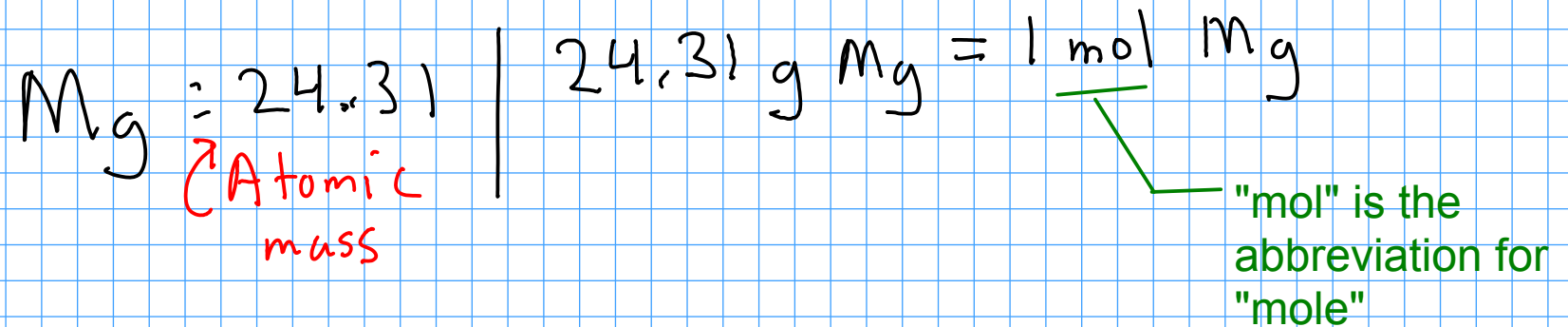
↓  
the mass of ONE MOLE of  
naturally-occurring carbon atoms

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?

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

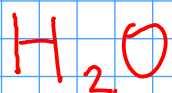
Example: You need 1.75 moles of iron. What mass of iron do you need to weigh out on the balance?

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

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

## WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

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



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

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

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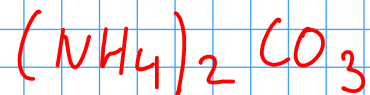
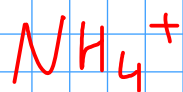
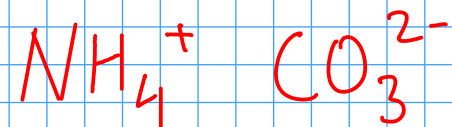
$$18.016 \quad | \text{— formula weight of water!}$$

Formula weight = mass of one mole of an element OR compound

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?



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

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

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

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

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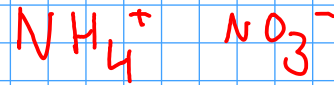
$$96.094 \text{ g} = 1 \text{ mol}$$

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

## 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.



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

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

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

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$$80.052 \text{ g NH}_4\text{NO}_3 \text{ per mole}$$

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

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

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

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

$$M = \text{molarity} = \frac{\text{moles of solute}}{\text{L solution}}$$

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

If you have 0.250 L (250 mL) of 6.0 M HCl, how many moles of HCl do you have?

$$6.0 \text{ mol HCl} = 1 \text{ L solution}$$

$$0.250 \cancel{\text{L}} \times \frac{6.0 \text{ mol HCl}}{1 \cancel{\text{L solution}}} = 1.5 \text{ mol HCl}$$



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

$$0.0555 \text{ mol HCl} = 1 \text{ L}$$

$$0.657 \text{ mol HCl} \times \frac{1 \text{ L}}{0.0555 \text{ mol HCl}} = \boxed{11.8 \text{ L}}$$

What if we used 6.00 M HCl?

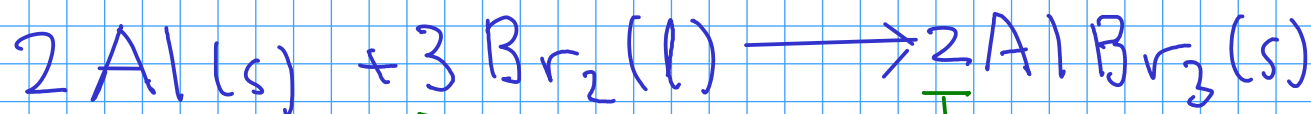
$$6.00 \text{ mol HCl} = 1 \text{ L}$$

$$0.657 \text{ mol HCl} \times \frac{1 \text{ L}}{6.00 \text{ mol HCl}} = \boxed{0.110 \text{ L}}$$

(110. mL)

## 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

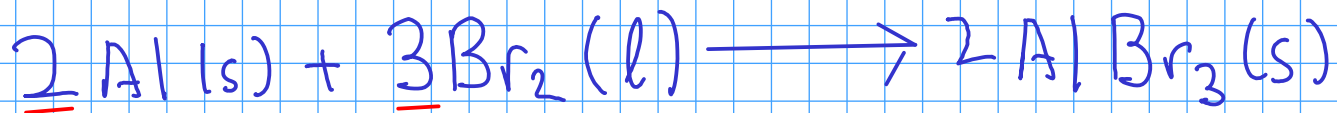


coefficients are in terms of atoms and molecules!

2 atoms Al = 3 molecules Br<sub>2</sub> = 2 formula units AlBr<sub>3</sub>

2 mol Al = 3 mol Br<sub>2</sub> = 2 mol AlBr<sub>3</sub>

- 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



\* 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  $\text{Br}_2: \frac{2 \times 79.90}{159.80}$

$$159.80 \text{ g Br}_2 = 1 \text{ mol Br}_2$$
$$25.0 \text{ g Br}_2 \times \frac{1 \text{ mol Br}_2}{159.80 \text{ g Br}_2} = 0.15645 \text{ mol Br}_2$$

② Use the chemical equation to relate moles of bromine to moles of aluminum

$$2 \text{ mol Al} = 3 \text{ mol Br}_2$$
$$0.15645 \text{ mol Br}_2 \times \frac{2 \text{ mol Al}}{3 \text{ mol Br}_2} = 0.10430 \text{ mol Al}$$

③ Convert moles aluminum to mass: Need formula weight  $\text{Al}: 26.98$

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

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

$$25.0 \text{ g Br}_2 \times \frac{1 \text{ mol Br}_2}{159.80 \text{ g Br}_2} \times \frac{2 \text{ mol Al}}{3 \text{ mol Br}_2} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 2.81 \text{ g Al}$$

①                      ②                      ③

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$$\begin{array}{r} 25.0 \text{ g Br}_2 \\ + 2.81 \text{ g Al} \\ \hline 27.8 \text{ g AlBr}_3 \end{array}$$

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?

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$$25.0 \text{ g Br}_2 \times \frac{1 \text{ mol Br}_2}{159.80 \text{ g Br}_2} \times \frac{2 \text{ mol AlBr}_3}{3 \text{ mol Br}_2} \times \frac{266.694 \text{ g AlBr}_3}{1 \text{ mol AlBr}_3} = 27.8 \text{ g}$$

①

convert mass  
bromine  
to moles

②

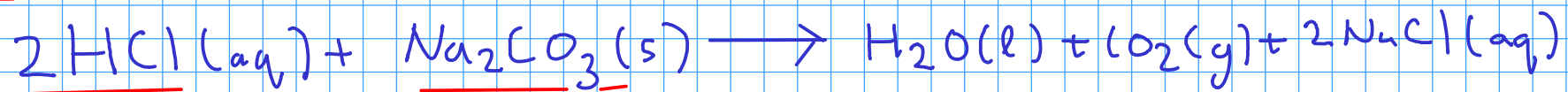
convert moles  
bromine to  
moles aluminum  
bromide

③

convert moles  
aluminum  
bromide  
to mass

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?



- 1 - Convert mass of sodium carbonate to moles using formula weight
  - 2 - Convert moles of sodium carbonate to moles hydrochloric acid using chemical equation
  - 3 - Convert moles of hydrochloric acid to volume using concentration (M = moles/L)
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- Convert mass of sodium carbonate to moles using formula weight

$$\begin{array}{l} \text{Na}_2\text{CO}_3: \text{Na} : 2 \times 22,99 \\ \quad \quad \quad \text{C} : 1 \times 12,01 \\ \quad \quad \quad \text{O} : 3 \times 16,00 \\ \hline \quad \quad \quad 105,99 \end{array}$$

$$105,99 \text{ g Na}_2\text{CO}_3 = 1 \text{ mol Na}_2\text{CO}_3$$

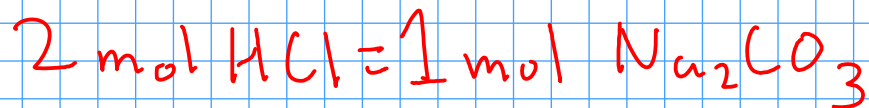
$$25,0 \text{ g Na}_2\text{CO}_3 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{105,99 \text{ g Na}_2\text{CO}_3} = 0,23587 \text{ mol Na}_2\text{CO}_3$$

Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?



- Convert moles of sodium carbonate to moles hydrochloric acid using chemical equation



$$0,235871 \text{ mol Na}_2\text{CO}_3 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Na}_2\text{CO}_3} = 0,471743 \text{ mol HCl}$$

- Convert moles of hydrochloric acid to volume using concentration (M = moles/L)

$$6.00 \text{ M HCl}: \quad 6.00 \text{ mol HCl} = 1 \text{ L}$$

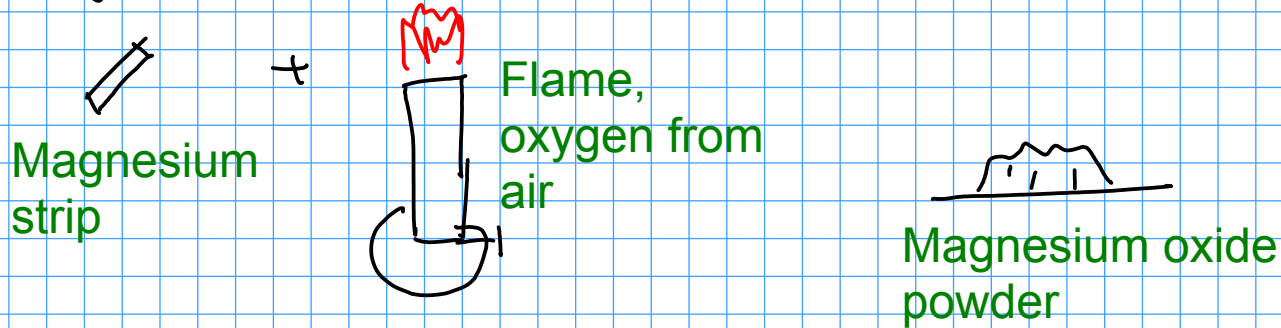
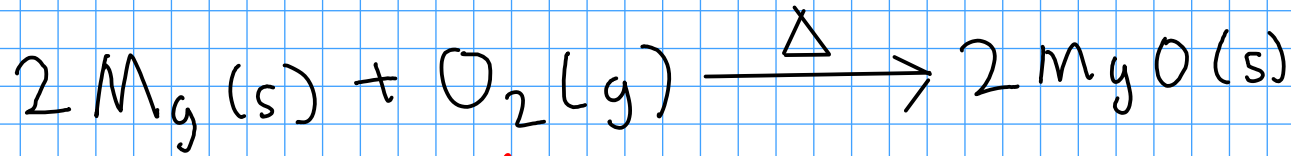
$$0,471743 \text{ mol HCl} \times \frac{1 \text{ L}}{6.00 \text{ mol HCl}} = 0,0786 \text{ L of solution}$$

$\text{mL} = 10^{-3} \text{ L}$  Convert liters to milliliters!

$$0,0786 \text{ L of solution} \times \frac{\text{mL}}{10^{-3} \text{ L}} = \boxed{78,6 \text{ mL solution}}$$

# CONCEPT OF LIMITING REACTANT

- When does a chemical reaction STOP?



- When does this reaction stop? When burned in open air, this reaction stops when all the MAGNESIUM STRIP is gone. We say that the magnesium is LIMITING.

- This reaction is controlled by the amount of available magnesium

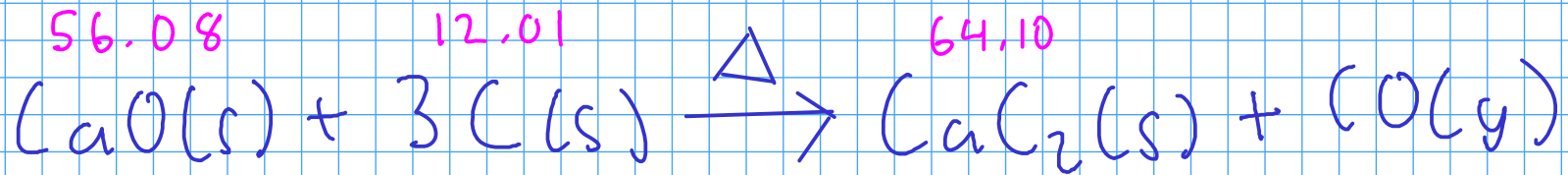
- At the end of a chemical reaction, the LIMITING REACTANT will be completely consumed, but there may be amount of OTHER reactants remaining. We do chemical calculations in part to minimize these "leftovers".

These are often called "excess" reactants, or reactants present "in excess"

## LIMITING REACTANT CALCULATIONS

- To find the limiting reactant, calculate how much product would be produced from ALL given reactants. Whichever produces the SMALLEST amount of product is the limiting reactant, and the smallest amount of product is the actual amount of product produced.

Example:



If you start with 100. g of each reactant, how much calcium carbide would be produced?

$$56.08 \text{ g CaO} = 1 \text{ mol CaO} \quad 64.10 \text{ g CaC}_2 = 1 \text{ mol CaC}_2$$

$$1 \text{ mol CaO} = 1 \text{ mol CaC}_2$$

$$100 \cancel{\text{g CaO}} \times \frac{1 \cancel{\text{mol CaO}}}{56.08 \cancel{\text{g CaO}}} \times \frac{1 \cancel{\text{mol CaC}_2}}{1 \cancel{\text{mol CaO}}} \times \frac{64.10 \text{ g CaC}_2}{1 \cancel{\text{mol CaC}_2}} = 114 \text{ g CaC}_2$$

$$12.01 \text{ g C} = 1 \text{ mol C} \quad 64.10 \text{ g CaC}_2 = 1 \text{ mol CaC}_2 \quad 3 \text{ mol C} = 1 \text{ mol CaC}_2$$

$$100 \cancel{\text{g C}} \times \frac{1 \cancel{\text{mol C}}}{12.01 \cancel{\text{g C}}} \times \frac{1 \cancel{\text{mol CaC}_2}}{3 \cancel{\text{mol C}}} \times \frac{64.10 \text{ g CaC}_2}{1 \cancel{\text{mol CaC}_2}} = 178 \text{ g CaC}_2$$

114g of calcium carbide are produced. Calcium oxide is limiting, and the reaction stops when all 100g of calcium carbide are gone! (There will be leftover carbon at the end of this reaction!)