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CHEMICAL QUANTITIES

Chapter 9

9.1 Information Given by Chemical Equations

Objective: To understand the molecular ad mass information given in a balanced equation

- Reactions are described by equations that give the identities o the reactants and products and show how much of each reactant and product participates in the reaction
- The coefficients tell us how much product we can get from a certain quantity of reactants
- Example: 2 pieces bread + 3 slices meat + 1 slice cheese→ 1 sandwich
  If you need 50 sandwiches:
  50 (2 bread) + 50 (3 meat) + 50 (1 cheese) → 50 sandwiches
  That is: 100 bread + 150 meat + 50 cheese → 50 sandwiches
  Notice that 100:150:50 is the same ratio as 2:3:1
- For a chemical equation:
  Unbalanced: CO (g) + H_2 (g) → CH_3OH (l)
  To balance the equation you choose coefficients that give the same number of each type of atom on both sides
  Balanced: CO (g) + $2H_2$(g) → CH_3OH (l)
- The coefficients give a *relative* number of molecules… you could multiply the equation by a number and still have a balanced equation

CO (g) +	$2H_2$(g) →	CH_3OH (l)
1 molecule CO	2 molecules H_2	1 molecule CH_3OH
1 dozen CO molecules	2 dozen H_2 molecules	1 dozen CH_3OH molecules
1 mol CO molecules	2 mol H_2 molecules	1 mol CH_3OH molecules

9.2 Mole-Mole Relationships

Objective: To learn to use a balanced equation to determine relationships between moles of reactants and moles of products

- We can use a balanced equation to predict the moles of products that a given number of moles of reactants with yield
  Example: $2H_2O$ (l) → $2H_2$ (g) + O_2 (g)
  This equation means that 2 mol of H_2O yields 2 mol of H_2 and 1 mol of O_2
  Suppose there are 4 mol of water, how many moles of product do we get?
  2[$2H_2O$ (l) → $2H_2$ (g) + O_2 (g)]
  4 H_2O → $4H_2$ (g) + $2O_2$ (g)
  Now you can say that: 4 mol of H_20 yields 4 mol of H_2 plus 2 mol of O_2

If you wanted to find 5.8 mol of water:

H_2O (l) → H_2 (g) + ½ O_2 (g)

5.8[H_2O (l) → H_2 (g) + ½ O_2 (g)]

5.8 H_2O (l) → 5.8 H_2 (g) + 2.9 O_2 (g)

- Determining Mole Rations

  $2H_2O$ (l) → $2H_2$ (g) + O_2 (g)

  5.8 mol H_2O yields ___ mol of O_2

  2 mol H_2O yields 1 mol of O_2

$$5.8\ \cancel{\text{mol } H_2O} \times \frac{1\ \text{mol } O_2}{2\ \cancel{\text{mol } H_2O}} = 2.9\ \text{mol } O_2$$

9.3 Mass Calculations

Objective: To learn to relate masses of reactants and products in a chemical reaction

- C_3H_8 (g) + O_2 (g) → CO_2 (g) + H_2O (g)
- What mass of oxygen is required to react exactly with 44.1 g of propane?
  1. Balance equation:
     C_3H_8 (g) + $5O_2$ (g) → $3CO_2$ (g) + $4H_2O$ (g)
  2. What we know: the balanced equation for the reaction, the mass of propane available (44.1 g)
     What we want to calculate: the mass of oxygen required to react exactly with all the propane
  3. Convert propane (C_3H_8) to moles
  4. Use the coefficients in the balanced equation to determine the moles of oxygen required
  5. Use the molar mass of oxygen to calculate grams of oxygen

9.4 Mass Calculations Using Scientific Notation

Objective: To carry out mass calculations that involve scientific notation

Steps For Calculating the Masses of Reactants and Products in Chemical Reactions:

1. Balance the equation for the reaction
2. Convert the masses of reactants or products to moles
3. Use the balanced equation to set up the appropriate mole ratio(s)
4. Use the mole ratio(s) to calculate the number of moles of the desired reactant or product
5. Convert from moles back to mass

Stoichiometry: the process of using a chemical equation to calculate the relative masses of reactants and products involved in a reaction

1. Unbalanced: $LiOH$ (s) + CO_2 (g) → Li_2CO_3 (s) + H_2O (l)

2. Balanced: $2LiOH$ (s) + CO_2 (g) → Li_2CO_3 (s) + H_2O (l)

3. Convert the given mass of LiOH to moles, using the molar mass of LiOH, which is 6.941 g + 16.00 g + 1.008 g = 23.95 g

$$1.00 \times 10^3 \cancel{\text{g LiOH}} \times \frac{1 \text{ mol LiOH}}{23.95 \cancel{\text{g LiOH}}} = 41.8 \text{ mol LiOH}$$

4. The appropriate mol ratio is

$$\frac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}}$$

5. Using the mole ratio we can calculate the moles of CO_2 needed to react with the given mass of LiOH

$$41.8 \cancel{\text{mol LiOH}} \times \frac{1 \text{ mol } CO_2}{2 \cancel{\text{mol LiOH}}} = 20.9 \text{ mol } CO_2$$

6. We calculate the mass of CO_2 by using its molar mass (44.01 g)

$$20.9 \cancel{\text{mol } CO_2} \times \frac{44.01 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}} = 920. \text{ g } CO_2 = 9.20 \times 10^2 \text{ g } CO_2$$

Thus 1.00×10^3 g of LiOH (s) can absorb 920. g CO_2 (g)

9.5 Mass Calculations: Comparing Two Reactions

Objective: To compare the stoichiometry of two reactions

- Baking soda, $NaHCO_3$, is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach
  $NaHCO_3$ (s) + HCl (aq) → NaCl (aq) + H_2O (l) + CO_2 (g)
- Mil of magnesia, which is an aqueous suspension of magnesium hydroxide, $Mg(OH)_2$, is also used as an antacid
  $Mg(OH)_2$ (s) + 2HCl (aq) → $2H_2O$ (l) + $MgCl_2$ (aq)
- Which antacid can consume the most stomach acid, 1.00 g of $NaHCO_3$ or 1.00 g of $Mg(OH)_2$?
  How many moles of HCl will react with 1.00 g of each antacid?
  The molar mass of $NaHCO_3$ is 84.01 g
  $$1.00 \cancel{\text{g } NaHCO_3} \times \frac{1 \text{ mol } NaHCO_3}{84.01 \cancel{\text{g } NaHCO_3}} = 0.0119 \ (1.19 \times 10^{-2}) \text{ mol } NaHCO_3$$
  $$1.19 \times 10^{-2} \cancel{\text{mol } NaHCO_3} \times \frac{1 \text{ mol HCl}}{\cancel{1 \text{ mol } NaHCO_3}} = 1.19 \times 10^{-2} \text{ mol HCL}$$
  Thus 1.00 g $NaHCO_3$ neutralizes 1.19×10^{-2} mol of HCl. Now we compare this to the number of moles of HCl that 1.00 g of $Mg(OH)_2$ neutralizes.
  The molar mass of $Mg(OH)_2$ is 58.33 g

$1.00\ \cancel{g\ Mg(OH)_2} \times \frac{1\ mol\ Mg(OH)_2}{58.33\ \cancel{g\ Mg(OH)_2}} = (1.71 \times 10^{-2})\ mol\ Mg(OH)_2$

$1.71 \times 10^{-2}\ \cancel{mol\ Mg(OH)_2} \times \frac{2\ mol\ HCl}{1\ \cancel{mol\ Mg(OH)_2}} = 3.42 \times 10^{-2}\ mol\ HCl$

Therefore, 1.00 g of $Mg(OH)_2$ is a more effective antacid than $NaHCO_3$ on a mass basis.

9.6 The Concept of Limiting Reactants

Objective: To understand the concept of limiting reactants

20 slices of bread, 24 slices of meat, 12 slices of cheese

How many sandwiches can be made? What will be left over?

Bread: 20 slices bread x $\frac{\text{1 sandwich}}{\text{2 slices bread}}$ = 10 sandwiches

Meat: 24 slices meat x $\frac{\text{1 sandwich}}{\text{3 slices meat}}$ = 8 sandwiches

Cheese: 12 slices cheese x $\frac{\text{1 sandwich}}{\text{1 slice cheese}}$ = 12 sandwiches

How many sandwiches can be made? The answer is 8. Once you run out of meat

you must stop making sandwiches. The meat is the limiting ingredient. 4 pieces

of bread are left over and 4 pieces of cheese. This can also be applied to

chemical reactions.

Limiting Reactant (limiting reagent) → the reactant that runs out first and thus limits the amounts of products that can be formed.

9.7 Calculations Involving a Limiting Reactant

Objective: To learn to recognize the limiting reactant in a reaction. To learn to use the limiting reactant to do stoichiometric calculations.

When chemicals are mixed together so that they can undergo a reaction, they are mixed in stoichiometric quantities—that is, in exactly the correct amounts so that all reactants "run out" at the same time.

- Example: $CH_4\ (g) + H_2O\ (g) \rightarrow 3H_2\ (g) + CO\ (g)$
- *What mass of water is needed to react exactly with 249 g of methane?*

  Molar mass CH_4 is 16.04 g/mol

  $249\ \cancel{g\ CH_4} \times \underline{1\ mol\ CH_4} = 15.5\ mol\ CH_4$

16.04 g CH_4

15.5 ~~mol CH_4~~ x $\frac{1 \text{ mol } H_2O}{1 \cancel{\text{mol } CH_4}}$ = 15.5 mol H_2O

15.5 ~~mol H_2O~~ x $\frac{18.02 \text{ g } H_2O}{1 \cancel{\text{mol } H_2O}}$ = 279 g H_2O

- If 249 g of methane is mixed with 279 g of water, both reactants will "run out" at the same time. If 249 g of methane is mixed with 300 g of water, the water will be in *excess*. The amount of methane in this case, limits the amount of products that can be formed.
    1. Write and balance the equation for the reaction
    2. Convert known masses of reactants to moles
    3. Using the numbers of moles of reactants and the appropriate mole ratios, determine which reactant is limiting
    4. Using the amount of limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product
    5. Convert from moles of product to grams of product, using the molar mass (if this is required by the problem)

9.8 Percent Yield

Objective: To learn to calculate actual yield as a percentage of theoretical yield

The limiting reagent controls the amount of product formed. The amount of product calculated in this way is called the **theoretical yield.** It is the amount of product predicted from the amounts of reactants used.

The **actual yield** of product, is the amount of product actually obtained.

The comparison of the two yields is called the **percent yield.**

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \text{percent yield}$$

Ex. $2NH_3$ (g) + $3CuO$ (s) → N_2 (g) + 3 Cu (s) + $3H_2O$ (g)

If this example gives 6.63 g of nitrogen instead of the predicted 10.6 g, the percent yield of nitrogen would be:

$$\frac{6.63 \cancel{\text{g } N_2}}{10.6 \cancel{\text{g } N_2}} \times 100\% = 62.5\%$$