Ch. 12.1 Review Notes
Stoichiometry

- ***Stoichiometry***: calculations of chemical quantities from balanced equations.

Reactants $\xrightarrow{\text{STOICHIOMETRY}}$ **Products**

Interpreting Chemical Equations

The first thing that must be done is to **BALANCE** the equation!

$$\underline{\ 1\ }N_{2\,(g)} + \underline{\ 3\ }H_{2\,(g)} \rightarrow \underline{\ 2\ }NH_{3\,(g)}$$

MOLE RATIOS!!

Information in an equation:

__1__ mole N_2 + __3__ moles H_2 → __2__ moles NH_3

__1__ molecule N_2 + __3__ molecules H_2 → __2__ molecules NH_3

__28.02__ grams N_2 + __6.06__ grams H_2 → __34.08__ grams NH_3

__34.08__ grams reactants → __34.08__ grams products

Interpreting Chemical Equations

Balance and interpret this equation in terms of moles, particles, and mass. Show *Law of Conservation of Mass* is proved!

_2_ $NaOH_{(aq)}$ + ___ $H_2SO_{4(aq)}$ → ___ $Na_2SO_{4\,(aq)}$ + _2_ $H_2O_{(l)}$

- 2 mol NaOH + 1 mol H_2SO_4 → 1 mol Na_2SO_4 + 2 mol H_2O
- 2 f.u. NaOH + 1 molec H_2SO_4 → 1 f.u. Na_2SO_4 + 2 molec. H_2O
- 80.00 g NaOH + 98.09 g H_2SO_4 → 142.05 g Na_2SO_4 + 36.04 g H_2O
- 178.09 grams reactants → 178.09 grams products

Ch. 12.2 Review Notes

Mole *to* Mole Conversions

- Conversion factor is a **mole ratio**.
- Mole ratios come from the coefficients on balanced equations.

Step 1: Identify the Known(A) and Unknown(B).

Step 2: Set up a mole ratio to change from moles A *to* moles B

Practice Problems: $N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$

1) How many moles of ammonia can be made from 7 moles of nitrogen reacting with an excess of hydrogen?

K(A): **7 mol N_2**

?(B): **mol NH_3**

$$7\ \cancel{\text{mol}\ N_2} \times \frac{2\ \text{mol}\ NH_3}{1\ \cancel{\text{mol}\ N_2}} = \boxed{14\ \text{moles of}\ NH_3}$$

Mole *to* Mole Conversions

Step 1: Identify the known(A) and unknown(B).

Step 2: Set up a mole ratio to change from moles A *to* moles B

Practice Problems: $\mathbf{N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}}$

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2) How many moles of hydrogen are required to completely react with 8 moles of nitrogen to produce ammonia?

K(A): **8 mol N_2**

?(B): **mol H_2**

$$8 \cancel{\text{ mol } N_2} \times \frac{3 \text{ mol } H_2}{1 \cancel{\text{ mol } N_2}} = \boxed{24 \text{ moles of } H_2}$$

3) How many moles of hydrogen are needed to react with an excess of nitrogen to make 10 moles of ammonia?

K(A): **10 mol NH_3**

?(B): **mol H_2**

$$10 \cancel{\text{ mol } NH_3} \times \frac{3 \text{ mol } H_2}{2 \cancel{\text{ mol } NH_3}} = \boxed{15 \text{ moles of } H_2}$$

Ch. 12.2 Review Notes pt. 2
Stoichiometry

Mass A to Mass B Conversion Problems

Step 1: Make sure you have a balanced equation!

Step 2: Identify the known (A) and unknown (B).

Step 3: Convert using the chart!

mass A $\rightarrow$ moles A $\rightarrow$ moles B $\rightarrow$ mass B

Practice Problem: How many grams of ammonia can be made from reacting 39.0 grams of nitrogen with an excess of hydrogen?

K(A): **39.0 g N_2**

$$N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$$

?(B): **g NH_3**

$$39.0\ \cancel{g\ N_2} \times \frac{1\ \cancel{mol\ N_2}}{28.02\ \cancel{g\ N_2}} \times \frac{2\ \cancel{mol\ NH_3}}{1\ \cancel{mol\ N_2}} \times \frac{17.04\ g\ NH_3}{1\ \cancel{mol\ NH_3}} = \boxed{47.4\ g\ NH_3}$$

Practice Problem: What mass of calcium hydroxide will be produced from the reaction of 2.50 mol of calcium hydride with excess water?

$$\mathbf{CaH_{2\,(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2\,(aq)} + 2H_{2\,(g)}}$$

Known-A	**Unknown-B**
2.50 mol CaH_2	**g $Ca(OH)_2$**

moles A → moles B → mass B

$$2.50\ \cancel{\text{mol CaH}_2} \times \frac{1\ \cancel{\text{mol Ca(OH)}_2}}{1\ \cancel{\text{mol CaH}_2}} \times \frac{74.10\ \text{g Ca(OH)}_2}{1\ \cancel{\text{mol Ca(OH)}_2}} = \boxed{185\ \text{g Ca(OH)}_2}$$

Practice Problem: What mass of water is needed to react 2.50 mol of calcium hydride?

$$\mathbf{CaH_{2\,(s)} + 2H_2O_{(l)} \rightarrow Ca(OH)_{2\,(aq)} + 2H_{2\,(g)}}$$

Known-A	**Unknown-B**
2.50 mol CaH_2	**g H_2O**

moles A → moles B → mass B

$$2.50\ \cancel{\text{mol CaH}_2} \times \frac{2\ \cancel{\text{mol H}_2\text{O}}}{1\ \cancel{\text{mol CaH}_2}} \times \frac{18.02\ \text{g H}_2\text{O}}{1\ \cancel{\text{mol H}_2\text{O}}} = \boxed{90.1\ \text{g H}_2\text{O}}$$

- **Carbon dioxide reacts with a calcium oxide solution to produce calcium carbonate. What mass of calcium carbonate would be produced if you blew 0.0900 mol CO_2 into a test tube of CaO?**

$$CO_{2\,(g)} + CaO_{(aq)} \rightarrow CaCO_{3\,(s)}$$

Known-A	**Unknown-B**
0.0900 mol CO_2	**g $CaCO_3$**

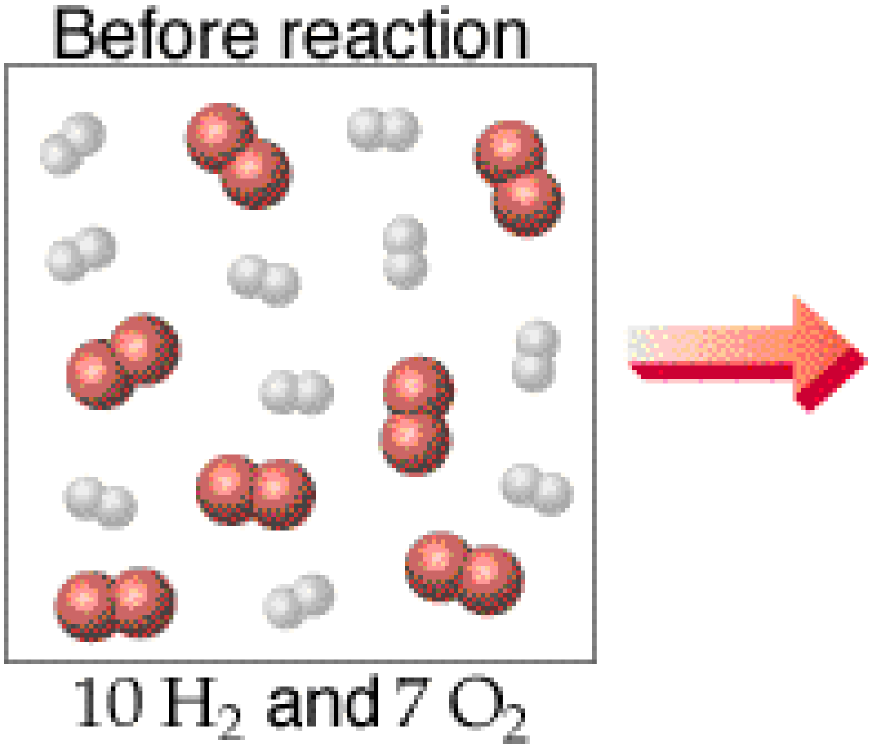
Moles A → Moles B → Mass B

$$0.0900\ \cancel{\text{mol } CO_2} \times \frac{1\ \cancel{\text{mol } CaCO_3}}{1\ \cancel{\text{mol } CO_2}} \times \frac{100.08\text{ g } CaCO_3}{1\ \cancel{\text{mol } CaCO_3}} = \boxed{9.01\text{ g } CaCO_3}$$

Ch. 12.3 Review Notes
Stoichiometry

Limiting Reagent (or Limiting Reactant)

- The ***limiting reagent*** is the reactant that ____runs____ ___out___ first.
- The reactant that is in abundance is called the ____excess____ reagent.



$$2H_2 + O_2 \rightarrow 2\,H_2O$$

Limiting Reagent Calculations

Step 1: Do a mole-mole conversion for one of the substances.

- ***The answer is how much of it you need.***

Step 2: Compare your answer to how much reactant was given.

- ***Do you have enough? If not, this reactant is your limiting reagent!***

Practice Problems: $N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$

1) If 2.7 moles of nitrogen reacts with 6.3 moles of hydrogen, which will you run out of first?

$$2.7 \cancel{\text{mol } N_2} \times \frac{3 \text{ mol } H_2}{1 \cancel{\text{mol } N_2}} =$$ 8.1 moles of H_2 are needed for the reaction.

Do we have enough H_2?
No ! (6.3<8.1)

Limiting Reagent = H_2

Limiting Reagent Calculations

Step 1: Do a mole-mole conversion for one of the substances.

- ***The answer is how much of it you need.***

Step 2: Compare your answer to how much reactant was given.

- ***Do you have enough? If not, this reactant is your limiting reagent!***

Practice Problems: $N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$

2) If 3.9 moles of nitrogen reacts with 12.1 moles of hydrogen, what is the limiting reagent?

$$3.9 \cancel{\text{ mol } N_2} \times \frac{3 \text{ mol } H_2}{1 \cancel{\text{ mol } N_2}} =$$

11.7 moles of H_2 are needed for the reaction.

Do we have enough H_2?
Yes ! (12.1>11.7)

Limiting Reagent = N_2

Calculating Excess Reagent

Step 1: Do a mole-mole conversion *starting with the limiting reagent*

The answer is how much of the excess reagent you need to completely react with the limiting reagent.

Step 2: Subtract this answer from the original amount given

Practice Problem: Find the number of moles of excess reagent from the first practice problem.

$$6.3\ \cancel{\text{mol}}\ H_2 \times \frac{1\ \text{mol}\ N_2}{3\ \cancel{\text{mol}\ H_2}} =$$

2.1 moles of N_2 are needed for the reaction.

2.7 moles of N_2 were originally given, so the excess will be…

2.7 moles given – 2.1 moles needed =

Excess Reagent (or Excess Reactant)

Step 1: Do a mole-mole conversion ***starting with the limiting reagent***

Step 2: Subtract this answer from the original amount given

Practice Problem: Find the number of moles of excess reagent from the second practice problems.

- We already know mol H_2… so just subtract.

12.1 mol H_2 given −11.7 mol H_2 needed= (

Ch. 12.4 Review Notes
Percent Yield

Percent Yield

- ***Percent Yield*** is a ratio that tells us how ____efficient____ a chemical reaction is.
- The higher the % yield, the more efficient the reaction is.

$$\textbf{\% Yield} = \frac{\textbf{Actual yield}}{\textbf{Theoretical yield}} \textbf{ x 100}$$

- The __**actual**__ **yield** is the **experimental data** from running the reaction in a lab.
- The __**theoretical**__ **yield** is the **ideal** amount! Calculate this amount using **stoichiometry**!

Percent Yield

Practice Problem: $2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(g)}$

1) A student **reacts** 40.0 grams of hydrogen with an excess of oxygen and **produces** 300.0 grams of water. Find the % yield for this reaction.

Step 1: Calculate mass A → mass B *starting with* ***reactant!***

The answer you get is how much water you "theoretically" should have produced.

$$40.0\ \cancel{g\ H_2} \times \frac{1\ \cancel{mol\ H_2}}{2.02\ \cancel{g\ H_2}} \times \frac{2\ \cancel{mol\ H_2O}}{2\ \cancel{mol\ H_2}} \times \frac{18.02\ g\ H_2O}{1\ \cancel{mol\ H_2O}} = 357\ g\ H_2O$$

Step 2: Identify *actual yield* then calculate % yield!

$$\%\ \text{Yield} = \frac{300\ g}{357\ g} \times 100 = \boxed{84.0\%}$$

Percent Yield

Practice Problem: $2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(g)}$

2) If 2.0 grams of hydrogen completely reacted with 16.0 grams of oxygen but only produced 17.5 grams of water, what is the % yield for the reaction?

Theoretical Yield = 2.0 g + 16.0 g = 18.0 grams

$$\% \text{ Yield} = \frac{17.5 \text{ g}}{18.0 \text{ g}} \times 100 = 97.2\%$$