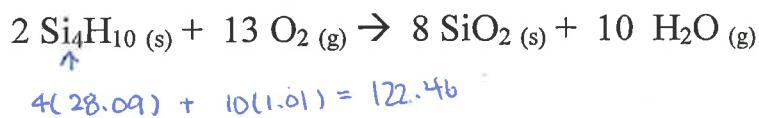


Name KEY
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Chemistry 11

Chapter 4 Review – Stoichiometry

1. Given the balanced equation:



- a) What volume of oxygen (STP) is required to react with 204.0 g of Si_4H_{10} ?

$$204.0 \text{ g } \text{Si}_4\text{H}_{10} \times \frac{1 \text{ mol } \text{Si}_4\text{H}_{10}}{122.46 \text{ g } \text{Si}_4\text{H}_{10}} \times \frac{13 \text{ mol O}_2}{2 \text{ mol } \text{Si}_4\text{H}_{10}} \times \frac{22.4 \text{ L O}_2}{1 \text{ mol O}_2} = \boxed{242.5 \text{ L O}_2}$$

- b) What mass of SiO_2 is formed when 345.0 g of H_2O are formed?

$$345.0 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{8 \text{ mol SiO}_2}{10 \text{ mol H}_2\text{O}} \times \frac{60.09 \text{ g SiO}_2}{1 \text{ mol SiO}_2} = \boxed{920.4 \text{ g SiO}_2}$$

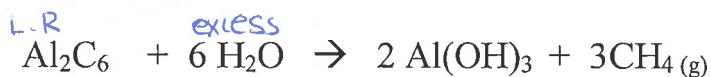
- c) How many molecules of H_2O are formed when 17.92 L of O_2 are used at STP?

$$17.92 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} \times \frac{10 \text{ mol H}_2\text{O}}{13 \text{ mol O}_2} \times \frac{6.02 \times 10^{23} \text{ molecules H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{3.71 \times 10^{23} \text{ molecules H}_2\text{O}}$$

- d) How many moles of Si_4H_{10} are needed to just react with 1.204×10^{26} molecules of oxygen?

$$1.204 \times 10^{26} \text{ molec. O}_2 \times \frac{1 \text{ mol O}_2}{6.02 \times 10^{23} \text{ molec. O}_2} \times \frac{2 \text{ mol Si}_4\text{H}_{10}}{13 \text{ mol O}_2} = \boxed{30.8 \text{ mol Si}_4\text{H}_{10}}$$

2. Given the balanced equation:



- a) If 34.5 grams of Al_2C_6 is mixed with 72.0 grams of water, which reactant is in excess? Show by calculations.

$$34.5 \text{ g Al}_2\text{C}_6 \times \frac{1 \text{ mol Al}_2\text{C}_6}{126.02 \text{ g Al}_2\text{C}_6} \times \frac{2 \text{ mol Al(OH)}_3}{1 \text{ mol Al}_2\text{C}_6} \times \frac{78.01 \text{ g Al(OH)}_3}{1 \text{ mol Al(OH)}_3} = 42.71 \text{ g Al(OH)}_3$$

$$72.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol Al(OH)}_3}{6 \text{ mol H}_2\text{O}} \times \frac{78.01 \text{ g Al(OH)}_3}{1 \text{ mol Al(OH)}_3} = 103.9 \text{ g Al(OH)}_3 \quad [\text{H}_2\text{O is excess}]$$

- b) If 34.5 grams of Al_2C_6 is mixed with 72.0 grams of water, what mass of Al(OH)_3 is formed?

$$42.71 \text{ g Al(OH)}_3$$

- c) If 34.5 grams of Al_2C_6 is mixed with 72.0 grams of water, what volume of CH_4 is formed at STP?

$$34.5 \text{ g Al}_2\text{C}_6 \times \frac{1 \text{ mol Al}_2\text{C}_6}{126.02 \text{ g Al}_2\text{C}_6} \times \frac{3 \text{ mol CH}_4}{1 \text{ mol Al}_2\text{C}_6} \times \frac{22.4 \text{ L CH}_4}{1 \text{ mol CH}_4} = 18.4 \text{ L CH}_4$$

3. Given the equation: $4 \text{NH}_3 + 5 \text{O}_2 \rightarrow 4 \text{NO} + 6 \text{H}_2\text{O}$

When 51.0 grams of NH_3 is burned in an excess of oxygen, 52.65 g of water are produced.

- a) Calculate the theoretical yield of H_2O .

$$51.0 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.04 \text{ g NH}_3} \times \frac{16 \text{ mol H}_2\text{O}}{4 \text{ mol NH}_3} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 80.99 \text{ g H}_2\text{O}$$

- b) Calculate the % yield of H_2O .

$$\% \text{ Yield} = \frac{\text{AY}}{\text{TY}} \times 100 = \frac{52.65}{80.99} \times 100\% = 65.1\% \text{ Yield}$$

4. Given the equation: $\text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3$ When 4.0 grams of hydrogen is combined with an excess of nitrogen, a 92% yield of NH_3 is obtained.

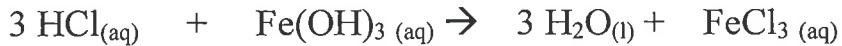
- a) Calculate the theoretical yield of NH_3

$$4.0 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.02 \text{ g H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{17.04 \text{ g NH}_3}{1 \text{ mol NH}_3} = 22.49 \text{ g NH}_3$$

- b) Calculate the actual yield of NH_3

$$\text{AY} = \frac{\% \text{ Y}}{100} \times \text{TY} = \frac{92}{100} \times 22.49 = 20.4 \text{ g NH}_3$$

5. Given the balanced equation:



- a) It takes 19.56 mL of 0.50 M HCl to titrate a 25.0 mL sample of a solution of Fe(OH)_3 . Calculate the $[\text{Fe(OH)}_3]$

$$\text{moles HCl} = 0.50 \text{ M} \times 0.01956 \text{ L} = 0.00978 \text{ mol HCl}$$

$$\text{moles Fe(OH)}_3 = 0.00978 \text{ mol HCl} \times \frac{1 \text{ mol Fe(OH)}_3}{3 \text{ mol HCl}} = 0.00326 \text{ mol Fe(OH)}_3$$

$$[\text{Fe(OH)}_3] = \frac{0.00326 \text{ mol Fe(OH)}_3}{0.0250 \text{ L Fe(OH)}_3} = 0.13 \text{ M}$$

- b) What mass of Fe(OH)_3 is needed to completely react with 10.0 mL of 0.50M HCl solution?

$$\text{moles HCl} = 0.50 \text{ M} \times 0.0100 \text{ L} = 0.00500 \text{ mol HCl}$$

$$0.00500 \text{ mol HCl} \times \frac{1 \text{ mol Fe(OH)}_3}{3 \text{ mol HCl}} \times \frac{106.88 \text{ g Fe(OH)}_3}{1 \text{ mol Fe(OH)}_3}$$

$$= 0.18 \text{ g Fe(OH)}_3$$

- c) What volume of 0.50M HCl is required to titrate a 21.36 gram sample of iron (III) hydroxide?

$$21.36 \text{ g Fe(OH)}_3 \times \frac{1 \text{ mol Fe(OH)}_3}{106.88 \text{ g Fe(OH)}_3} \times \frac{3 \text{ mol HCl}}{1 \text{ mol Fe(OH)}_3} = 0.600 \text{ mol HCl}$$



$$L = \frac{\text{mol}}{\text{M}} = \frac{0.600 \text{ mol}}{0.50 \text{ M}} = 1.2 \text{ L}$$

