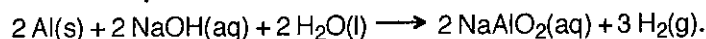


$$\text{volume of KOH} = \frac{n}{c} = \frac{0.0625 \text{ mol}}{0.200 \text{ mol/L}} = 0.313 \text{ L}$$

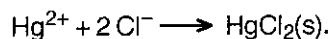
**EXERCISES:**

17. A student wants to put 50.0 L of hydrogen gas at STP into a plastic bag by reacting excess aluminum metal with 3.00 M sodium hydroxide solution according to the reaction



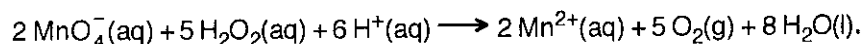
What volume of NaOH solution is required?

18. What volume of 0.250 M HCl is required to completely neutralize 25.0 mL of 0.318 M NaOH? [Hint: what is the balanced equation for the reaction between HCl and NaOH?]
19. A technician analyzes a sample of water from the "tailings" pond of a mine for the presence of mercury. After treating and concentrating the water sample, the technician carries out the titration reaction



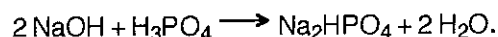
A 25.0 mL sample of water containing mercury reacts with 15.4 mL of 0.0148 M  $\text{Cl}^-$  (as NaCl).

- a) What is the molar concentration of the mercury in the water sample?  
b) What mass of  $\text{HgCl}_2$  is formed in the reaction?
20. A 10.0 mL sample of a saturated solution of  $\text{Ca(OH)}_2$  is titrated with 23.5 mL of 0.0156 M HCl.  
a) What is the molarity of the  $\text{Ca(OH)}_2$  in the saturated solution?  
b) What mass of  $\text{Ca(OH)}_2$  is dissolved in 250.0 mL of saturated  $\text{Ca(OH)}_2$ ?
21. A student titrates a 2.00 mL sample of hydrogen peroxide solution,  $\text{H}_2\text{O}_2\text{(aq)}$ , according to the reaction

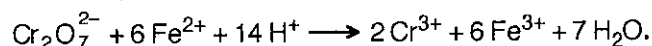


The supply bottle of  $\text{H}_2\text{O}_2$  is labelled as "3.00% by volume" (3.00 mL of  $\text{H}_2\text{O}_2$  per 100 mL of solution), which the student calculates to have  $[\text{H}_2\text{O}_2] = 1.24 \text{ M}$ .

- a) What volume of 0.0496 M  $\text{MnO}_4^-$  is required for the titration?  
b) What volume of  $\text{O}_2\text{(g)}$  at STP is produced during the reaction?
22. A 1.00 mL sample of pure phosphoric acid,  $\text{H}_3\text{PO}_4$ , is titrated with 43.8 mL of 0.853 M NaOH according to the reaction



- a) What is the molar concentration of pure  $\text{H}_3\text{PO}_4$ ?  
b) Calculate the density of pure  $\text{H}_3\text{PO}_4$ .
23. The iron present in a sample of iron ore is converted to  $\text{Fe}^{2+}$  and titrated with dichromate ion



If 17.6 mL of 0.125 M dichromate ion is required to titrate a 25.0 mL sample of  $\text{Fe}^{2+}$  solution,

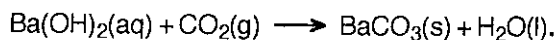
- a) what is the molarity of the  $\text{Fe}^{2+}$ ?      b) what mass of iron is present in the 25.0 mL sample?
24. Prior to analyzing a fertilizer sample containing  $\text{NH}_4\text{NO}_3$ , a chemist makes a test solution by dissolving 15.5 g of pure  $\text{NH}_4\text{NO}_3$  and diluting it to 500.0 mL. If the chemist wishes to carry out the titration reaction



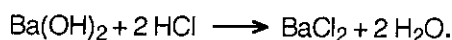
such that the reaction requires 25.0 mL of NaOH when 10.0 mL of the  $\text{NH}_4\text{NO}_3$  solution is titrated,

- a) what is the molarity of the NaOH she should use?  
b) what volume of  $\text{NH}_3\text{(g)}$  at STP would be produced?

25. CARE! The  $\text{CO}_2$  content of a 10.0 L sample of air at STP is determined as follows. The air is pumped through a flask containing 25.0 mL of 0.0538 M  $\text{Ba}(\text{OH})_2$ , precipitating the  $\text{CO}_2$  present as  $\text{BaCO}_3$ :



- a) How many moles of  $\text{Ba}(\text{OH})_2$  are present in the original  $\text{Ba}(\text{OH})_2$  solution?  
 b) Only a small amount of the  $\text{Ba}(\text{OH})_2$  present reacts with the added  $\text{CO}_2$ . The remaining, unreacted,  $\text{Ba}(\text{OH})_2$  is titrated with hydrochloric acid according to the equation



If the titration requires 23.0 mL of 0.104 M HCl, how many moles of  $\text{Ba}(\text{OH})_2$  remain in the  $\text{Ba}(\text{OH})_2$  solution after reaction with the  $\text{CO}_2$  in the air?

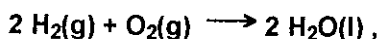
- c) How many moles of  $\text{Ba}(\text{OH})_2$  are reacted by the  $\text{CO}_2$ ?  
 d) How many moles of  $\text{CO}_2$  are in the sample of air?  
 e) How many litres of  $\text{CO}_2$  at STP are contained in the 10.0 L sample of air? What percentage of the air sample's volume is  $\text{CO}_2$ ?

#### VII.4. STOICHIOMETRY OF EXCESS QUANTITIES

The stoichiometry calculations in the previous section assume that a given reactant is completely used up during a reaction. Nevertheless, reactions frequently are carried out in such a way that one or more of the reactants actually are present in EXCESS amounts. Some reasons for having an excess amount include:

- (i) deliberately adding an excess of one reactant to make sure all of a second reactant is used (the second reactant may be too expensive to waste or harmful to the environment if left unreacted).  
 (ii) unavoidably having a reactant in excess because a limited amount of another reactant is available.

**EXAMPLE:** If 20.0 g of  $\text{H}_2(\text{g})$  react with 100.0 g of  $\text{O}_2(\text{g})$  according to the reaction



which reactant is present in excess and by how many grams?

A simple way to find the excess reactant is to calculate the mass of some **arbitrarily-selected product**. In this case, you can find how much  $\text{H}_2\text{O}$  can be formed.

$$\text{mass of H}_2\text{O (based on H}_2\text{)} = 20.0 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.0 \text{ g H}_2} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 180.0 \text{ g}$$

$$\text{mass of H}_2\text{O (based on O}_2\text{)} = 100.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 112.5 \text{ g}$$

There is enough  $\text{H}_2$  to make 180.0 g of  $\text{H}_2\text{O}$ , but only enough  $\text{O}_2$  to make 112.5 g, so that the  $\text{O}_2$  **sets a limit** on the amount of  $\text{H}_2\text{O}$  formed. Therefore,  $\text{O}_2$  is called the **LIMITING REACTANT**. Since there is only enough  $\text{O}_2$  to make 112.5 g of  $\text{H}_2\text{O}$ , not all of the  $\text{H}_2$  can be used and  $\text{H}_2$  is called the **EXCESS REACTANT**.

To find the mass of  $\text{H}_2$  in excess, find the mass of  $\text{H}_2$  which actually reacts based on either the mass of limiting reactant ( $\text{O}_2$ ), or the mass of a product ( $\text{H}_2\text{O}$ ) formed by the limiting reactant. Then, subtract the mass of  $\text{H}_2$  which reacts from the starting mass of  $\text{H}_2$ . We **arbitrarily** find the mass of  $\text{H}_2$  used by starting with the mass of the limiting reactant,  $\text{O}_2$ .

$$\text{mass of H}_2 \text{ reacted} = 100.0 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \times \frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \times \frac{2.0 \text{ g H}_2}{1 \text{ mol H}_2} = 12.5 \text{ g}$$

$$\text{And: mass of H}_2 \text{ (in excess)} = \text{mass H}_2 \text{ (at start)} - \text{mass H}_2 \text{ (reacted)} = 20.0 - 12.5 = 7.5 \text{ g}$$