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

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

$$2\,Al(s) + 2\,NaOH(aq) + 2\,H_2O(l) \longrightarrow 2\,NaAlO_2(aq) + 3\,H_2(g).$$

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

$$Hg^{2+} + 2\,Cl^- \longrightarrow HgCl_2(s).$$

A 25.0 mL sample of water containing mercury reacts with 15.4 mL of 0.0148 M Cl^-(as NaCl).
a) What is the molar concentration of the mercury in the water sample?
b) What mass of $HgCl_2$ is formed in the reaction?

20. A 10.0 mL sample of a saturated solution of $Ca(OH)_2$ is titrated with 23.5 mL of 0.0156 M HCl.
a) What is the molarity of the $Ca(OH)_2$ in the saturated solution?
b) What mass of $Ca(OH)_2$ is dissolved in 250.0 mL of saturated $Ca(OH)_2$?

21. A student titrates a 2.00 mL sample of hydrogen peroxide solution, $H_2O_2(aq)$, according to the reaction

$$2\,MnO_4^-(aq) + 5\,H_2O_2(aq) + 6\,H^+(aq) \longrightarrow 2\,Mn^{2+}(aq) + 5\,O_2(g) + 8\,H_2O(l).$$

The supply bottle of H_2O_2 is labelled as "3.00% by volume" (3.00 mL of H_2O_2 per 100 mL of solution), which the student calculates to have $[H_2O_2] = 1.24$ M.

a) What volume of 0.0496 M MnO_4^- is required for the titration?

b) What volume of $O_2(g)$ at STP is produced during the reaction?

22. A 1.00 mL sample of pure phosphoric acid, H_3PO_4 , is titrated with 43.8 mL of 0.853 M NaOH according to the reaction

$$2\,NaOH + H_3PO_4 \longrightarrow Na_2HPO_4 + 2\,H_2O.$$

a) What is the molar concentration of pure H_3PO_4?
b) Calculate the density of pure H_3PO_4.

23. The iron present in a sample of iron ore is converted to Fe^{2+} and titrated with dichromate ion

$$Cr_2O_7^{2-} + 6\,Fe^{2+} + 14\,H^+ \longrightarrow 2\,Cr^{3+} + 6\,Fe^{3+} + 7\,H_2O.$$

If 17.6 mL of 0.125 M dichromate ion is required to titrate a 25.0 mL sample of Fe^{2+} solution,
a) what is the molarity of the Fe^{2+}? b) what mass of iron is present in the 25.0 mL sample?

24. Prior to analyzing a fertilizer sample containing NH_4NO_3 , a chemist makes a test solution by dissolving 15.5 g of pure NH_4NO_3 and diluting it to 500.0 mL. If the chemist wishes to carry out the titration reaction

$$NH_4NO_3(aq) + NaOH(aq) \longrightarrow NH_3(g) + H_2O(l) + NaNO_3(aq)$$

such that the reaction requires 25.0 mL of NaOH when 10.0 mL of the NH_4NO_3 solution is titrated,
a) what is the molarity of the NaOH she should use?
b) what volume of $NH_3(g)$ at STP would be produced?

25. CARE! The 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 $Ba(OH)_2$, precipitating the CO_2 present as $BaCO_3$:

$$Ba(OH)_2(aq) + CO_2(g) \longrightarrow BaCO_3(s) + H_2O(l).$$

a) How many moles of $Ba(OH)_2$ are present in the original $Ba(OH)_2$ solution?

b) Only a small amount of the $Ba(OH)_2$ present reacts with the added CO_2. The remaining, unreacted, $Ba(OH)_2$ is titrated with hydrochloric acid according to the equation

$$Ba(OH)_2 + 2\,HCl \longrightarrow BaCl_2 + 2\,H_2O.$$

If the titration requires 23.0 mL of 0.104 M HCl, how many moles of $Ba(OH)_2$ remain in the $Ba(OH)_2$ solution after reaction with the CO_2 in the air?

c) How many moles of $Ba(OH)_2$ are reacted by the CO_2?

d) How many moles of CO_2 are in the sample of air?

e) How many litres of CO_2 at STP are contained in the 10.0 L sample of air? What percentage of the air sample's volume is 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 $H_2(g)$ react with 100.0 g of $O_2(g)$ according to the reaction**

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l),$$

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 H_2O can be formed.

$$\text{mass of } H_2O \text{ (based on } H_2) = 20.0\text{ g } H_2 \times \frac{1\text{ mol } H_2}{2.0\text{ g } H_2} \times \frac{2\text{ mol } H_2O}{2\text{ mol } H_2} \times \frac{18.0\text{ g } H_2O}{1\text{ mol } H_2O} = \mathbf{180.0\ g}$$

$$\text{mass of } H_2O \text{ (based on } O_2) = 100.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{2\text{ mol } H_2O}{1\text{ mol } O_2} \times \frac{18.0\text{ g } H_2O}{1\text{ mol } H_2O} = \mathbf{112.5\ g}$$

There is enough H_2 to make 180.0 g of H_2O, but only enough O_2 to make 112.5 g, so that the O_2 **sets a limit** on the amount of H_2O formed. Therefore, O_2 is called the **LIMITING REACTANT.** Since there is only enough O_2 to make 112.5 g of H_2O, not all of the H_2 can be used and H_2 is called the **EXCESS REACTANT.**

To find the mass of H_2 in excess, find the mass of H_2 which actually reacts based on either the mass of limiting reactant (O_2), or the mass of a product (H_2O) formed by the limiting reactant. Then, subtract the mass of H_2 which reacts from the starting mass of H_2. We **arbitrarily** find the mass of H_2 used by starting with the mass of the limiting reactant, O_2.

$$\text{mass of } H_2 \text{ reacted} = 100.0\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.0\text{ g } O_2} \times \frac{2\text{ mol } H_2}{1\text{ mol } O_2} \times \frac{2.0\text{ g } H_2}{1\text{ mol } H_2} = \mathbf{12.5\ g}$$

And: mass of H_2 (in excess) = mass H_2 (at start) – mass H_2 (reacted) = 20.0 – 12.5 = **7.5 g**