Molar mass, M, of the unknown gas:

$$\begin{aligned} M &= \frac{m}{n} \\ &= \frac{10.00 \text{ g}}{0.088037 \text{ mol}} \\ &= 113.16057 \text{ g/mol} \end{aligned}$$

Molar mass, M, of the empirical formula C_4H_9:

$$\begin{aligned} M_{C_4H_9} &= 4M_C + 9M_H \\ &= 4(12.01 \text{ g/mol}) + 9(1.01 \text{ g/mol}) \\ &= 57.13 \text{ g/mol} \end{aligned}$$

Ratio of molar masses: $\dfrac{113.16057 \text{ g/mol}}{57.13 \text{ g/mol}} = \dfrac{1.9807}{1} = \dfrac{2}{1}$

Since the ratio of the molar masses is 2:1, the molecular formula of the unknown liquid is $C_4H_9 \times 2 = C_8H_{18}$.

Check Your Solution

The simple integer ratio of the two molar masses makes the molecular formula a reasonable answer. The variables are substituted correctly into the ideal gas law equation and the units are correct.

Section 12.2 The Ideal Gas Law

Solutions for Practice Problems

Student Edition page 560

31. Practice Problem (page 560)

What volume of hydrogen gas will be produced at 93.0 kPa and 23°C from the reaction of 33 mg of magnesium with hydrochloric acid?

What Is Required?

You need to find the volume of gas that is produced when magnesium reacts with hydrochloric acid under specific temperature and pressure conditions.

What Is Given?

You know that the reactants are hydrochloric acid, HCl(aq), and magnesium, Mg(s).

You known that the products are magnesium chloride, $MgCl_2$(aq), and hydrogen gas, H_2(g).

You know the mass of the magnesium: $m = 33$ mg
You know the temperature and pressure:
$T = 23°C$
$P = 93.0$ kPa

Plan Your Strategy
Write the balanced equation for the chemical reaction.
Convert the mass of magnesium to grams.
Determine the molar mass of magnesium using the atomic mass from the periodic table.

Calculate the amount in moles of magnesium using the relationship $n = \frac{m}{M}$.

Use the amount in moles and the mole ratio from the balanced chemical equation to calculate the amount in moles of H_2(g) produced.
Convert the temperature from the Celsius scale to the Kelvin scale.

Use the universal gas constant: $R = 8.314 \frac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$

Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable V.
Substitute the given data into the equation and solve for V_{H_2}.

Act on Your Strategy
Balanced chemical equation: $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
Mole ratio: 1 mole 2 moles 1 mole 1 mole

Mass conversion:
Mass, m, of Mg(s):

$$\begin{aligned} m_{Mg} &= 33 \cancel{\text{mg}} \times 1 \times 10^{-3} \text{ g}/\cancel{\text{mg}} \\ &= 0.033 \text{ g} \end{aligned}$$

Molar mass, M, of magnesium: 24.31 g/mol (from the periodic table)

Amount in moles, n, of Mg(s):

$$\begin{aligned} n_{Mg} &= \frac{m}{M} \\ &= \frac{0.033 \cancel{\text{g}}}{24.31 \cancel{\text{g}}/\text{mol}} \\ &= 0.0013574 \text{ mol} \end{aligned}$$

Amount in moles, n, of $H_2(g)$:

$$\frac{1 \text{ mol } H_2}{1 \text{ mol Mg}} = \frac{n_{H_2}}{0.0013574 \text{ mol Mg}}$$

$$n_{H_2} = \frac{1 \text{ mol } H_2 \times 0.0013574 \cancel{\text{mol Mg}}}{1 \cancel{\text{mol Mg}}}$$

$$= 0.0013574 \text{ mol}$$

Temperature conversion:

$$T = 23°\text{C} + 273.15$$

$$= 296.15 \text{ K}$$

Isolation of variable V:

$$PV = nRT$$

$$\frac{\cancel{P}V}{\cancel{P}} = \frac{nRT}{P}$$

$$V = \frac{nRT}{P}$$

Substitution to solve for V_{H_2}:

$$V_{H_2} = \frac{nRT}{P}$$

$$= \frac{0.0013574 \cancel{\text{mol}} \times 8.314 \frac{\cancel{\text{kPa}} \bullet \text{L}}{\cancel{\text{mol}} \bullet \cancel{\text{K}}} \times 296.15 \cancel{\text{K}}}{93.0 \cancel{\text{kPa}}}$$

$$= 0.035937 \text{ L}$$

$$= 0.036 \text{ L}$$

The volume of hydrogen that will be produced is 0.036 L.

Check Your Solution

The small volume of hydrogen gas produced seems reasonable given the mole ratio in the balanced chemical equation and the small mass of magnesium that reacted. The answer correctly shows two significant digits.

Check Your Solution
The balanced chemical equation indicates that there should be twice as much HCl(g) produced as either reactant. There is almost an equal amount of each reactant and the volume of HCl(g) is expected to be about 16 L. The answer correctly shows two significant digits.

33. Practice Problem (page 560)
Determine the volume of nitrogen gas produced when 120 g of sodium azide, NaN_3(s), decomposes at 27°C and 100.5 kPa. Sodium metal is the other product.

What Is Required?
You need to find the volume of nitrogen that is produced when sodium azide decomposes.

What Is Given?
You know that the reactant is sodium azide, NaN_3(s).
You know that the products are sodium, Na(s), and nitrogen gas, N_2(g).
You know the mass of the NaN_3(s): m = 120 g
You know the temperature and pressure:
T = 27°C
P = 100.5 kPa

Plan Your Strategy
Write the balanced equation for the chemical reaction.
Determine the molar mass of NaN_3(s).
Calculate the amount in moles of NaN_3(s) using the relationship $n = \frac{m}{M}$.
Use the amount in moles of NaN_3(s) and the mole ratio from the balanced chemical equation to calculate the amount in moles of N_2(g) produced.
Convert the temperature from the Celsius scale to the Kelvin scale.
Use the universal gas constant: $R = 8.314 \frac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$
Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable V.
Substitute the given data into the equation and solve for V_{N_2}.

Act on Your Strategy
Balanced chemical equation: $2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$
Mole ratio: 2 moles 2 moles 3 moles

Molar mass, M, of NaN_3(s):

$$\begin{aligned} M_{NaN_3} &= 1M_{Na} + 3M_{N} \\ &= 1(22.99 \text{ g/mol}) + 3(14.01 \text{ g/mol}) \\ &= 65.02 \text{ g/mol} \end{aligned}$$

Amount in moles, n, of NaN_3(s):

$$\begin{aligned} n_{NaN_3} &= \frac{m}{M} \\ &= \frac{120 \cancel{\text{g}}}{65.02 \cancel{\text{g}}/\text{mol}} \\ &= 1.845585 \text{ mol} \end{aligned}$$

Amount in moles, n, of N_2(g):

$$\frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} = \frac{n_{N_2}}{1.845585 \text{ mol NaN}_3}$$

$$\begin{aligned} n_{N_2} &= \frac{3 \text{ mol N}_2 \times 1.845585 \cancel{\text{mol NaN}_3}}{2 \cancel{\text{mol NaN}_3}} \\ &= 2.76837 \text{ mol} \end{aligned}$$

Temperature conversion:

$$\begin{aligned} T &= 27°\text{C} + 273.15 \\ &= 300.15 \text{ K} \end{aligned}$$

Isolation of the variable V:

$$PV = nRT$$

$$\frac{\cancel{P}V}{\cancel{P}} = \frac{nRT}{P}$$

$$V = \frac{nRT}{P}$$

Substitution to solve for V_{N_2}:

$$V_{N_2} = \frac{nRT}{P}$$

$$= \frac{2.76837 \cancel{\text{mol}} \times 8.314 \frac{\cancel{\text{kPa}} \bullet \text{L}}{\cancel{\text{mol}} \bullet \cancel{\text{K}}} \times 300.15 \cancel{\text{K}}}{100.5 \cancel{\text{kPa}}}$$

$$= 68.739511 \text{ L}$$

$$= 69 \text{ L}$$

The volume of nitrogen gas produced is 69 L.

Check Your Solution
The volume of nitrogen gas produced seems reasonable given the mole ratio in the balanced chemical equation and the mass of sodium azide that reacted. The answer correctly shows two significant digits.

34. Practice Problem (page 560)
When calcium carbonate, $CaCO_3$(s), is heated, it decomposes to form calcium oxide, CaO(s), and carbon dioxide gas. How many litres of carbon dioxide will be produced at STP if 2.38 kg of calcium carbonate reacts completely?

What Is Required?
You need to find the volume of carbon dioxide gas that is produced when calcium carbonate decomposes.

What Is Given?
You know that the reactant is calcium carbonate, $CaCO_3$(s).
You know that the products are calcium oxide, CaO(s), and carbon dioxide gas, CO_2(g).
You know the mass of calcium carbonate that reacts: m = 2.38 kg
You know the temperature and pressure at STP:
T = 273.15 K
P = 101.325 kPa

Plan Your Strategy
Write the balanced equation for the chemical reaction.
Convert the mass of $CaCO_3$(s) to grams.
Determine the molar mass of $CaCO_3$(s) using the atomic masses from the periodic table.

Calculate the amount in moles of $CaCO_3$(s) using the relationship $n = \frac{m}{M}$.

Isolation of the variable V:

$$PV = nRT$$

$$\frac{\cancel{P}V}{\cancel{P}} = \frac{nRT}{P}$$

$$V = \frac{nRT}{P}$$

Substitution to solve for V_{CO_2}:

$$V_{CO_2} = \frac{nRT}{P}$$

$$= \frac{23.77859\ \cancel{\text{mol}} \times 8.314 \frac{\cancel{\text{kPa}} \bullet \text{L}}{\cancel{\text{mol}} \bullet \cancel{\text{K}}} \times 273.15\ \cancel{\text{K}}}{101.325\ \cancel{\text{kPa}}}$$

$$= 532.94293\ \text{L}$$

$$= 533\ \text{L}$$

The volume of carbon dioxide gas that is produced is 533 L.

Check Your Solution

The volume of carbon dioxide gas produced seems reasonable given the mole ratio in the balanced chemical equation and the mass of calcium carbonate that reacted. The answer correctly shows three significant digits.

35. Practice Problem (page 560)

When iron rusts, it undergoes a reaction with oxygen to form solid iron(III) oxide. Calculate the volume of oxygen gas at STP that is required to completely react with 52.0 g of iron.

What Is Required?

You need to find the volume of oxygen gas that reacts to produce iron(III) oxide.

What Is Given?

You know that the reactants are iron, Fe(s), and oxygen gas, $O_2(g)$.
You know that the product is iron(III) oxide, $Fe_2O_3(s)$.
You know the mass of the iron, Fe(s): m = 52.0 g
You know the temperature and pressure at STP:
T = 273.15 K
P = 101.325 kPa

Plan Your Strategy

Write the balanced equation for the chemical reaction.
Determine the molar mass of Fe(s) using the atomic mass in the period table.
Calculate the amount in moles of Fe(s) using the relationship $n = \dfrac{m}{M}$.
Use the amount in moles of Fe(s) and the mole ratio from the balanced chemical equation to calculate the amount in moles of O_2(g) that reacts.
Convert the temperature from the Celsius scale to the Kelvin scale.
Use the universal gas constant: $R = 8.314 \dfrac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$
Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable V.
Substitute the given data into the equation and solve for V_{O_2}.

Act on Your Strategy

Balanced chemical equation: $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
Mole ratio: 4 moles 3 moles 2 moles

Molar mass, M, of Fe(s): 55.85 g/mol (from the periodic table)

Amount in moles, n, of Fe(s):

$$\begin{aligned} n_{Fe} &= \frac{m}{M} \\ &= \frac{52.0\ \cancel{\text{g}}}{55.85\ \cancel{\text{g}}/\text{mol}} \\ &= 0.93106\ \text{mol} \end{aligned}$$

Amount in moles, n, of O_2(g):

$$\frac{3\ \text{mol O}_2}{4\ \text{mol Fe}} = \frac{n_{O_2}}{0.93106\ \text{mol Fe}}$$

$$\begin{aligned} n_{O_2} &= \frac{3\ \text{mol O}_2 \times 0.93106\ \cancel{\text{mol Fe}}}{4\ \cancel{\text{mol Fe}}} \\ &= 0.69829\ \text{mol} \end{aligned}$$

Isolation of the variable V:

$$\begin{aligned} PV &= nRT \\ \frac{\cancel{P}V}{\cancel{P}} &= \frac{nRT}{P} \\ V &= \frac{nRT}{P} \end{aligned}$$

Substitution to solve for V_{O_2}:

$$V_{O_2} = \frac{nRT}{P}$$

$$= \frac{0.69829\ \cancel{\text{mol}} \times 8.314 \frac{\cancel{\text{kPa}} \bullet \text{L}}{\cancel{\text{mol}} \bullet \cancel{\text{K}}} \times 273.15\ \cancel{\text{K}}}{101.325\ \cancel{\text{kPa}}}$$

$$= 15.65069\ \text{L}$$

$$= 15.6\ \text{L}$$

The volume of oxygen that is required is 15.6 L.

Check Your Solution

The volume of oxygen gas seems reasonable given the mole ratio in the balanced chemical equation and the mass of iron that reacted. The answer correctly shows three significant digits.

36. Practice Problem (page 560)

Oxygen gas and magnesium react to form 2.43 g of magnesium oxide, MgO(s). What volume of oxygen gas at 94.9 kPa and 25.0°C would be consumed to produce this mass of magnesium oxide?

What Is Required?

You need the volume of oxygen consumed in a reaction with magnesium to produce magnesium oxide.

What Is Given?

You know that the reactants are magnesium, Mg(s), and oxygen gas, O_2(g).
You know that the product is magnesium oxide, MgO(s).
You know the mass of the MgO(s): m = 2.43 g
You know the temperature and pressure:
T = 25.0°C
P = 94.9 kPa

Plan Your Strategy

Write the balanced equation for the chemical reaction.
Determine the molar mass of MgO(s) using the atomic masses from the period table.
Calculate the amount in moles of MgO(s) using the relationship $n = \frac{m}{M}$.
Use the amount in moles of MgO(s) and the mole ratio from the balanced chemical equation to calculate the amount in moles of O_2(g) consumed.
Convert the temperature from the Celsius scale to the Kelvin scale.

Substitution to solve for V_{O_2}:

$$V_{O_2} = \frac{nRT}{P}$$

$$= \frac{0.0301414 \cancel{\text{mol}} \times 8.314 \frac{\cancel{\text{kPa}} \bullet \text{L}}{\cancel{\text{mol}} \bullet \cancel{\text{K}}} \times 298.15 \cancel{\text{K}}}{94.9 \cancel{\text{kPa}}}$$

$$= 0.7873029 \text{ L}$$

$$= 0.787 \text{ L}$$

The volume of oxygen gas that would be consumed is 0.787 L.

Check Your Solution
The volume of oxygen gas seems reasonable given the mole ratio in the balanced chemical equation and the mass of magnesium oxide produced. The answer correctly shows three significant digits.

37. Practice Problem (page 560)
In the semiconductor industry, hexafluoroethane, C_2F_6(g), is used to remove silicon dioxide, SiO_2(s), according to the following chemical equation:

$$2SiO_2(s) + 2C_2F_6(g) + O_2(g) \rightarrow 2SiF_4(g) + 2COF_2(g) + 2CO_2(g)$$

What mass of silicon dioxide reacts with 1.270 L of hexafluoroethane at 0.200 kPa and 400.0°C?

What Is Required?
You need to find the mass of silicon dioxide that reacts with a specific volume of hexafluoroethane.

What Is Given?
You know the balanced chemical equation:
$2SiO_2(s) + 2C_2F_6(g) + O_2(g) \rightarrow 2SiF_4(g) + 2COF_2(g) + 2CO_2(g)$
You know the volume of hexafluoroethane that reacts: V = 1.270 L
You know the temperature and pressure:
T = 400.0°C
P = 0.200 kPa

Plan Your Strategy
Write the balanced equation for the chemical reaction.
Convert the temperature from the Celsius scale to the Kelvin scale.
Use the universal gas constant: $R = 8.314 \frac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$

Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable n.
Substitute the given data into the equation and solve for $n_{C_2F_6}$.
Use the amount in moles of C_2F_6(g) and mole ratio from the balanced chemical equation to calculate the amount in moles of SiO_2(s) that react.
Determine the molar mass of SiO_2(s) using the atomic masses from the periodic table.
Calculate the mass of SiO_2(s) using the relationship $m = n \times M$.

Act on Your Strategy
Balanced chemical equation: $2SiO_2(s) + 2C_2F_6(g) + O_2(g) \rightarrow 2SiF_4(g) + 2COF_2(g) + 2CO_2(g)$ Mole ratio: 2 moles 2 moles 1 mole 2 moles 2 moles 2 moles

Temperature conversion:
$T = 400.0°C + 273.15$
$= 673.15$ K

Isolation of the variable n:

$$PV = nRT$$

$$\frac{PV}{RT} = \frac{n\cancel{RT}}{\cancel{RT}}$$

$$n = \frac{PV}{RT}$$

Substitution to solve for $n_{C_2F_6}$:

$$\begin{aligned} n_{C_2F_6} &= \frac{PV}{RT} \\ &= \frac{0.200\ \cancel{\text{kPa}} \times 1.270\ \cancel{\text{L}}}{8.314 \dfrac{\cancel{\text{kPa}} \bullet \cancel{\text{L}}}{\text{mol} \bullet \cancel{\text{K}}} \times 673.15\ \cancel{\text{K}}} \\ &= 4.53849 \times 10^{-5}\ \text{mol} \end{aligned}$$

Amount in moles, n, of SiO_2(s):

$$\frac{2\ \text{mol SiO}_2}{2\ \text{mol C}_2\text{F}_6} = \frac{n_{\text{SiO}_2}}{4.53849 \times 10^{-5}\ \text{mol C}_2\text{F}_6}$$

$$\begin{aligned} n_{\text{SiO}_2} &= \frac{2\ \text{mol SiO}_2 \times 4.53849 \times 10^{-5}\ \cancel{\text{mol C}_2\text{F}_6}}{2\ \cancel{\text{mol C}_2\text{F}_6}} \\ &= 4.53849 \times 10^{-5}\ \text{mol} \end{aligned}$$

Molar mass, M, of SiO_2(s):

$$\begin{aligned} M_{SiO_2} &= 1M_{Si} + 2M_O \\ &= 1(28.09 \text{ g/mol}) + 1(16.00 \text{ g/mol}) \\ &= 60.09 \text{ g/mol} \end{aligned}$$

Mass, m, of SiO_2(s):

$$\begin{aligned} m_{SiO_2} &= n \times M \\ &= 4.53849 \times 10^{-5} \cancel{\text{mol}} \times 60.09 \text{ g/}\cancel{\text{mol}} \\ &= 2.72717 \times 10^{-3} \text{ g} \\ &= 2.73 \times 10^{-3} \text{ g} \end{aligned}$$

The mass of silicon dioxide that reacts is 2.73×10^{-3} g.

Check Your Solution
The small mass of silicon dioxide seems reasonable given the mole ratio in the balanced chemical equation, the conditions of the reaction, and the volume of hexafluoroethane that reacted. The answer correctly shows three significant digits.

38. Practice Problem (page 560)
What mass of oxygen gas reacts with hydrogen gas to produce 0.62 L of water vapour at 100.0°C and 101.3 kPa?

What Is Required?
You need to find the mass of oxygen gas that reacts with hydrogen to produce a specific volume of water vapour, H_2O(g).

What Is Given?
You know the volume of the water vapour: $V = 0.62$ L
You know the temperature and pressure:
$T = 100.0$°C
$P = 101.3$ kPa

Plan Your Strategy
Convert the temperature from the Celsius scale to the Kelvin scale.
Use the universal gas constant: $R = 8.314 \dfrac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$
Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable n.
Substitute the given data into the equation and solve for n_{H_2O}.

Mass, *m*, of O_2(g):

$$m_{O_2} = n \times M$$
$$= 1.01222 \times 10^{-2}\ \cancel{\text{mol}} \times 32.00\ \text{g}/\cancel{\text{mol}}$$
$$= 0.32391\ \text{g}$$
$$= 0.32\ \text{g}$$

The mass of oxygen that reacts is 0.32 g.

Check Your Solution
The small mass of oxygen seems reasonable, given the mole ratio in the balanced chemical equation and the small volume of water vapour that was produced. The answer correctly shows two significant digits.

39. Practice Problem (page 560)
One method of producing ammonia gas involves the reaction of ammonium chloride, NH_4Cl(aq), with sodium hydroxide, NaOH(aq); water and aqueous sodium chloride are also products of the reaction. During an experiment, 98 mL of ammonia gas was collected using water displacement. If the gas was collected at 20.0°C and 780 mmHg, determine the amount of sodium hydroxide that must have reacted.

What Is Required?
You need to find the amount, in grams, of NaOH(s) that reacts to produce a specific volume of ammonia, NH_3(g), collected over water.

What Is Given?
You know that the reactants are ammonium chloride, NH_4Cl(aq), and sodium hydroxide, NaOH(aq).
You know that the products are water, $H_2O(\ell)$; sodium chloride, NaCl(aq); and ammonia, NH_3(g).
You know the volume of the ammonia gas: V = 98 mL
You know the temperature and the total pressure:
T = 20.0°C
P_{total} = 780 mmHg
Partial pressure of water vapour at 20.0°C: P_{H_2O} = 2.33 kPa

Plan Your Strategy
Write the balanced equation for the chemical reaction.
Convert the temperature from the Celsius scale to the Kelvin scale.
Convert the total pressure from mmHg to kilopascals.

Use Dalton's law of partial pressures to determine the partial pressure of the dry ammonia gas: $P_{\text{total}} = P_{\text{gas}} + P_{\text{water vapour}}$
Convert the volume of ammonia gas to litres.
Use the universal gas constant: $R = 8.314 \dfrac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$
Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable n.
Substitute the given data into the equation and solve for n_{NH_3}.
Use the amount in moles of NH_3(g) and the mole ratio from the balanced chemical equation to calculate the amount in moles of NaOH(s) that reacts.
Determine the molar mass of NaOH(s) using the atomic masses from the periodic table.
Calculate the mass of NaOH(s) using the relationship $m = n \times M$.

Act on Your Strategy
Balanced chemical equation: NH_4Cl(aq) + NaOH(aq) → NaCl(aq) + H_2O(ℓ) + NH_3(g)
Mole ratio: 1 mole 1 mole 1 mole 1 mole 1 mole

Temperature conversion:

$$T = 20.0°\text{C} + 273.15$$
$$= 293.15 \text{ K}$$

Pressure conversion:

$$P_{\text{total}} = 780 \cancel{\text{mmHg}} \times \frac{101.325 \text{ kPa}}{760 \cancel{\text{mmHg}}}$$
$$= 103.99144 \text{ kPa}$$

Partial pressure of the dry ammonia gas:

$$P_{\text{total}} = P_{NH_3} + P_{\text{water vapour}}$$
$$P_{NH_3} = P_{\text{total}} - P_{\text{water vapour}}$$
$$= 103.99144 \text{ kPa} - 2.33 \text{ kPa}$$
$$= 101.66 \text{ kPa}$$

Volume conversion:

$$V = 98 \cancel{\text{mL}} \times 1 \times 10^{-3} \text{ L}/\cancel{\text{mL}}$$
$$= 0.098 \text{ L}$$

Isolation of the variable *n*:

$$PV = nRT$$

$$\frac{PV}{RT} = \frac{n\cancel{RT}}{\cancel{RT}}$$

$$n = \frac{PV}{RT}$$

Substitution to solve for n_{NH_3}:

$$n_{NH_3} = \frac{PV}{RT}$$

$$= \frac{101.66\ \cancel{kPa} \times 0.098\ \cancel{L}}{8.314 \frac{\cancel{kPa} \bullet \cancel{L}}{mol \bullet \cancel{K}} \times 293.15\ \cancel{K}}$$

$$= 4.08767 \times 10^{-3}\ mol$$

Amount in moles, *n*, of NaOH(s):

$$\frac{1\ mol\ NaOH}{1\ mol\ NH_3} = \frac{n_{NaOH}}{4.08767 \times 10^{-3} mol\ NH_3}$$

$$n_{NaOH} = \frac{1\ mol\ NaOH \times 4.08767 \times 10^{-3}\ \cancel{mol\ NH_3}}{1\ \cancel{mol\ NH_3}}$$

$$= 4.08767 \times 10^{-3}\ mol$$

Molar mass, *M*, of NaOH(s):

$$M_{NaOH} = 1M_{Na} + 1M_{O} + 1M_{H}$$

$$= 1(22.99\ g/mol) + 1(16.00\ g/mol) + 1(1.01\ g/mol)$$

$$= 40.00\ g/mol$$

Mass, *m*, of NaOH(s):

$$m_{NaOH} = n \times M$$

$$= 4.08767 \times 10^{-3}\ \cancel{mol} \times 40.00\ g/\cancel{mol}$$

$$= 0.16350\ g$$

$$= 0.16\ g$$

The mass of sodium hydroxide that reacts is 0.16 g.

Check Your Solution
The small mass of sodium hydroxide seems reasonable given the mole ratio in the balanced chemical equation and the volume of ammonia that reacted. The answer correctly shows two significant digits.

40. Practice Problem (page 560)
A student reacts 0.15 g of magnesium metal with excess dilute hydrochloric acid to produce hydrogen gas, which she collects over water. What volume of dry hydrogen gas does she collect over water at 28°C and 101.8 kPa?

What Is Required?
You need to find the volume of hydrogen gas, $H_2(g)$, that is collected over water.

What Is Given?
You know that the reactants are magnesium, Mg(s), and hydrochloric acid, HCl(aq).
You know that the products are magnesium chloride, $MgCl_2(aq)$, and hydrogen gas, $H_2(g)$.
You know the mass of the Mg(s): $m = 0.15$ g
You know the temperature and the total pressure
$T = 28°C$
$P_{total} = 101.8$ kPa
Partial pressure of water vapour at 28°C: $P_{H_2O} = 3.78$ kPa

Plan Your Strategy
Write the balanced equation for the chemical reaction.
Find the molar mass of Mg(s) from the atomic mass in the periodic table.
Calculate the amount in moles of Mg(s) using the relationship $n = \frac{m}{M}$.
Use the amount in moles of Mg(s) and the mole ratio from the balanced chemical equation to calculate the amount in moles of $H_2(g)$ that is produced.
Convert the temperature from the Celsius scale to the Kelvin scale.
Use Dalton's law of partial pressures to determine the partial pressure of the dry hydrogen gas: $P_{total} = P_{gas} + P_{water\ vapour}$
Use the universal gas constant: $R = 8.314 \frac{\text{kPa} \bullet \text{L}}{\text{mol} \bullet \text{K}}$
Use the ideal gas law: $PV = nRT$
Rearrange the equation to isolate the variable V.
Substitute the given data into the equation and solve for V_{H_2}.