Chapter 7

Quantities in Chemical Reactions

Solutions for Practice Problems

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1. **Problem**

Consider the following reaction.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)}$$

(a) Write the ratio of H_2 molecules: O_2 molecules: H_2O molecules.

(b) How many molecules of O_2 are required to react with 100 molecules of H_2, according to your ratio in part (a)?

(c) How many molecules of water are formed when 2478 molecules of O_2 react with H_2?

(d) How many molecules of H_2 are required to react completely with 6.02×10^{23} molecules of O_2?

What Is Required?

(a) You need to give the ratio of molecules in the reaction.

(b) You need to find the number of O_2 molecules that will react with the given number of H_2 molecules.

(c) You need to find the number of water molecules that will form from the given number of reactants.

(d) You need to find the number of H_2 molecules that will react with the given number of O_2 molecules.

What Is Given?

The balanced equation for the reaction is given. The number of molecules of reactants or products is given in each question.

Plan Your Strategy

(a) The ratio is given by the whole number in front of each reactant and product in the balanced equation.

(b) Equate the ratio of number of molecules to the whole number ratio obtained in (a) for hydrogen:oxygen, and solve for the number of oxygen molecules.

(c) As in (b), equate the ratio of the number of molecules to the whole number ratio obtained in (a) for oxygen:water. Solve for the number of water molecules.

(d) As in (b) and (c), equate the ratio of the number of molecules to the whole number ratio obtained in (a) for hydrogen:oxygen. Solve for the number of hydrogen molecules.

Act on Your Strategy

	$2H_{2(g)}$	$+ O_{2(g)}$	$\rightarrow$	$2H_2O_{(\ell)}$
(a)	2	1		2
(b)	100 molecules	50 molecules		
(c)		2478 molecules		4956 molecules
(d)	1.204×10^{24} molecules	6.02×10^{23} molecules		

(a) The ratio is 2: 1: 2

(b) $\frac{2 \text{ molecules } H_2}{1 \text{ molecule } O_2} = \frac{N \text{ molecules } O_2}{100 \text{ molecules } H_2}$

$N = 50$ molecules O_2

(c) $\frac{1 \text{ molecule } O_2}{2 \text{ molecules } H_2O} = \frac{2478 \text{ molecules } O_2}{N \text{ molecules } H_2O}$

$N = 4956$ molecules H_2O

(d) $\frac{2 \text{ molecules } H_2}{1 \text{ molecule } O_2} = \frac{N \text{ molecules } H_2}{6.02 \times 10^{23}}$

$N = 1.204 \times 10^{24}$ molecules H_2

Check Your Solution

(a) Try drawing the molecules of hydrogen and oxygen and fitting them into water molecules. You will see that it takes 2 full molecules of gaseous H_2 and one full molecule of gaseous O_2 to form 2 full molecules of water.

(b) The ratio of hydrogen to oxygen in this reaction is always 2:1, no matter what. Dividing both 100 molecules of hydrogen and 50 molecules of oxygen by 50 will give the lowest ratio of 2:1. Your answer is consistent.

(c) The ratio of oxygen to water in this reaction is always 1:2, no matter what. Dividing both the 2478 molecules of oxygen and 4956 molecules of water by 2478 will give the lowest ratio of 1:2. Your answer is consistent.

(d) The ratio of hydrogen to oxygen in this reaction is always 2:1, no matter what. Dividing both 1.204×10^{24} molecules of hydrogen and 6.02×10^{23} molecules of oxygen by 6.02×10^{23} will give the lowest ratio of 2:1. Your answer is consistent.

2. Problem

Iron reacts with chlorine gas to form iron(III) chloride, $FeCl_3$.

$$2Fe_{(s)} + 3Cl_{2(g)} \rightarrow 2FeCl_{3(s)}$$

(a) How many atoms of Fe are needed to react with three molecules of Cl_2?

(b) How many molecules of $FeCl_3$ are formed when 150 atoms of Fe react with sufficient Cl_2?

(c) How many Cl_2 molecules are needed to react with 1.204×10^{24} atoms of Fe?

(d) How many molecules of $FeCl_3$ are formed when 1.806×10^{24} molecules of Cl_2 react with sufficient Fe?

What Is Required?

(a) You need to give the ratio of Fe to Cl_2 molecules in the reaction.

(b) You need to find the number of $FeCl_3$ molecules that will form from the given number of Fe atoms.

(c) You need to find the number of Cl_2 molecules that will react with the given number of Fe.

(d) You need to find the number of $FeCl_3$ molecules that will form from the given number of Cl_2 molecules.

What Is Given?

The balanced equation for the reaction is given. The number of molecules of reactants or products is given in each question.

Plan Your Strategy

(a) The ratio is given by the whole number in front of each reactant and product in the balanced equation.

(b) Equate the ratio of number of molecules to the whole number ratio for Fe:$FeCl_3$, and solve for the number of $FeCl_3$ molecules.

(c) As in (b), equate the ratio of the number of molecules to the whole number ratio for Fe:Cl_2. Solve for the number of Cl_2 molecules.

(d) As in (b) and (c), equate the ratio of the number of molecules to the whole number ratio for Cl_2:$FeCl_3$. Solve for the number of $FeCl_3$ molecules.

Act on Your Strategy

	$2Fe_{(s)}$	$+ 3Cl_{2(g)}$	$\rightarrow$	$2FeCl_{3(s)}$
(a)	2 atoms	3 molecules		
(b)	150 atoms			150 molecules
(c)	1.204×10^{24} atoms	1.806×10^{24} molecules		
(d)		1.806×10^{24} molecules		1.204×10^{24} molecules

(a) According to the balanced equation, only two atoms of Fe are required to react with three molecules of Cl_2.

(b) $\dfrac{2 \text{ atoms of Fe}}{2 \text{ formula units of } FeCl_3} = \dfrac{150 \text{ atoms of Fe}}{N \text{ formula units of } FeCl_3}$

$N = 150$ formula units of $FeCl_3$

(c) $\dfrac{2 \text{ atoms of Fe}}{3 \text{ molecules of } Cl_2} = \dfrac{1.204 \times 10^{24} \text{ atoms of Fe}}{N \text{ molecules of } Cl_2}$

$N = 1.806 \times 10^{24}$ molecules of Cl_2

(d) $\dfrac{3 \text{ molecules of } Cl_2}{2 \text{ molecules of } FeCl_3} = \dfrac{1.806 \times 10^{24} \text{ molecules of } Cl_2}{N \text{ molecules of } FeCl_3}$

$N = 1.204 \times 10^{24}$ molecules of $FeCl_3$

Check Your Solution

(a) Try drawing the atoms and molecules of iron and chloride gas and fitting them into iron(III) chloride molecules. You will see that it takes 2 atoms of Fe and one 3 molecules of gaseous Cl_2 to form 2 full molecules of $FeCl_3$.

(b) The ratio of iron to iron(III) chloride in this reaction is always 2:2, no matter what, which is 1:1 in the lowest ratio. Dividing both the 150 atoms of Fe and 150 molecules of $FeCl_3$ by 150 will give the lowest ratio of 1:1. Your answer is consistent.

(c) The ratio of Fe to Cl_2 in this reaction is always 2:3, no matter what, which is 1:1.5 in the lowest ratio. Dividing both the 1.204×10^{24} atoms of Fe and the 1.806×10^{24} molecules of Cl_2 by 1.204×10^{24} will give the lowest ratio of 1:1.5. Your answer is consistent.

(d) The ratio of Cl_2 to $FeCl_3$ in this reaction is always 3:2, no matter what, which is 1.5:1 in the lowest ratio. Dividing both 1.806×10^{24} molecules of Cl_2 and 1.204×10^{24} molecules of $FeCl_3$ by 1.204×10^{24} will give the lowest ratio of 1.5:1. Your answer is consistent.

3. Problem

Consider the following reaction.

$$Ca(OH)_{2(aq)} + 2HCl_{(aq)} \rightarrow CaCl_2 + 2H_2O_{(\ell)}$$

(a) How many formula units of calcium chloride, $CaCl_2$, would be produced by 6.7×10^{25} molecules of hydrochloric acid, HCl?

(b) How many molecules of water would be produced in the reaction in part (a)?

What Is Required?

(a) You need to find the number of $CaCl_2$ formula units produced by the given amount of HCl.

(b) You need to find the number of water molecules produced by the given amount of HCl.

What Is Given?

The balanced equation for the reaction is given. The number of molecules of HCl is given.

Plan Your Strategy

(a) Equate the ratio of number of molecules to the whole number ratio for $HCl:CaCl_2$ and solve for the number of $CaCl_2$ formula units.

(b) As in (a), equate the ratio of the number of molecules to the whole number ratio for $HCl:H_2O$. Solve for the number of H_2O molecules.

Act on Your Strategy

	$Ca(OH)_{2(aq)}$	$+ 2HCl_{(aq)}$	$\rightarrow$	$CaCl_{2(g)}$	$+ 2H_2O_{(\ell)}$
mole ratio	1	2		1	2
molecule or formula unit ratio	6.02×10^{23}	1.2×10^{24}		6.02×10^{23}	1.2×10^{24}
given (a)		6.7×10^{25}		N	
given (b)		6.7×10^{25}			N

(a) $N = (6.7 \times 10^{25}) \times \frac{6.02 \times 10^{23}}{1.2 \times 10^{24}} = 3.35 \times 10^{25}$ formula units of $CaCl_2$

(b) $N = (6.7 \times 10^{25}) \times \frac{1.2 \times 10^{24}}{1.2 \times 10^{24}} = 6.7 \times 10^{25}$ molecules H_2O

Check Your Solution

(a) The mol ratio of HCl to $CaCl_2$ in this reaction is always 2:1, no matter what. Dividing both the 6.7×10^{25} molecules of HCl and 3.35×10^{25} molecules of water by 3.35×10^{25} will give the lowest ratio of 2:1. Your answer is consistent

(b) The mol ratio of HCl to water in this reaction is always 2:2, no matter what, which is 1:1 in the lowest ratio. Dividing both the 6.7×10^{25} of HCl and 6.7×10^{25} molecules of water by 6.7×10^{25} will give the lowest ratio of 1:1. Your answer is consistent.

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4. **Problem**

Aluminum bromide can be prepared by reacting small pieces of aluminum foil with liquid bromine at room temperature. The reaction is accompanied by flashes of red light.

$$2Al_{(s)} + 3Br_{2(\ell)} \rightarrow 2AlBr_{3(s)}$$

How many moles of Br_2 are needed to produce 5 mol of $AlBr_3$, if sufficient Al is present?

What Is Required?

You need to find the number of moles of Br_2 to produce the given number of moles of product.

What Is Given?

The balanced equation for the reaction is given. The number of moles of product is given.

Plan Your Strategy

From the balanced equation, determine the mol ratio of reactants to products, given by the whole number in front of each molecule. Equate the equation mol ratio of Br_2:$AlBr_3$ to the same ratio using the number of mol of $AlBr_3$ given, and solve for the number of moles of Br_2.

Act on Your Strategy

	$2Al_{(s)}$	$+ 3Br_{2(l)}$	$\rightarrow$	$2AlBr_{3(s)}$
mole ratio	2	3		2
given		n		5 mol

$$\frac{3 \text{ mol } Br_2}{2 \text{ mol } AlBr_3} = \frac{n \text{ mol } Br_2}{5 \text{ mol } AlBr_3}$$

$$n = 5 \text{ mol } AlBr_3 \times \frac{3 \text{ mol } Br_2}{2 \text{ mol } AlBr_3} = 7.5 \text{ mol } Br_2$$

Check Your Solution

The mol ratio of Br_2 to $AlBr_3$ in this reaction is always 3:2, no matter what, which is 1.5:1 in the lowest ratio. Dividing both the 7.5 mol of Br_2 and 5.0 mol of $AlBr_3$ by 5.0 will give the lowest ratio of 1.5:1. Your answer is consistent.

5. Problem

Hydrogen cyanide gas, $HCN_{(g)}$, is used to prepare clear, hard plastics, such as Plexiglas™. Hydrogen cyanide is formed by reacting ammonia, NH_3, with oxygen and methane, CH_4.

$$2NH_{3(g)} + 3O_{2(g)} + 2CH_{4(g)} \rightarrow 2HCN_{(g)} + 6H_2O_{(g)}$$

(a) How many moles of O_2 are needed to react with 1.2 mol of NH_3?

(b) How many moles of H_2O can be expected from the reaction of 12.5 mol of CH_4? Assume that sufficient NH_3 and O_2 are present.

What Is Required?

(a) You need to find the number of moles of O_2 that will react with the given number of moles of ammonia.

(b) You need to find the number of moles of water that will be produced from the given number of moles of methane.

What Is Given?

The balanced equation for the reaction is given. The number of moles of reactants is given.

Plan Your Strategy

(a) From the balanced equation, determine the mol ratio of reactants to products, given by the whole number in front of each molecule. Equate the equation mol ratio of NH_3:O_2 to the same ratio using the number of mol of NH_3 given, and solve for O_2.

(b) As in (a), equate the ratio of the number of moles given to the whole number ratio for CH_4:H_2O. Solve for the number of moles of H_2O.

Act on Your Strategy

	$2NH_{3(g)}$	$+ 3O_{2(g)}$	$2CH_{4(g)}$	$\rightarrow$	$2HCN_{(g)}$	$+ 6H_2O_{(g)}$
Mole ratio	2	3	2		2	6
given	1.2 mol	n				
given			12.5 mol			n

(a) $$\frac{2 \text{ mol } NH_3}{3 \text{ mol } O_2} = \frac{1.2 \text{ mol } NH_3}{n \text{ mol } O_2}$$

$$n = 1.2 \text{ mol } NH_3 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } NH_3} = 1.8 \text{ mol } O_2$$

(b) $$\frac{2 \text{ mol } CH_4}{6 \text{ mol } H_2O} = \frac{12.5 \text{ mol } CH_4}{n \text{ mol } H_2O}$$

$$n = 12.5 \text{ mol } CH_4 \times \frac{6 \text{ mol } H_2O}{2 \text{ mol } CH_4} = 37.5 \text{ mol } H_2O$$

Check Your Solution

(a) The mol ratio of NH_3:O_2 in this reaction is always 2:3, no matter what, the lowest ratio of which is 1:1.5. Dividing both the 1.2 mol of NH_3 and 1.8 mol of O_2 by 1.2 will give the lowest ratio of 1.5:1. Your answer is consistent.

(b) The mol ratio of CH_4:water in this reaction is always 2:6, no matter what, which is 1:3 in the lowest ratio. Dividing both the 12.5 mol of CH_4 and 37.5 mol of water by 12.5 will give the lowest ratio of 1:3. Your answer is consistent.

6. Problem

Ethane gas, C_2H_6, is present in small amounts in natural gas. It undergoes complete combustion to produce carbon dioxide and water.

$$2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(\ell)}$$

(a) How many moles of O_2 are required to react with 13.9 mol of C_2H_6?

(b) How many moles of H_2O would be produced by 1.40 mol of O_2 and sufficient ethane?

What Is Required?

(a) You need to find the number of moles of O_2 that will react with the given number of moles of ethane.

(b) You need to find the number of moles of water that will be produced from the given number of moles of oxygen.

What Is Given?

The balanced equation for the reaction is given. The number of moles of reactants is given.

Plan Your Strategy

(a) From the balanced equation, determine the mol ratio of reactants to products, given by the whole number in front of each molecule. Equate the equation mol ratio of C_2H_6:O_2 to the same ratio using the number of mol of ethane given, and solve for O_2.

(b) As in (a), equate the ratio of the number of moles given to the whole number ratio for O_2:H_2O. Solve for the number of moles of H_2O.

Act on Your Strategy

	$2C_2H_{6(g)}$	$+ 7O_{2(g)}$	$\rightarrow$	$4CO_{2(g)}$	$+ 6H_2O_{(g)}$
mole ratio	2	7		4	6
given	13.9 mol	n			
given		1.40 mol			n

(a) $$\frac{2 \text{ mol } C_2H_6}{7 \text{ mol } O_2} = \frac{13.9 \text{ mol } C_2H_6}{n \text{ mol } O_2}$$

$$n = 13.9 \text{ mol } C_2H_6 \times \frac{7 \text{ mol } O_2}{2 \text{ mol } C_2H_6} = 48.7 \text{ mol } O_2$$

(b) $$\frac{7 \text{ mol } O_2}{6 \text{ mol } H_2O} = \frac{1.4 \text{ mol } O_2}{n \text{ mol } O_2}$$

$$n = 1.4 \text{ mol } O_2 \times \frac{6 \text{ mol } H_2O}{7 \text{ mol } O_2} = 1.2 \text{ mol } H_2O$$

Check Your Solution

(a) The mol ratio of C_2H_6:O_2 in this reaction is always 2:7, no matter what, the lowest ratio of which is 1:3.5. Dividing both the 13.9 mol of C_2H_6 and 48.7 mol of O_2 by 13.9 will give the lowest ratio of 1:3.5. Your answer is consistent.

(b) The mol ratio of O_2:water in this reaction is always 7:6, no matter what, which is 1.67:1 in the lowest ratio. Dividing both the 1.40 mol of O_2 and 1.20 mol of water by 1.20 will give the lowest ratio of 1.67:1. Your answer is consistent.

7. Problem

Magnesium nitride reacts with water to produce magnesium hydroxide and ammonia gas, NH_3 according to the balanced chemical equation

$$Mg_3N_{2(s)} + 6H_2O_{(\ell)} \rightarrow 3Mg(OH)_{2(s)} + 2NH_{3(g)}$$

(a) How many molecules of water are required to react with 2.3 mol Mg_3N_2?

(b) How many molecules of $Mg(OH)_2$ will be expected in part (a)?

What Is Required?

(a) You need to find the number of molecules of water that will react with the magnesium nitride.

(b) You need to find the number of molecules of magnesium hydroxide that will form from the given amount of magnesium nitride.

What Is Given?

The balanced equation for the reaction is given. The number of moles of magnesium nitride is given. Avogadro's number, $N_A = 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy

(a) From the balanced equation, determine the mol ratio of reactants to products, given by the whole number in front of each molecule. Equate the equation mol ratio of Mg_3N_2:H_2O to the same ratio using the number of mol of magnesium nitride given, and solve for the number of moles of O_2. Multiply this by Avogadro's number to obtain the number of molecules needed.

(b) As in (a), equate the ratio of the number of moles given to the whole number ratio for Mg_3N_2:$Mg(OH)_2$. Solve for the number of moles of $Mg(OH)_2$. Multiply this by Avogadro's number to obtain the number of molecules produced.

Act on Your Strategy

	$Mg_3N_{2(s)}$	$+ 6H_2O_{(\ell)}$	$\rightarrow$	$3Mg(OH)_{2(s)}$	$+ 2NH_{3(g)}$
mole ratio	1	6		3	2
given	2.3 mol	*n*		*n*	

(a) $\dfrac{1 \text{ mol } Mg_3N_2}{6 \text{ mol } H_2O} = \dfrac{2.3 \text{ mol } Mg_3N_2}{n \text{ mol } H_2O}$

$n = 2.3 \text{ mol } Mg_3N_2 \times \dfrac{6 \text{ mol } H_2O}{1 \text{ mol } Mg_3N_2} = 13.8 \text{ mol } H_2O$

$N = 13.8 \text{ mol } H_2O \times (6.02 \times 10^{23}) \text{ molecules/mol}$

$= 4.2 \times 10^{24} \text{ molecules } H_2O$

(b) $\dfrac{1 \text{ mol } Mg_3N_2}{3 \text{ mol } Mg(OH)_2} = \dfrac{2.3 \text{ mol } Mg_3N_2}{n \text{ mol } Mg(OH)_2}$

$n = 2.3 \text{ mol } Mg_3N_2 \times \dfrac{3 \text{ mol } Mg(OH)_2}{1 \text{ mol } Mg_3N_2} = 6.9 \text{ mol } Mg(OH)_2$

$N = 6.9 \text{ mol} \times (6.02 \times 10^{23}) \text{ molecules/mol}$

$= 4.2 \times 10^{24} \text{ molecules } Mg(OH)_2$

Check Your Solution

(a) The mol ratio of $Mg_3N_2{:}H_2O$ in this reaction is always 1:6, no matter what. Dividing both the 2.3 mol of Mg_3N_2 and 13.8 mol of H_2O by 2.3 will give the lowest ratio of 1:6. Your answer is consistent.

(b) The mol ratio of $Mg_3N_2{:}Mg(OH)_2$ in this reaction is always 1:3, no matter what. Dividing both the 2.3 mol of Mg_3N_2 and 6.9 mol of $Mg(OH)_2$ by 2.3 will give the lowest ratio of 1:3. Your answer is consistent.

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8. Problem

Vanadium can form several different compounds with oxygen, including V_2O_5, VO_2, and V_2O_3.

(a) How many moles of V are needed to produce 7.47 mol of VO_2? Assume that sufficient O_2 is present.

(b) How many moles of V are needed to react with 5.39 mol of O_2 to produce V_2O_3?

What Is Required?

(a) You need to find the number of moles of V to produce the given amount of VO_2.

(b) You need to find the number of moles of V that will react with the given amount of oxygen to form V_2O_3.

What Is Given?

The number of moles of product or reactant is given. The type of reactants and products that make up each reaction are given.

Plan Your Strategy

(a) Write out the full reaction and balance the equation. Determine the mole ratio of reactant to products, which is the whole number in front of each reactant and product. Equate the ratio of $V{:}VO_2$ in the equation to the given amounts, and solve for the number of moles of V.

(b) As in (a), write out the full reaction and balance the equation. Determine the mole ratio of reactant to products. Equate the ratio of $V{:}O_2$ in the equation to the given amounts, and solve for the number of moles of V.

Act on Your Strategy

(a)

	$V_{(s)}$	$+\ O_{2(g)}$	$\rightarrow$	$VO_{2(s)}$
mole ratio	1	1		1
given	n			7.47 mol

$$\frac{1 \text{ mol V}}{1 \text{ mol } VO_2} = \frac{n \text{ mol V}}{7.47 \text{ mol } VO_2}$$

$$n = 7.47 \text{ mol } VO_2 \times \frac{1 \text{ mol V}}{1 \text{ mol } VO_2} = 7.47 \text{ mol V}$$

(b)

	$4V_{(s)}$	$+\ 3O_{2(g)}$	$\rightarrow$	$2V_2O_{3(s)}$
mole ratio	4	3		2
given	n	5.39 mol		

$$\frac{4 \text{ mol V}}{3 \text{ mol } O_2} = \frac{n \text{ mol V}}{5.39 \text{ mol } O_2}$$

$$n = 5.39 \text{ mol } O_2 \times \frac{4 \text{ mol V}}{3 \text{ mol } O_2} = 7.19 \text{ mol V}$$

Check Your Solution

(a) The mol ratio of $V:VO_2$ in this reaction is always 1:1, no matter what. Dividing both the 7.47 mol of V and 7.47 mol of VO_2 by 7.47 will give the lowest ratio of 1:1. Your answer is consistent.

(b) The mol ratio of $V:O_2$ in this reaction is always 4:3, no matter what, which is 1.3:1 in the lowest ratio. Dividing both the 7.19 mol of V and 5.39 mol of O_2 by 5.39 will give the lowest ratio of 1.3:1 Your answer is consistent.

9. Problem

Nitrogen, N_2, can combine with oxygen, O_2, to form several different oxides of nitrogen. These oxides include NO_2, No, and N_2O.

(a) How many moles of O_2 are required to react with 9.35×10^{-2} moles of N_2 to form N_2O?

(b) How many moles of O_2 are required to react with 9.35×10^{-2} moles of N_2 to form NO_2?

What Is Required?

(a) You need to find the number of moles of oxygen needed to produce N_2O from the given amount of nitrogen.

(b) You need to find the number of moles of oxygen needed to produce NO_2 from the given amount of nitrogen.

What Is Given?

The number of moles of nitrogen is given. The type of reactants and products that make up each reaction are given.

Plan Your Strategy

(a) Write out the full reaction and balance the equation. Determine the mole ratio of reactant to products, which is the whole number in front of each reactant and product. Equate the ratio of $N_2:O_2$ in the equation to the given amounts, and solve for the number of moles of O_2.

(b) As in (a), write out the full reaction and balance the equation. Determine the mole ratio of reactant to products. Equate the ratio of $N_2:O_2$ in the equation to the given amounts, and solve for the number of moles of O_2.

Act on Your Strategy

(a)

	$2N_{2(g)}$	$+ O_{2(g)}$	$\rightarrow$	$2N_2O_{(g)}$
mole ratio	2	1		2
given	$(9.35)(10^{-2})$ mol	n		

$$\frac{2 \text{ mol } N_2}{1 \text{ mol } O_2} = \frac{9.35 \times 10^{-2} \text{ mol } N_2}{n \text{ mol } O_2}$$

$$n = (9.35 \times 10^{-2}) \text{ mol } N_2 \times \frac{1 \text{ mol } O_2}{1 \text{ mol } N_2} = 4.68 \times 10^{-2} \text{ mol } O_2$$

(b)

	$N_{2(g)}$	$+ 2O_{2(g)}$	$\rightarrow$	$2NO_{2(g)}$
mole ratio	1	2		2
given	$(9.35)(10^{-2})$ mol	n		

$$\frac{1 \text{ mol } N_2}{2 \text{ mol } O_2} = \frac{9.35 \times 10^{-2} \text{ mol } N_2}{n \text{ mol } O_2}$$

$$n = 9.35 \times 10^{-2} \text{ mol } N_2 \times \frac{2 \text{ mol } O_2}{1 \text{ mol } N_2} = 0.19 \text{ mol } O_2$$

Check Your Solution

(a) The mol ratio of N_2:O_2 in this reaction is always 2:1, no matter what. Dividing both the 9.35×10^{-2} mol of N_2 and 4.68×10^{-2} mol of O_2 by 4.68×10^{-2} will give the lowest ratio of 2:1. Your answer is consistent.

(b) The mol ratio of N_2:O_2 in this reaction is always 1:2, no matter what. Dividing both the 9.35×10^{-2} mol of N_2 and 0.19 mol of O_2 by 9.35×10^{-2} will give the lowest ratio of 1:2 Your answer is consistent.

10. Problem

When heated in a nickel vessel to 400°C, xenon can be made to react with fluorine to produce colourless crystals of xenon tetrafluoride.

(a) How many moles of fluorine gas, F_2, would be required to react with 3.54×10^{-1} mol of xenon?

(b) Under somewhat similar reaction conditions, xenon hexafluoride can also be obtained. How many moles of fluorine would be required to react with the amount of xenon given in part (a) to produce xenon hexafluoride?

What Is Required?

(a) You need to find the number of moles of F_2 that will react with the given amount of xenon to form xenon tetrafluoride.

(b) You need to find the number of moles of F_2 that will react with the given amount of xenon to form xenon hexafluoride.

What Is Given?

The number of moles of xenon is given. The type of reactants and products that make up each reaction are given.

Plan Your Strategy

(a) Write out the full reaction and balance the equation. Determine the mole ratio of reactant to products, which is the whole number in front of each reactant and product. Equate the ratio of Xe:F_2 in the equation to the given amounts, and solve for the number of moles of F_2.

(b) As in (a), write out the full reaction and balance the equation. Determine the mole ratio of reactant to products. Equate the ratio of Xe:F_2 in the equation to the given amounts, and solve for the number of moles of F_2.

Act on Your Strategy

(a)

	$Xe_{(g)}$	$+ 2F_{2(g)}$	$\rightarrow$	$XeF_{4(s)}$
mole ratio	1	2		1
given	3.45×10^{-1} mol	n		

$$\frac{1 \text{ mol Xe}}{2 \text{ mol } F_2} = \frac{3.54 \times 10^{-1} \text{ mol Xe}}{n \text{ mol } F_2}$$

$$n = (3.54 \times 10^{-1}) \text{ mol Xe} \times \frac{2 \text{ mol } F_2}{1 \text{ mol Xe}} = 7.08 \times 10^{-1} \text{ mol } F_2$$

(b)

	$Xe_{(g)}$	$+ 3F_{2(g)}$	$\rightarrow$	$XeF_{6(s)}$
mole ratio	1	3		1
given	3.45×10^{-1} mol	n		

$$\frac{1 \text{ mol Xe}}{3 \text{ mol } F_2} = \frac{3.54 \times 10^{-1} \text{ mol Xe}}{n \text{ mol } F_2}$$

$$n = (3.54 \times 10^{-1}) \text{ mol Xe} \times \frac{3 \text{ mol } F_2}{1 \text{ mol Xe}} = 1.06 \text{ mol } F_2$$

Check Your Solution

(a) The mol ratio of $Xe:F_2$ in this reaction is always 1:2, no matter what. Dividing both the 0.345 mol of Xe and 0.708 mol of F_2 by 0.345 will give the lowest ratio of 1:2. Your answer is consistent.

(b) The mol ratio of $Xe:F_2$ in this reaction is always 1:3, no matter what. Dividing both the 0.345 mol of Xe and 1.06 mol of F_2 by 0.345 will give the lowest ratio of 1:3. Your answer is consistent.

Solutions for Practice Problems

Student Textbook pages 244–246, 248–249

11. Problem

Ammonium sulfate, $(NH_4)_2SO_4$, is used as a source of nitrogen in some fertilizers. It reacts with sodium hydroxide to produce sodium sulfate, water, and ammonia.

$$(NH_4)_2SO_{4(s)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2NH_{3(g)} + 2H_2O_{(\ell)}$$

What mass of sodium hydroxide is required to react completely with 15.4 g of $(NH_4)_2SO_4$?

What Is Required?

You need to find the mass of NaOH that will complete the reaction with the given amount of ammonium sulfate.

What Is Given?

The mass of the ammonium sulfate is given. The balanced equation is given.

Plan Your Strategy

Step 1 Convert the given mass (m) of ammonium sulfate to the number of moles (n) of ammonium sulfate, using its molar mass (M). Use the formula $n = m / M$

Step 2 Determine the mole ratio of NaOH to $(NH_4)_2SO_4$ from the balanced equation. Using the result in Step 1, solve for the number of moles of NaOH.

Step 3 Convert the number of moles of NaOH to mass using its molar mass. Use the formula $m = n \times M$ of the NaOH.

Act on Your Strategy

	$(NH_4)_2SO_{4(s)}$	$+ 2NaOH_{(aq)}$	$\rightarrow$	$Na_2SO_{4(aq)}$	$+ NH_{3(g)}$	$+ 2H_2O_{(\ell)}$
mole ratio	1	2		1	1	2
molar mass	132.17 g/mol	40.0 g/mol		142.05 g/mol	17.04 g/mol	18.02 g/mol
given	15.4 g	m				

Step 1 $n = \dfrac{15.4\ \cancel{g}\ (NH_4)_2SO_4}{132.17\ \cancel{g}/mol} = 0.1165\ mol\ (NH_4)_2SO_4$

Step 2 $\dfrac{1\ mol\ (NH_4)_2SO_4}{2\ mol\ NaOH} = \dfrac{0.1165\ mol\ (NH_4)_2SO_4}{n\ NaOH}$

$n = 0.1165\ mol\ (NH_4)_2SO_4 \times \dfrac{2\ mol\ NaOH}{1\ mol\ (NH_4)_2SO_4} = 0.233\ mol\ NaOH$

Step 3 $m = 0.233\ \cancel{mol}\ NaOH \times 40.0\ g/\cancel{mol} = 9.32\ g\ NaOH$

Therefore, the mass of NaOH that will react completely with the ammonium sulfate is 9.32 g.

Check Your Solution

The mol ratio of $(NH_4)_2SO_4$:NaOH in this reaction is always 1:2, no matter what. Dividing both the 0.1165 mol of $(NH_4)_2SO_4$ and 0.233 mol of NaOH by 0.1165 will give the lowest ratio of 1:2. Your answer is consistent.

12. Problem

Iron(III) oxide, also known as rust, can be removed from iron by reacting it with hydrochloric acid to produce iron(III) chloride and water.

$$Fe_2O_{3(s)} + 6HCl_{(aq)} \rightarrow 2FeCl_{3(aq)} + 3H_2O_{(\ell)}$$

What mass of hydrogen chloride is required to react with 1.00×10^2 g of rust?

What Is Required?

You need to find the mass of HCl that will complete the reaction with the given amount of rust.

What Is Given?

The mass of the HCl is given. The balanced equation is given.

Plan Your Strategy

Step 1 Convert the given mass (*m*) of rust to the number of moles (*n*) of rust, using its molar mass (*M*). Use the formula $n = m \,/\, M$

Step 2 Determine the mole ratio of rust to HCl from the balanced equation. Using the result in Step 1, solve for the number of moles of HCl.

Step 3 Convert the number of moles of HCl to mass using its molar mass. Use the formula $m = n \times M$ of the HCl.

Act on Your Strategy

	$Fe_2O_{3(s)}$	$+\ 6HCl_{(aq)}$	$\rightarrow$	$2FeCl_{3(aq)}$	$+\ 3H_2O_{(\ell)}$
mole ratio	1	6		2	3
molar mass	159.7 g/mol	36.46 g/mol		162.2 g/mol	18.02 g/mol
given	$(1.00)(10^2)$ g	*m*			

Step 1 $n = \dfrac{(1.00)(10^2)\ \cancel{g}\ Fe_2O_3}{159.7\ \cancel{g}/mol} = 0.626\ mol\ Fe_2O_3$

Step 2 $\dfrac{1\ mol\ Fe_2O_3}{6\ mol\ HCl} = \dfrac{0.626\ mol\ Fe_2O_3}{n\ mol\ HCl}$

$n = 0.626\ mol\ Fe_2O_3 \times \dfrac{6\ mol\ HCl}{1\ mol\ Fe_2O_3} = 3.756\ mol\ HCl$

Step 3 $m = 3.756\ \cancel{mol}\ HCl \times 36.46\ g/\cancel{mol} = 137.0\ g\ HCl$

Therefore, the mass of HCl needed to bring this reaction to completoin is 137.0 g.

Check Your Solution

The mol ratio of Fe_2O_3:HCl in this reaction is always 1:6, no matter what. Dividing both the 0.626 mol of rust and 3.756 mol of HCl by 0.626 will give the lowest ratio of 1:6. Your answer is consistent.

13. Problem

Iron reacts slowly with hydrochloric acid to produce iron(III) chloride and hydrogen gas.

$$Fe_{(s)} + 2HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$$

What mass of HCl is required to react with 3.56 g of iron?

What Is Required?

You need to find the mass of HCl that will complete the reaction with the given amount of iron.

What Is Given?

The mass of the iron is given. The balanced equation is given.

Plan Your Strategy

Step 1 Convert the given mass (*m*) of iron to the number of moles (*n*)of iron, using its molar mass (*M*). Use the formula $n = m \,/\, M$

Step 2 Determine the mole ratio of Fe to HCl from the balanced equation. Using the result in Step 1, solve for the number of moles of HCl.
Step 3 Convert the number of moles of HCl to mass using its molar mass. Use the formula $m = n \times M$ of the HCl.

Act on Your Strategy

	$Fe_{(s)}$	$+ 2HCl_{(aq)}$	$\rightarrow$	$FeCl_{2(aq)}$	$+ H_{2(g)}$
mole ratio	1	2		1	1
molar mass	55.85 g/mol	36.46 g/mol		126.75 g/mol	2.02 g/mol
given	3.56 g	m			

Step 1 $n = \dfrac{3.56\ \cancel{g}\ Fe}{55.85\ \cancel{g}/mol} = 0.0637$ mol Fe
Step 2 $\dfrac{1\ mol\ Fe}{2\ mol\ HCl} = \dfrac{0.0637\ mol\ Fe}{n\ mol\ HCl}$
$n = 0.0637$ mol Fe $\times \dfrac{2\ mol\ HCl}{1\ mol\ Fe} = 0.1274$ mol HCl
Step 3 $m = 0.1274\ \cancel{mol}$ HCl $\times$ 36.46 g/$\cancel{mol}$ = 4.65 g HCl
Therefore, the mass of HCl needed to react fully with the Fe is 4.65 g.

Check Your Solution
The mol ratio of Fe:HCl in this reaction is always 1:2, no matter what. Dividing both the 0.0637 mol of Fe and 0.1274 mol of HCl by 0.0637 will give the lowest ratio of 1:2. Your answer is consistent.

14. Problem
Dinitrogen pentoxide is a white solid. When heated it decomposes to produce nitrogen dioxide oxygen.

$$N_2O_{5(s)} \rightarrow 2NO_{2(g)} + O_{2(g)}$$

How many grams of oxygen gas will be produced in this reaction when 2.34 g of NO_2 are made?

What Is Required?
You need to find the mass of oxygen that will be produced together with the given amount of NO_2.

What Is Given?
The mass of the NO_2 product is given. The balanced equation is given.

Plan Your Strategy
Step 1 Convert the given mass (m) of NO_2 to the number of moles (n) of NO_2, using its molar mass (M). Use the formula $n = m\ /\ M$
Step 2 Determine the mole ratio of NO_2 to O_2 from the balanced equation. Using the result in Step 1, solve for the number of moles of O_2.
Step 3 Convert the number of moles of O_2 to mass using its molar mass. Use the formula $m = n \times M$ of the O_2.

Act on Your Strategy

	$N_2O_{5(s)}$	$\rightarrow$	$2NO_{2(g)}$	$+ O_{2(g)}$
mole ratio	1		2	1
molar mass	108.02 g/mol		46.01 g/mol	32.00 g/mol
given			2.34 g	m

Step 1 $n = \dfrac{2.34\ \cancel{g}\ NO_2}{46.01\ \cancel{g}/mol} = 0.0508$ mol NO_2
Step 2 $\dfrac{2\ mol\ NO_2}{1\ mol\ O_2} = \dfrac{0.0508\ mol\ NO_2}{n\ mol\ O_2}$

$$n = 0.0508 \text{ mol } NO_2 \times \frac{1 \text{ mol } O_2}{2 \text{ mol } NO_2} = 0.0254 \text{ mol } O_2$$

Step 3 $m = 0.0254 \cancel{\text{mol}}\ O_2 \times 32.00 \text{ g}/\cancel{\text{mol}} = 0.813 \text{ g } O_2$

Therefore, the mass of O_2 that will be produced is 0.813 g.

Check Your Solution

The mol ratio of NO_2:O_2 in this reaction is always 2:1, no matter what. Dividing both the 0.0508 mol of NO_2 and 0.0254 mol of O_2 by 0.0254 will give the lowest ratio of 2:1. Your answer is consistent.

15. Problem

Powdered zinc reacts rapidly with powdered sulfur in a highly exothermic reaction.

$$8Zn_{(s)} + S_{8(s)} \rightarrow 8ZnS_{(s)}$$

What mass of zinc sulfide is expected when 32.0 g of S_8 reacts with sufficient zinc?

What Is Required?

You need to find the mass of ZnS that will be produced from the given amount of sulfur.

What Is Given?

The mass of the S_8 is given. The balanced equation is given.

Plan Your Strategy

Step 1 Convert the given mass (*m*) of sulfur to the number of moles (*n*) of sulfur, using its molar mass (*M*). Use the formula $n = m / M$

Step 2 Determine the mole ratio of sulfur to zinc sulfide from the balanced equation. Using the result in Step 1, solve for the number of moles of ZnS.

Step 3 Convert the number of moles of ZnS to mass using its molar mass. Use the formula $m = n \times M$ of the ZnS.

Act on Your Strategy

	$8Zn_{(s)}$	$+ S_{8(s)}$	$\rightarrow$	$8ZnS_{(s)}$
mole ratio	8	1		8
molar mass	65.39 g/mol	256.56 g/mol		97.46 g/mol
given		32.0 g		*m*

Step 1 $n = \dfrac{32.0 \cancel{\text{g}}\ S_8}{256.56 \cancel{\text{g}}/\text{mol}} = 0.1247 \text{ mol } S_8$

Step 2 $\dfrac{1 \text{ mol } S_8}{8 \text{ mol ZnS}} = \dfrac{0.1247 \text{ mol } S_8}{n \text{ mol ZnS}}$

$$n = 0.1247 \text{ mol } S_8 \times \frac{8 \text{ mol ZnS}}{1 \text{ mol } S_8} = 0.9976 \text{ mol ZnS}$$

Step 3 $m = 0.9976 \cancel{\text{mol}}\ \text{ZnS} \times 97.46 \text{ g}/\cancel{\text{mol}} = 97.2 \text{ g ZnS}$

Therefore, the mass of ZnS produced is 97.2 g.

Check Your Solution

The mol ratio of S_8:ZnS in this reaction is always 1:8, no matter what. Dividing both the 0.1247 mol of S_8 and 0.9976 mol of ZnS by 0.1247 will give the lowest ratio of 1:8. Your answer is consistent.

16. Problem

The addition of concentrated hydrochloric acid to manganese(IV) oxide leads to the production of chlorine gas.

$$4HCl_{(aq)} + MnO_{2(g)} \rightarrow MnCl_{2(aq)} + Cl_{2(g)} + 2H_2O_{(\ell)}$$

What mass of chlorine can be obtained when 4.76×10^{-2} g pf HCl react with sufficient MnO_2?

What Is Required?
You need to find the mass of Cl_2 that will be produced from the given amount of HCl.

What Is Given?
The mass of the HCl is given. The balanced equation is given.

Plan Your Strategy
Step 1 Convert the given mass (m) of HCl to the number of moles (n) of HCl, using its molar mass (M). Use the formula $n = m / M$
Step 2 Determine the mole ratio of HCl to Cl_2 from the balanced equation. Using the result in Step 1, solve for the number of moles of Cl_2.
Step 3 Convert the number of moles of Cl_2 to mass using its molar mass. Use the formula $m = n \times M$ of the Cl_2.

Act on Your Strategy

	$4HCl_{(aq)}$	$+ MnO_{2(g)}$	$\rightarrow$	$MnCl_{2(aq)}$	$+ Cl_{2(g)}$	$+ 2H_2O_{(\ell)}$
mole ratio	4	1		1	1	2
molar mass	36.46 g/mol	86.94 g/mol		125.84 g/mol	68.9 g/mol	36.04 g/mol
given	$(4.76)(10^{-2})$ g				m	

Step 1 $n = \dfrac{(4.76 \times 10^{-2})\ \cancel{g}\ HCl}{36.46\ \cancel{g}/mol} = 0.0013$ mol HCl

Step 2 $\dfrac{4\ mol\ HCl}{1\ mol\ Cl_2} = \dfrac{0.0013\ mol\ HCl}{n\ mol\ Cl_2}$

$n = 0.0013\ mol\ HCl \times \dfrac{1\ mol\ Cl_2}{4\ mol\ HCl} = 3.25 \times 10^{-4}\ mol\ Cl_2$

Step 3 $m = 3.25 \times 10^{-4}\ \cancel{mol}\ Cl_2 \times 68.9\ g/\cancel{mol} = (2.25 \times 10^{-2})\ g\ Cl_{2(g)}$

Check Your Solution
The mol ratio of HCl:Cl_2 in this reaction is always 4:1, no matter what. Dividing both the 0.0013 mol of HCl and 3.25×10^{-4} mol of Cl_2 by 3.25×10^{-4} will give the lowest ratio of 4:1. Your answer is consistent.

17. Problem
Aluminum carbide, Al_4C_3, is a yellow powder that reacts with water to produce aluminum hydroxide and methane.

$$Al_4C_{3(s)} + 12H_2O_{(\ell)} \rightarrow 4Al(OH)_{3(s)} + 3CH_{4(g)}$$

What mass of water is required to react completely with 25.0 g of aluminum carbide?

What Is Required?
You need to find the mass of H_2O that is needed to react with the given amount of Al_4C_3.

What Is Given?
The mass of the Al_4C_3 is given. The balanced equation is given.

Plan Your Strategy
Step 1 Convert the given mass (m) of Al_4C_3 to the number of moles (n) of Al_4C_3, using its molar mass (M). Use the formula $n = m / M$
Step 2 Determine the mole ratio of Al_4C_3 to H_2O from the balanced equation. Using the result in Step 1, solve for the number of moles of H_2O.
Step 3 Convert the number of moles of H_2O to mass using its molar mass. Use the formula $m = n \times M$ of the H_2O.

Act on Your Strategy

	$Al_4C_{3(s)}$	$+ 12H_2O_{(\ell)}$	$\rightarrow$	$4Al(OH)_{3(s)}$	$+ 3CH_{4(g)}$
mole ratio	1	12		4	3
molar mass	g/mol	g/mol		g/mol	g/mol
given	25.0 g	m			

Step 1 $n = \dfrac{25.0\ \cancel{g}\ Al_4C_3}{143.95\ \cancel{g}/mol} = 0.174\ mol\ Al_4C_3$

Step 2 $\dfrac{1\ mol\ Al_4C_3}{12\ mol\ H_2O} = \dfrac{0.174\ mol\ Al_4C_3}{n\ mol\ H_2O}$

$n = 0.174\ mol\ Al_4C_3 \times \dfrac{12\ mol\ H_2O}{1\ mol\ Al_4C_3} = 2.088\ mol\ H_2O$

Step 3 $m = 2.088\ \cancel{mol}\ H_2O \times 18.02\ g/\cancel{mol} = 37.63\ g$

Therefore, the mass of water that will react completely with the aluminum carbide is 37.63 g.

Check Your Solution

The mol ratio of Al_4C_3:H_2O in this reaction is always 1:12, no matter what. Dividing both the 0.174 mol of Al_4C_3 and 2.088 mol of H_2O by 0.174 will give the lowest ratio of 1:12. Your answer is consistent.

18. Problem

Magnesium oxide reacts with phosphoric acid, H_3PO_4, to produce magnesium phosphate and water.

$$3MgO_{(s)} + 2H_3PO_{4(aq)} \rightarrow Mg_3(PO_4)_{2(s)} + 3H_2O_{(\ell)}$$

How many grams of magnesium oxide are required to react completely with 33.5 g of phosphoric acid?

What Is Required?

You need to find the mass of MgO that is needed to react with the given amount of phosphoric acid.

What Is Given?

The mass of the H_3PO_4 is given. The balanced equation is given.

Plan Your Strategy

Step 1 Convert the given mass (m) of H_3PO_4 to the number of moles (n) of H_3PO_4, using its molar mass (M). Use the formula $n = m / M$

Step 2 Determine the mole ratio of H_3PO_4 to MgO from the balanced equation. Using the result in Step 1, solve for the number of moles of MgO.

Step 3 Convert the number of moles of MgO to mass using its molar mass. Use the formula $m = n \times M$ of the MgO.

Act on Your Strategy

	$3MgO_{(s)}$	$+ 2H_3PO_{4(aq)}$	$\rightarrow$	$Mg_3(PO_4)_{2(s)}$	$+ 3H_2O_{(\ell)}$
mole ratio	3	2		1	3
molar mass	40.3 g/mol	98.0 g/mol			
given	m	33.5 g			

Step 1 $n = \dfrac{33.5\ \cancel{g}\ H_3PO_4}{98.0\ \cancel{g}/mol} = 0.342\ mol\ H_3PO_4$

Step 2 $\dfrac{3\ mol\ MgO}{2\ mol\ H_3PO_4} = \dfrac{n\ mol\ MgO}{0.342\ mol\ H_3PO_4}$

$n = 0.342\ mol\ H_3PO_4 \times \dfrac{3\ mol\ MgO}{2\ mol\ H_3PO_4} = 0.513\ mol\ MgO$

Step 3 $m = 0.513\ \cancel{mol}\ MgO \times 40.3\ g/\cancel{mol} = 20.7\ g$

Therefore, the mass of MgO that will react completely with the phosphoric acid is 20.7 g.

Check Your Solution
The mol ratio of $MgO:H_3PO_4$ in this reaction is always 3:2, no matter what, which is 1.5:1 in the lowest ratio. Dividing both the 0.513 mol of MgO and 0.342 mol of H_3PO_4 by 0.342 will give the lowest ratio of 1.5:1. Your answer is consistent.

Solutions for Practice Problems

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19. Problem
Nitrogen gas is produced in an automobile air bag. It is generated by the decomposition of sodium azide, NaN_3.

$$2NaN_{3(s)} \rightarrow 3N_{2(g)} + 2Na_{(s)}$$

(a) To inflate the air bag on the driver's side of a certain car, 80.0 g of N_2 is required. What mass of NaN_3 is needed to produce 80.0 g of N_2?
(b) How many atoms of Na are produced when 80.0 g of N_2 are generated in this reaction?

What Is Required?
(a) You need to find the mass of NaN_3 needed to produce the required amount of nitrogen.
(b) You need to find the number of Na atoms produced in this reaction amount in (a).

What Is Given?
The mass of the nitrogen produced is given. The balanced equation is given.
Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy
(a) To obtain the mass of NaN_3, perform the following steps below:
Step 1 Convert the given mass (m) of N_2 to the number of moles (n) of N_2, using its molar mass (M). Use the formula $n = m \ / \ M$
Step 2 Determine the mole ratio of NaN_3 to N_2 from the balanced equation. Using the result in Step 1, solve for the number of moles of NaN_3.
Step 3 Convert the number of moles of NaN_3 to mass using its molar mass. Use the formula $m = n \times M$ of the NaN_3.
(b) Repeat Step 2 in (a) using the mole ratio of N_2 to Na, to determine the number of moles of Na that is produced. Multiply this by the Avogadro's number to obtain the number of atoms of Na.

Act on Your Strategy
(a)

	$2NaN_{3(s)}$	$\rightarrow$	$3N_{2(g)}$	$+ 2Na_{(s)}$
mole ratio	2		3	2
molar mass	65.02 g/mol		28.02 g/mol	22.99 g/mol
given	m		80.0 g	

Step 1 $n = \dfrac{80.0 \cancel{g} \ N_2}{28.02 \cancel{g}/mol} = 2.855 \text{ mol } N_2$

Step 2 $\dfrac{3 \text{ mol } N_2}{2 \text{ mol } NaN_3} = \dfrac{2.855 \text{ mol } N_2}{n \text{ mol } NaN_3}$

$n = 2.855 \text{ mol } N_2 \times \dfrac{2 \text{ mol } NaN_3}{3 \text{ mol } N_2} = 1.903 \text{ mol } NaN_3$

Step 3 $m = 1.903 \cancel{\text{mol}}\ NaN_3 \times 65.02\ \text{g}/\cancel{\text{mol}} = 124\ \text{g}\ NaN_3$
Therefore, the mass of NaN_3 needed is 124 g.

(b) $\dfrac{3\ \text{mol}\ N_2}{2\ \text{mol Na}} = \dfrac{2.855\ \text{mol}\ N_2}{n\ \text{mol Na}}$

$n = 2.855\ \text{mol}\ N_2 \times \dfrac{2\ \text{mol Na}}{3\ \text{mol}\ N_2} = 1.903\ \text{mol Na}$

$N = 1.903 \cancel{\text{mol}}\ \text{Na} \times (6.02 \times 10^{23})\ \text{atoms}/\cancel{\text{mol}} = 1.15 \times 10^{24}\ \text{atoms Na}$
Therefore, the number of atoms of Na produced is 1.15×10^{24} atoms.

Check Your Solution

(a) The mol ratio of NaN_3:N_2 in this reaction is always 2:3, no matter what, which is 1:1.5 in the lowest ratio. Dividing both the 1.903 mol of NaN_3 and 2.855 mol of N_2 by 1.903 will give the lowest ratio of 1:1.5. Your answer is consistent.

(b) Likewise, the mol ratio of N_2:Na in this reaction is always 3:2, no matter what, which is 1.5:1 in the lowest ratio. Dividing both the 2.855 mol of N_2 and 1.903 mol of Na by 1.903 will give the lowest ratio of 1.5:1. Your answer is consistent.

20. Problem

The reaction of iron(III) oxide with powered aluminum is known as the thermite reaction.

$$2Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(\ell)}$$

(a) Calculate the mass of aluminum oxide, Al_2O_3, that is produced when 1.42×10^{24} atoms of Al react with Fe_2O_3.

(b) How many molecules of Fe_2O_3 are needed to react with 0.134 g of Al?

What Is Required?

(a) You need to find the mass of Al_2O_3 produced from the given number of atoms of Al.

(b) You need to find the number of Fe_2O_3 molecules needed to react with the given amount of Al.

What Is Given?

(a) The number of atoms of Al is given. The balanced equation is given. Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

(b) The mass of the aluminum is given. The balanced equation is given. Avogadro's number, $N_A = 6.02 \times 10^{23}$ formula units/mol.

Plan Your Strategy

(a) Use the following steps below:

Step 1 Divide the number of Al atoms by Avogadro's number to obtain the number of moles in the reaction.

Step 2 Using the ratio of Al to Al_2O_3 from the balanced equation and the answer in Step 1, solve for the number of moles of Al_2O_3.

Step 3 Multiply the number of moles of Al_2O_3 by its molar mass to obtain the mass in grams.

(b) Use the following steps below:

Step 1 Convert the given mass (m) of Al to the number of moles (n) of Al, using its molar mass (M). Use the formula $n = m\ /\ M$

Step 2 Determine the mole ratio of Al to Fe_2O_3 from the balanced equation. Using the result in Step 1, solve for the number of moles of Fe_2O_3.

Step 3 Multiply the number of moles of Fe_2O_3 to the Avogadro number to obtain the number of molecules in the reaction.

Act on Your Strategy

(a)

	$2Al_{(s)}$	$+ Fe_2O_{3(s)}$	$\rightarrow$	$Al_2O_{3(s)}$	$+ 2Fe_{(\ell)}$
mole ratio	2	1		1	2
molar mass	26.98 g/mol	159.7 g/mol		101.96 g/mol	55.85 g/mol
given	1.42×10^{24} atoms			*m*	

Step 1 $n = \dfrac{(1.42 \times 10^{24})\ \cancel{\text{atoms}}\ \text{Al}}{(6.02 \times 10^{23})\ \cancel{\text{atoms}}/\text{mol}} = 2.36 \text{ mol Al}$

Step 2 $\dfrac{2 \text{ mol Al}}{1 \text{ mol } Al_2O_3} = \dfrac{2.36 \text{ mol Al}}{n \text{ mol } Al_2O_3}$

$n = 2.36 \text{ mol Al} \times \dfrac{1 \text{ mol } Al_2O_3}{2 \text{ mol Al}} = 1.18 \text{ mol } Al_2O_3$

Step 3 $m = 1.18\ \cancel{\text{mol}}\ Al_2O_3 \times 101.96 \text{ g}/\cancel{\text{mol}} = 120 \text{ g } Al_2O_3$

Therefore, the mass of aluminum oxide produced is 120 g.

(b)

	$2Al_{(s)}$	$+ Fe_2O_{3(s)}$	$\rightarrow$	$Al_2O_{3(s)}$	$+ 2Fe_{(\ell)}$
mole ratio	2	1		1	2
molar mass	26.98 g/mol	159.7 g/mol		93.96 g/mol	55.85 g/mol
given	0.134 g	*N*			

Step 1 $n = \dfrac{0.134\ \cancel{\text{g}}\ \text{Al}}{26.98\ \cancel{\text{g}}/\text{mol}} = 0.005 \text{ mol Al}$

Step 2 $\dfrac{2 \text{ mol Al}}{1 \text{ mol } Fe_2O_3} = \dfrac{0.005 \text{ mol Al}}{n \text{ mol } Fe_2O_3}$

$n = 0.005 \text{ mol Al} \times \dfrac{1 \text{ mol } Fe_2O_3}{2 \text{ mol Al}} = 0.0025 \text{ mol } Fe_2O_3$

Step 3 $N = 0.0025\ \cancel{\text{mol}}\ Fe_2O_3 \times (6.02 \times 10^{23}) \text{ molecules}/\cancel{\text{mol}}$

$= 1.5 \times 10^{21} \text{ molecules } Fe_2O_3$

Therefore, the number of Fe_2O_3 molecules needed is 1.5×10^{21} molecules.

Check Your Solution

(a) The mol ratio of $Al{:}Al_2O_3$ in this reaction is always 2:1, no matter what. Dividing both the 2.36 mol of Al and 1.18 mol of Al_2O_3 by 1.18 will give the lowest ratio of 2:1. Your answer is consistent.

(b) Likewise, the mol ratio of $Al{:}Fe_2O_3$ in this reaction is always 2:1, no matter what. Dividing both the 0.005 mol of Al and 0.0025 mol of Fe_2O_3 by 0.0025 will give the lowest ratio of 2:1. Your answer is consistent.

21. Problem

The thermal decomposition of ammonium dichromate is an impressive reaction. When heated with a Bunsen burner or propane torch, the orange crystals of ammonium dichromate slowly decompose to green chromium(III) oxide in a volcano-like display. Colourless nitrogen gas and water vapour are also given off.

$$(NH_4)_2Cr_2O_{7(s)} \rightarrow Cr_2O_{3(s)} + N_{2(g)} + 4H_2O_{(g)}$$

(a) Calculate the number of molecules of Cr_2O_3 that is produced from the decomposition of 10.0 g of $(NH_4)_2Cr_2O_7$.

(b) In a different reaction, 16.9 g of N_2 is produced when a sample of $(NH_4)_2Cr_2O_7$ is decomposed. How many water molecules are also produced in this reaction?

(c) How many molecules of $(NH_4)_2Cr_2O_7$ are needed to produce 1.45 g of H_2O?

What Is Required?

(a) You need to find the number of molecules of Cr_2O_3 in the reaction.

(b) You need to find the number of water molecules produced in the reaction.

(c) You need to find the number of molecules of ammonium dichromate reacted.

What Is Given?

(a) The mass of ammonium dichromate is 10.0 g.
(b) The mass of N_2 is 16.9 g.
(c) The mass of water is 1.45 g.
The balanced equation is given. Avogadro's number is 6.02×10^{23} molecules/mol.

Plan Your Strategy

(a) Apply the following steps:
Step 1 Convert the given mass (m) of ammonium dichromate to the number of moles (n), using its molar mass (M). Use the formula $n = m / M$
Step 2 Determine the mole ratio of $(NH_4)_2Cr_2O_7$ to Cr_2O_3 from the balanced equation. Using the result in Step 1, solve for the number of moles of Cr_2O_3.
Step 3 Multiply the number of moles of Cr_2O_3 to the Avogadro number to obtain the number of molecules in the reaction.
(b) Repeat the steps in (a), this time working with the mass of N_2 to obtain its number of moles, and then using the mole ratio of N_2 to H_2O to obtain the number of moles of water. Multiply this value by the Avogadro number to establish the number of molecules of water produced.
(c) Repeat the steps in (a), this time working with the mass of water given to obtain its number of moles, and then using the mole ratio of $(NH_4)_2Cr_2O_7$ to H_2O to obtain the number of moles of $(NH_4)_2Cr_2O_7$. Multiply this value by the Avogadro number to establish the number of molecules of $(NH_4)_2Cr_2O_7$ needed for the full reaction to proceed.

Act on Your Strategy

(a)

	$(NH_4)_2Cr_2O_{7(s)}$	$\rightarrow$	$Cr_2O_{3(s)}$	$+ N_{2(g)}$	$+ 4H_2O_{(g)}$
mole ratio	1		1	1	4
molar mass	252.1 g/mol		152 g/mol	28.02 g/mol	18.02 g/mol
given	10.0 g		N		

Step 1 $n = \dfrac{10.0\ \cancel{g}\ (NH_4)_2Cr_2O_7}{252.1\ \cancel{g}/mol} = 0.0396\ mol\ (NH_4)_2Cr_2O_7$

Step 2 $\dfrac{1\ mol\ (NH_4)_2Cr_2O_7}{1\ mol\ Cr_2O_3} = \dfrac{0.0396\ mol\ (NH_4)_2Cr_2O_7}{n\ mol\ Cr_2O_3}$

$n = 0.0396\ mol\ (NH_4)_2Cr_2O_7 \times \dfrac{1\ mol\ Cr_2O_3}{1\ mol\ (NH_4)_2Cr_2O_7}$

$n = 3.97 \times 10^{-2}\ mol\ Cr_2O_3$

Step 3 $N = 3.97 \times 10^{-2}\ \cancel{mol}\ Cr_2O_3 \times (6.02 \times 10^{23})\ molecules/\cancel{mol}$

$= 2.39 \times 10^{22}$ molecules Cr_2O_3

Therefore, the number of Cr_2O_3 molecules produced from the decomposition reaction is 2.39×10^{22}.

(b)

	$(NH_4)_2Cr_2O_{7(s)}$	$\rightarrow$	$Cr_2O_{3(s)}$	$+ N_{2(g)}$	$+ 4H_2O_{(g)}$
mole ratio	1		1	1	4
molar mass	252.1 g/mol		152 g/mol	28.02 g/mol	18.02 g/mol
given				16.9 g	N

Step 1 $n = \dfrac{16.9\ \cancel{g}\ N_2}{28.02\ \cancel{g}/mol} = 0.603\ mol\ N_2$

Step 2 $\frac{1 \text{ mol } N_2}{4 \text{ mol } H_2O} = \frac{0.603 \text{ mol } N_2}{n \text{ mol } H_2O}$

$n = 0.603 \text{ mol } N_2 \times \frac{4 \text{ mol } H_2O}{1 \text{ mol } N_2} = 2.41 \text{ mol } H_2O$

Step 3 $N = 2.41 \cancel{\text{mol}}\ H_2O \times (6.02 \times 10^{23}) \text{ molecules}/\cancel{\text{mol}}$

$= 1.45 \times 10^{24} \text{ molecules } H_2O$

Therefore, the number of water molecules produced is 1.45×10^{24}.

(c)

	$(NH_4)_2Cr_2O_{7(s)}$	$\rightarrow$	$Cr_2O_{3(s)}$	$+ N_{2(g)}$	$+ 4H_2O_{(g)}$
mole ratio	1		1	1	4
molar mass	252.1 g/mol		152 g/mol	28.02 g/mol	18.02 g/mol
given	*N*				1.45 g

Step 1 $n = \frac{1.45 \cancel{\text{g}}\ H_2O}{18.02 \cancel{\text{g}}/\text{mol}} = 0.0805 \text{ mol } H_2O$

Step 2 $\frac{1 \text{ mol } (NH_4)_2Cr_2O_7}{4 \text{ mol } H_2O} = \frac{n \text{ mol } (NH_4)_2Cr_2O_7}{0.0805 \text{ mol } H_2O}$

$n = 0.0805 \text{ mol } H_2O \times \frac{1 \text{ mol } (NH_4)_2Cr_2O_7}{4 \text{ mol } H_2O} = 0.020 \text{ mol } (NH_4)_2Cr_2O_7$

Step 3 $N = 0.020\ \cancel{\text{mol}}\ (NH_4)_2Cr_2O_7 \times (6.02 \times 10^{23}) \text{ molecules}/\cancel{\text{mol}}$

$= 1.21 \times 10^{22} \text{ molecules } (NH_4)_2Cr_2O_7$

Therefore, the number of $(NH_4)_2Cr_2O_7$ molecules needed in the reaction is 1.21×10^{22}.

Check Your Solution

(a) The mol ratio of $(NH_4)_2Cr_2O_7$ to Cr_2O_3 in this reaction is always 1:1, no matter what. Dividing both the 0.0396 mol of $(NH_4)_2Cr_2O_7$ and 3.97×10^{-2} mol of Cr_2O_3 by 0.0396 will give the lowest ratio of 1:1. Your answer is consistent.

(b) The mol ratio of N_2:H_2O in this reaction is always 1:4, no matter what. Dividing both the 0.020 mol of $(NH_4)_2Cr_2O_7$ and 0.0805 mol of H_2O by 0.020 will give the lowest ratio of 1:4. Your answer is consistent.

(c) Likewise, the mol ratio of $(NH_4)_2Cr_2O_7$:H_2O in this reaction is always 1:4, no matter what. Dividing both the 0.603 mol of N_2 and 2.41 mol of H_2O by 0.603 will give the lowest ratio of 1:4. Your answer is consistent.

22. Problem

Ammonia gas reacts with oxygen to produce water and nitrogen oxide. This reaction can be catalyzed, or sped up, by Cr_2O_3, produced in the reaction in problem 21.

$$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(\ell)}$$

(a) How many molecules of oxygen are is required to react with 34.0 g of ammonia?

(b) What mass of nitrogen oxide is expected from this reaction of 8.95×10^{24} molecules of oxygen with sufficient ammonia?

What Is Required?

(a) You need to find the number of molecules of oxygen required for the reaction.

(b) You need to find the mass of NO produced from the given number of oxygen molecules.

What Is Given?

(a) The mass of the ammonia is 34.0 g.

(b) The number of oxygen molecules is 8.95×10^{24}.

The balanced equation is given. Avogadro's number = 6.02×10^{23} molecules/mol.

Plan Your Strategy

(a) Apply the following steps:

Step 1 Convert the given mass (m) of ammonia to the number of moles (n), using its molar mass (M). Use the formula $n = m / M$
Step 2 Determine the mole ratio of NH_3 to O_2 from the balanced equation. Using the result in Step 1, solve for the number of moles of O_2.
Step 3 Multiply the number of moles of O_2 by the Avogadro number to obtain the number of molecules in the reaction.

(b) Apply the following steps:
Step 1 Divide the number of molecules of O_2 by the Avogadro number to obtain the number of moles in the reaction.
Step 2 Determine the mole ratio of O_2 to NO from the balanced equation. Using the result in Step 1, equate and solve for the number of moles of NO.
Step 3 Multiply the number of moles of NO by its molar mass, to obtain its mass in grams.

Act on Your Strategy

(a)

	$4NH_{3(g)}$	$+ 5O_{2(g)}$	$\rightarrow$	$4NO_{(g)}$	$6H_2O_{(\ell)}$
mole ratio	4	5		4	6
molar mass	17.04 g/mol	32.00 g/mol		30.01 g/mol	18.02 g/mol
given	34.0 g	N			

Step 1 $n = \dfrac{34.0\ \cancel{g}\ NH_3}{17.04\ \cancel{g}/mol} = 1.995\ mol\ NH_3$

Step 2 $\dfrac{4\ mol\ NH_3}{5\ mol\ O_2} = \dfrac{1.995\ mol\ NH_3}{n\ mol\ O_2}$

$n = 1.995\ mol\ NH_3 \times \dfrac{5\ mol\ O_2}{4\ mol\ NH_3} = 2.49\ mol\ O_2$

Step 3 $N = 2.49\ \cancel{mol}\ O_2 \times (6.02 \times 10^{23})\ molecules/\cancel{mol}$
$= 1.5 \times 10^{24}$ molecules O_2

Therefore, the number of oxygen molecules required for the reaction is 1.5×10^{24}.

(b)

	$4NH_{3(g)}$	$+ 5O_{2(g)}$	$\rightarrow$	$4NO_{(g)}$	$6H_2O_{(\ell)}$
mole ratio	4	5		4	6
molar mass	17.04 g/mol	32.00 g/mol		30.01 g/mol	18.02 g/mol
given		8.95×10^{24} molecules		m	

Step 1 $n = \dfrac{(8.95 \times 10^{24})\ \cancel{moleculses}\ O_2}{(6.02 \times 10^{23})\ \cancel{molecules}/mol} = 14.87\ mol\ O_2$

Step 2 $\dfrac{5\ mol\ O_2}{4\ mol\ NO} = \dfrac{14.87\ mol\ O_2}{n\ mol\ NO}$

$n = 14.87\ mol\ O_2 \times \dfrac{4\ mol\ NO}{5\ mol\ O_2} = 11.92\ mol\ NO$

Step 3 $m = 11.92\ \cancel{mol}\ NO \times 30.01\ g/\cancel{mol} = 357.7\ g\ NO$

Therefore, the mass of nitrogen oxide expected from the reaction is 357.7 g.

Check Your Solution

(a) The mol ratio of NH_3 to O_2 in this reaction is always 4:5, no matter what, which is 1:1.25 in its lowest ratio. Dividing both the 1.995 mol of NH_3 and 2.49 mol of O_2 by 1.995 will give the lowest ratio of 1:1.25. Your answer is consistent.

(b) The mol ratio of O_2:NO in this reaction is always 5:4, no matter what, which is 1.25:1 in its lowest ratio. Dividing both the 14.87 mol of O_2 and 11.92 mol of NO by 11.92 will give the lowest ratio of 1.25:1. Your answer is consistent.

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23. Problem

The following balanced chemical equation shows the reaction of aluminum with copper(II) chloride. If 0.25 g of aluminum reacts with 0.51 g of copper(II) chloride, determine the limiting reactant.

$$2Al_{(s)} + 3CuCl_{2(aq)} \rightarrow 3Cu_{(s)} + 2AlCl_{3(aq)}$$

What Is Required?

You need to find the limiting reagent in this reaction.

What Is Given?

Reactant Al = 0.25 g
Reactant $CuCl_2$ = 0.51 g
Products are Cu and $AlCl_3$
The equation is balanced.

Plan Your Strategy

Convert the given masses into moles (n). From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose $AlCl_3$. The reactant giving the smallest amount of $AlCl_3$ (in moles) is the limiting reactant.

Act on Your Strategy

	$2Al_{(s)}$	$+ 3CuCl_{2(aq)}$	$\rightarrow$	$3Cu_{(s)}$	$+ 2AlCl_{3(aq)}$
mole ratio	2	3		3	2
molar mass	26.98 g	134.5 g		63.55 g	133.33 g
given	0.25 g	0.51 g			n

Number of moles of Al $= \frac{0.25 \cancel{g} \text{ Al}}{26.98 \cancel{g}\text{/mol}} = 0.01$ mol

Number of moles of $CuCl_2 = \frac{0.51 \cancel{g} \text{ CuCl}_2}{134.5 \cancel{g}\text{/mol}} = 0.004$ mol

Therefore, the amount of $AlCl_3$ that should be produced based on:

Al $\rightarrow n = \frac{(0.01)(2)}{2} = 0.01$ mol of $AlCl_3$

$CuCl_2 \rightarrow n = \frac{(0.004)(2)}{3} = 0.003$ mol of $AlCl_3$

$CuCl_2$ will produce less $AlCl_3$ than Al, therefore, it is the limiting reagent.

Check Your Solution

The same result should be obtained using Cu as the basis for your calculations instead of $AlCl_3$. The mole ratio of Al:Cu is 2:3; the mol ratio of $CuCl_2$:Cu is 3:3.

For Al: n of Cu $= \frac{0.01 \text{ mol Al} \times 3 \text{ mol Cu}}{2 \text{ mol Al}} = 0.015$ mol

For $CuCl_3$: n of Cu $= \frac{0.004 \text{ mol CuCl}_3 \times 3 \text{ mol CuCl}_3}{3 \text{ mol Cu}} = 0.004$ mol

Again, the $CuCl_3$ produced less amount of the copper product, making it the limiting reagent. Your result is reasonable.

24. Hydrogen fluoride, HF, is a highly toxic gas. It is produced by the double displacement reaction of calcium fluoride, CaF_2, with concentrated sulfuric acid, H_2SO_4.

$$CaF_{2(s)} + H_2SO_{4(\ell)} \rightarrow 2HF_{(g)} + CaSO_{4(s)}$$

Determine the limiting reactant when 10.0 g of CaF_2 reacts with 15.5 g of H_2SO_4.

What Is Required?
You need to find the limiting reagent in this reaction.

What Is Given?
Reactant CaF_2 = 10.0 g
Reactant H_2SO_4 = 15.5 g
Products are HF and $CaSO_4$
The equation is balanced.

Plan Your Strategy
Convert the given masses into moles (*n*). From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose HF. The reactant giving the smallest amount of HF (in moles) is the limiting reactant.

Act on Your Strategy

	$CaF_{2(s)}$	$+ H_2SO_{4(aq)}$	$\rightarrow$	$2HF_{(g)} +$	$+ CaSO_{4(s)}$
mole ratio	1	1		2	1
molar mass	78.08 g	98.09 g		20.01 g	98.09 g
given	10.0 g	15.5 g		*n*	

Number of moles of $CaF_2 = \frac{10.0 \cancel{g}\ CaF_2}{78.08 \cancel{g}/mol} = 0.13$ mol

Number of moles of $H_2SO_4 = \frac{15.5 \cancel{g}\ H_2SO_4}{98.09 \cancel{g}/mol} = 0.16$ mol

Therefore, the amount of HF that should be produced based on:

$CaF_2 \rightarrow n = \frac{(0.13)(2)}{1} = 0.26$ mol

$H_2SO_4 \rightarrow n = \frac{(0.16)(2)}{1} = 0.32$ mol

CaF_2 will produce less HF than H_2SO_4, therefore, it is the limiting reagent.

Check Your Solution
The same result should be obtained using $CaSO_4$ as the basis for your calculations instead of HF. The mole ratio of CaF_2:$CaSO_4$ is 1:1; the mol ratio of H_2SO_4:$CaSO_4$ is 1:1.

For CaF_2: *n* of $CaSO_4 = \frac{0.13 \text{ mol } CaF_2 \times 1 \text{ mol } CaSO_4}{1 \text{ mol } CaF_2} = 0.13$ mol

For H_2SO_4: *n* of $CaSO_4 = \frac{0.16 \text{ mol } CaSO_4 \times 1 \text{ mol } CaSO_4}{1 \text{ mol } H_2SO_4} = 0.16$ mol

Again, the CaF_2 produced less amount of the calcium sulfate product, making it the limiting reagent. Your result is reasonable.

25. Acrylic, a common synthetic fibre, is formed when 10.0 g of acrylonitrile C_3H_3N. Acrylonitrile can be prepared by the reaction of propylene, C_3H_6, with nitric oxide, NO.

$$4C_3H_{6(g)} + 6NO_{(g)} \rightarrow 4C_3H_3N_{(g)} + 6H_2O_{(g)} + N_{2(g)}$$

What is the limiting reactant when 126 g of C_3H_6 reacts with 175 g of NO?

What Is Required?
You need to find the limiting reagent in this reaction.

What Is Given?
Reactant C_3H_6 = 126 g
Reactant NO = 175 g
Products are C_3H_3N, H_2O, and N_2

The equation is balanced.

Plan Your Strategy
Convert the given masses into moles (n). From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose N_2. The reactant giving the smallest amount of N_2 (in moles) is the limiting reactant.

Act on Your Strategy

	$4C_3H_{6(g)}$	$+ 6NO_{(g)}$	$\rightarrow$	$4C_3H_3N_{(g)}$	$+ 6H_2O_{(g)}$	$+ N_{2(g)}$
mole ratio	4	6		4	6	1
molar mass	42.09 g/mol	30.01 g/mol				28.02 g/mol
given	126 g	175 g				n

Number of moles of $C_3H_6 = \frac{126 \cancel{g}\ C_3H_6}{42.09 \cancel{g}/mol} = 2.99$ mol

Number of moles of $NO = \frac{175 \cancel{g}\ NO}{30.01 \cancel{g}/mol} = 5.83$ mol

Therefore, the amount of N_2 that should be produced based on:

$C_3H_6 \rightarrow\ n = \frac{(2.99)(1)}{4} = 0.75$ mol

$NO \rightarrow\ n = \frac{(5.83)(1)}{6} = 0.97$ mol

Therefore, C_3H_6 is the limiting reagent.

Check Your Solution
The same result should be obtained using either acrylonitrile or water as the basis for your calculations instead of N_2. For example, the mole ratio of C_3H_6:H_2O is 4:6; the mol ratio of NO:H_2O is 6:6.

For C_3H_6: n of $H_2O = \frac{2.99 \text{ mol } C_3H_6 \times 6 \text{ mol } H_2O}{4 \text{ mol } C_3H_6} = 4.48$ mol

For NO: n of $H_2O = \frac{5.83 \text{ mol NO} \times 6 \text{ mol } H_2O}{6 \text{ mol NO}} = 5.83$ mol

Again, the C_3H_6 produced less amount of the water product, making it the limiting reagent. Your result is reasonable.

26. Problem
3.76 g of zinc reacts with 8.93×10^{23} molecules of hydrogen chloride. Which reactant is present in excess?

What Is Required?
You need to find the reactant in excess in the reaction.

What Is Given?
Reactant Zn = 3.76 g
Reactant HCl = 8.93×10^{23} molecules.
You know Avogadro's number = 6.02×10^{23} molecules/mol.

Plan Your Strategy
Step 1 Write the balanced equation for this reaction.
Step 2 Determine the number of moles (n) of each reactant. For Zn, divide the given mass by its molar mass. For HCl, divide the number of molecules by the Avogadro number.
Step 3 From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose H_2.

The reactant giving the largest amount of H_2 (in moles) is the reactant in excess.

Act on Your Strategy

	Zn	+ 2HCl	→	$ZnCl_2$	+ H_2
mole ratio	1	3		1	1
molar mass	65.39 g/mol	36.46 g/mol		136.29 g/mol	2.02 g/mol
given	3.76 g	$(8.93)(10^{23})$ molecules			*n*

Step 1 The balanced equation is

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Step 2 Number of moles of Zn $= \dfrac{3.76\ \cancel{g}\ Zn}{65.39\ \cancel{g}/mol} = 0.06$ mol

Number of moles of HCl $= \dfrac{8.93 \times 10^{23}\ \cancel{molecule}s\ HCl}{6.02 \times 10^{23}\ \cancel{molecule}s/mol} = 1.48$ mol

Step 3 Therefore, the amount of H_2 that should be produced based on:

$Zn \rightarrow n = \dfrac{(0.06)(1)}{1} = 0.06$ mol

$HCl \rightarrow n = \dfrac{(0.74)(1)}{2} = 0.74$ mol

Therefore, HCl is present in excess.

Check Your Solution

The same result should be obtained using zinc chloride as the basis for your calculations instead of hydrogen. For example, the mole ratio of $Zn:ZnCl_2$ is 1:1; the mol ratio of $HCl:ZnCl_2$ is 2:1.

For Zn: n of $ZnCl_2 = \dfrac{0.06\ mol\ Zn \times 1\ mol\ ZnCl_2}{1\ mol\ Zn} = 0.06$ mol

For HCl: n of $ZnCl_2 = \dfrac{1.48\ mol\ HCl \times 1\ mol\ ZnCl_2}{2\ mol\ HCl} = 0.74$ mol

Again, the HCl produced more $ZnCl_2$ product, making it the reagent in excess. Your result is reasonable.

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27. Problem

Chloride dioxide, ClO_2, is a reactive oxidizing agent. It is used to purify water.

$$6ClO_{2(g)} + 3H_2O_{(\ell)} \rightarrow 5HClO_{3(aq)} + HCl_{(aq)}$$

(a) If 71.00 g of ClO_2 is mixed with 19.00 g of water, what is the limiting reactant?
(b) What mass of $HClO_3$ expected in part (a)?
(c) How many molecules of HCl are expected in part (a)?

What Is Required?

(a) You need to find the limiting reactant in the reaction.
(b) You need to find the mass of the $HClO_3$ produced in the reaction.
(c) You need to find the number of molecules of HCl produced in the reaction.

What Is Given?

Reactant ClO_2 = 71.00 g
Reactant H_2O = 19.00 g
Products are $HClO_3$ and HCl
The balanced equation is given.
You know Avogadro's number = 36.02×10^{23} molecules/mol.

Plan Your Strategy

(a) Convert the given masses into moles (n). From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose $HClO_3$. The reactant giving the smallest amount of $HClO_3$ (in moles) is the limiting reactant.

(b) Multiply the mole amount of $HClO_3$, obtained from the limiting reactant in (a), by the molar mass of $HClO_3$ to obtain its mass in grams.

(c) Repeat step (a) to solve for the number of moles of HCl produced by the limiting reactant. Multiply this result by the Avogadro number to obtain the number of molecules of HCl produced.

Act on Your Strategy

	$6ClO_{2(g)}$	$+ 3H_2O_{(\ell)}$	$\rightarrow$	$5HClO_{3(aq)}$	$+ HCl_{(aq)}$
mole ratio	6	3		5	1
molar mass	67.54 g/mol	18.02 g/mol		84.46 g/mol	36.46 g/mol
given	71.00 g	19.00 g			

(a) Number of moles of $ClO_2 = \frac{71.00 \cancel{g}\ ClO_2}{67.54 \cancel{g}/mol} = 1.0526$ mol

Number of moles of $H_2O = \frac{19.00 \cancel{g}\ H_2O}{18.02 \cancel{g}/mol} = 1.0544$ mol

Therefore, the amount of $HClO_3$ that should be produced based on:

$ClO_2 \rightarrow n = \frac{1.0526 \times 5}{6} = 0.877$ mol

$H_2O \rightarrow n = \frac{1.0544 \times 5}{3} = 1.757$ mol

Therefore, ClO_2 is the limiting reagent.

(b) Number of moles of $HClO_3$ produced = 0.877 mol (from (a))

Therefore, $m = 0.877\ \cancel{mol}\ HClO_3 \times 84.46\ g/\cancel{mol} = 74.07\ g\ HClO_3$

(c) Using the number of moles of ClO_2 (the limiting reagent):

Number of moles of HCl produced $= \frac{1.0526\ mol \times 1}{6} = 0.146$ mol

Therefore, $N = 0.146\ \cancel{mol}\ HCl \times (6.02 \times 10^{23})\ molecules/\cancel{mol}$
$= 8.8 \times 10^{22}$ molecules HCl

Check Your Solution

The mole ratio of the reactants ClO_2:H_2O was 6:3 which is 2:1 in its lowest ratio. The calculated mole ratio from the given masses of reactants was 0.877 mol ClO_2:1.757 mol H_2O, which is 1:2 in its lowest ratio (dividing both by 0.877). Clearly, there was not enough ClO_2 available to satisfy the ratio needed by the balanced equation, therefore, it is the limiting reagent in this case. The results are reasonable.

28. Problem

Hydrazine, N_2H_4, reacts exothermically with hydrogen peroxide, H_2O_2.

$$N_2H_{4(\ell)} + 7H_2O_{2(aq)} \rightarrow 2HNO_{3(g)} + 8H_2O_{(g)}$$

(a) 120 g of N_2H_4 reacts with an equal mass of H_2O_2. Which is the limiting reactant?

(b) What mass of HNO_3 is expected?

(c) What mass, in grams, of the excess reactant remains at the end of the reaction?

What Is Required?

(a) You need to find the limiting reactant in the reaction.

(b) You need to find the mass of the HNO_3 produced in the reaction.

(c) You need to find the mass of excess reagent left over after the reaction.

What Is Given?
Reactant $N_2H_4 = 120$ g
Reactant $H_2O_2 = 120$ g
Products are HNO_3 and H_2O
The balanced equation is given.

Plan Your Strategy

(a) Convert the given masses into moles (n). From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose HNO_3. The reactant giving the smallest amount of HNO_3 (in moles) is the limiting reactant.

(b) Multiply the mole amount of HNO_3, obtained from the limiting reactant in (a), by the molar mass of HNO_3 to obtain its mass in grams.

(c) Using the mole ratio of the reactants N_2H_4 and H_2O_2 from the balanced equation and the number of moles of the limiting reagent calculated in (a), equate and solve for the number of moles of excess reagent used in this reaction. Subtract this mole amount from the given mole amount calculated in (a) for the identified excess reagent. Multiply the mole difference by the molar mass of the excess reagent to obtain the mass left over.

Act on Your Strategy

	$N_2H_{4(\ell)}$	$+ 7H_2O_{2(aq)}$	$\rightarrow$	$2HNO_{3(g)}$	$+ 6H_2O_{(g)}$
mole ratio	1	7		2	6
molar mass	32.06	34.02		63.02	18.02
given	120 g	120 g		n	

(a) Number of moles of $N_2H_4 = \frac{120 \cancel{g}\ N_2H_4}{32.06 \cancel{g}/mol} = 3.74$ mol

Number of moles of $H_2O_2 = \frac{120 \cancel{g}\ H_2O_2}{34.02 \cancel{g}/mol} = 3.52$ mol

Therefore, amount of HNO_3 that should be produced based on:

$N_2H_4 = \frac{3.74 \times 2}{1} = 7.48$ mol

$H_2O_2 = \frac{3.52 \times 2}{1} = 1.01$ mol

H_2O_2 is the limiting reagent.

(b) Number of moles of HNO_3 produced = 1.01 mol (from (a))
Therefore, $m = 1.01 \cancel{mol}\ HNO_3 \times 63.02\ g/\cancel{mol} = 63.65$ g HNO_3

(c) Number of moles of N_2H_2 used $= \frac{3.52 \text{ mol} \times 1}{7} = 0.503$ mol
Therefore, number of moles of N_2H_2 in excess=3.74 mol−0.503 mol=3.24 mol
Hence, $m = 3.24$ mol $N_2H_4 \times 32.06$ g/mol = 103.87 g
The mass of N_2H_4 left over after the reaction is 103.87 g.

Check Your Solution

The mole ratio of the reactants N_2H_4 to H_2O_2 from the balanced equation is 1:7. The calculated mole ratio from the given masses of reactants was 3.74 mol C N_2H_4:3.54 mol H_2O_2, which is 1.06:1 in its lowest ratio (dividing both by 3.54). Clearly, there was not enough H_2O_2 available to satisfy the 1:7 ratio needed by the balanced equation, therefore, it is the limiting reagent in this case and the N_2H_4 was in excess. The results are reasonable.

29. Problem

In the textile industry, chlorine is used to bleach fabrics. Any of the toxic chlorine that remains after the bleaching process is destroyed by reacting it with a sodium thiosulfate solution, $Na_2S_2O_{3(aq)}$.

$$Na_2S_2O_{3(aq)} + 4Cl_{2(g)} + 5H_2O_{(\ell)} \rightarrow 2NaHSO_{4(aq)} + 8HCl_{(aq)}$$

135 kg of $Na_2S_2O_3$ reacts with 50.0 kg of Cl_2 and 238 kg of water. How many grams of $NaHSO_4$ are expected?

What Is Required?

You need to find the mass of the $NaHSO_4$ produced in this reaction.

What Is Given?

Reactant $Na_2S_2O_3$ = 135 kg
Reactant Cl_2 = 50.0 kg
Reactant H_2O = 238 kg
Products are $NaHSO_4$ and HCl
The balanced equation is given.

Plan Your Strategy

Follow the steps below:

Step 1 Determine the identity of the limiting reagent for the given amounts of reactants. Convert the given masses into moles (n). From the balanced equation, use the mole ratio of the reactants and products to determine how much either of the products is produced from the calculated number of moles of each reactant. In this case, we will choose $NaHSO_4$. The reactant giving the smallest amount of $NaHSO_4$ (in moles) is the limiting reactant. Remember to convert the given kg quantities to grams before proceeding with the calculations.

Step 2 From Step 1, the number of moles of $NaHSO_4$ produced by the limiting reagent is the amount that will be produced by this reaction. Multiply this mole amount by the molar mass of $NaHSO_4$ to obtain its mass in grams.

Act on Your Strategy

	$Na_2S_2O_{3(aq)}$	$+ 4Cl_{2(g)}$	$+ 5H_2O_{(\ell)}$	$\rightarrow$	$2NaHSO_4$	$+ 8HCl_{(aq)}$
Mole ratio	1	4	5		2	8
given	158.12 g/mol	70.9 g/mol	18.02 g/mol		120.07 g/mol	36.46 g/mol
given	135 kg	50.0 kg	238 kg		n	

Step 1 Number of moles of $Na_2S_2O_3 = \dfrac{(1.35 \times 10^5)\ \cancel{g}\ Na_2S_2O_3}{158.12\ \cancel{g}/mol} = 853.778$ mol

Number of moles of $Cl_2 = \dfrac{(5.02 \times 10^4)\ \cancel{g}\ Cl_2}{70.9\ \cancel{g}/mol} = 705.22$ mol

Number of moles of $H_2O = \dfrac{(2.38 \times 10^5)\ \cancel{g}\ H_2O}{18.02\ \cancel{g}/mol} = 13207.55$ mol

Therefore, the mole amount of $NaHSO_4$ that should be produced based on:

$Na_2S_2O_3 \rightarrow n = \dfrac{853.778\ mol \times 2}{1} = 1707.6$ mol

$Cl_2 \rightarrow n = \dfrac{705.22\ mol \times 2}{4} = 352.6$ mol

$H_2O \rightarrow n = \dfrac{13207.55\ mol \times 2}{5} = 5282.8$ mol

The limiting reagent is Cl_2.

Step 2 From Step 1, 352.6 mol of $NaHSO_4$ will be produced.

Therefore, $m = 352.6\ \cancel{mol}\ NaHSO_4 \times 120.07\ g/\cancel{mol} = 4.2 \times 10^4$ g.

The mass of $NaHSO_4$ produced is 4.2×10^4 g.

Check Your Solution
The mole ratio of the reactants $Na_2S_2O_3$ to Cl_2 to H_2O from the balanced equation is 1:4:5. The calculated mole ratio from the given masses of reactants was 853.778:705.22: 13207.55, which is 1.2:1:18.7 in its lowest ratio (dividing all three by 705.22). Clearly, there is not enough Cl_2 available to satisfy the 1:4 ratio needed between it and $Na_2S_2O_3$, therefore, Cl_2 is the limiting reagent in this case. The results are reasonable.

30. Problem
Manganese(III) fluoride can be formed by the reaction of manganese(II) iodide with fluorine.

$$2MnI_{2(s)} + 13F_{2(g)} \rightarrow 2MnF_{3(s)} + 4IF_{5(\ell)}$$

(a) 1.23 g of MnI_2 reacts with 25.0 g of F_2. What mass of MnF_3 is expected?
(b) How many molecules of IF_5 are produced in part (a)?
(c) What reactant is in excess? How much of it remains at the end of the reaction?

What Is Required?
(a) You need to find the mass of the MnF_3 produced in this reaction.
(b) You need to find the number of IF_5 molecules produced in this reaction.
(c) You have to find the mass of excess reactant left over after the reaction.

What Is Given?
Reactant MnI_2 = 1.23 g
Reactant F_2 = 25.0 g
Products are MnF_3 and IF_5
The balanced equation is given.
You know Avogadro's number = 6.02×10^{23} molecules/mol.

Plan Your Strategy
(a) Determine the identity of the limiting reagent for the given amounts of reactants. Convert the given masses into moles (n). From the balanced equation, use the mole ratio of the reactants and products to determine how much MnF_3 is produced from the calculated number of moles of each reactant. The reactant giving the smallest amount of MnF_3 (in moles) is the limiting reactant. Take this smaller mole amount of MnF_3 and multiply it by its molar mass to obtain the mass in grams.
(b) Repeat (a) to establish the mole ratio of the limiting reagent to IF_5 from the balanced equation, and equate and solve for the number of moles of IF_5. Multiply this mole amount by the Avogadro number to establish the number of molecules of IF_5 produced.
(c) From (a), the excess reactant can be identified. Using the mole ratio of MnI_2 and F_2 from the balanced equation and the number of moles of limiting reagent used in the reaction, equate and solve for the mole amount of the excess reagent. Subtract this value from the given mole amount at the start of the reaction to obtain the mole amount in left over.

Act on Your Strategy

	$2MnI_{2(s)}$	$+ 13F_{2(g)}$	$\rightarrow$	$2MnF_{3(g)}$	$+ 4IF_{5(\ell)}$
mole ratio	2	13		2	4
molar mass	308.74 g/mol	38.00 g/mol		111.94 g/mol	221.90 g/mol
given	1.23 g	25.0 g		m	

(a) Number of moles of $MnI_2 = \dfrac{1.23 \cancel{g}\ MnI_2}{308.74 \cancel{g}/mol} = 0.004$ mol

Number of moles of $F_2 = \frac{25.0 \cancel{g} F_2}{38.00 \cancel{g}/mol} = 0.658$ mol

Therefore, the number of moles of MnF_3 that should be produced based on:

$MnI_2 \rightarrow n = \frac{0.004 \text{ mol} \times 2}{2} = 0.004$ mol

$F_2 \rightarrow n = \frac{0.658 \text{ mol} \times 2}{13} = 0.101$ mol

The limiting reagent is MnI_2.
Using the mole amount produced by the limiting reagent:
$m = 0.004 \cancel{\text{mol}}\ MnF_3 \times 111.94 \text{ g}/\cancel{\text{mol}} = 0.446 \text{ g } MnF_3$
Therefore, 0.446 g of MnF_3 will be produced.

(b) Number of moles of IF_5 produced $= \frac{0.004 \text{ mol} \times 4}{2} = 0.008$ mol

$N = 0.008 \text{ mol } IF_5 \times (6.02 \times 10^{23})$ molecules/mol $= 4.8 \times 10^{21}$ molecules
Therefore, 4.8×10^{21} molecules of IF_5 will be produced.

(c) From (a), the reactant in excess is F_2.

Number of moles of F_2 used $= \frac{0.004 \text{ mol} \times 13}{2} = 0.026$ mol

Therefore, number of moles of F_2 in excess=0.658 mol−0.026 mol=0.632 mol.

Check Your Solution

The mole ratio of the reactants MnI_2 to F_2 from the balanced equation is 2:13, which is 1:6.5 in its lowest ratio. The calculated mole ratio from the given masses of reactants was 0.004 mol MnI_2:0.65 mol F_2, which is 1:162 in its lowest ratio (dividing both by 0.004). Clearly, there is not enough MnI_2 available to react with the vast amount of F_2 present at the start of the reaction. Therefore, MnI_2 is the limiting reagent in this case. The results are reasonable.

Solutions for Practice Problems

Student Textbook pages 262, 264

31. Problem

20.0 g of bromic acid, $HBrO_3$, is reacted with excess HBr.

$$HBrO_{3(aq)} + 5HBr_{(aq)} \rightarrow 3H_2O_{(\ell)} + 3Br_{2(aq)}$$

(a) What is the theoretical yield of Br_2 for this reaction?
(b) If 47.3 g of Br_2 is produced, what is the percentage yield of Br_2?

What Is Required?

(a) You need to find the theoretical yield of Br in the reaction.
(b) You need to find the actual percentage yield of Br from the mass produced.

What Is Given?

(a) Mass of reactant bromic acid = 20.0 g.
(b) Actual yield of Br_2 = 47.3 g.
The balanced equation is given.

Plan Your Strategy

(a) Divide the mass of bromic acid by its molar mass to obtain the mole amount. Use this value and the mole ratio of $HBrO_3$ to Br_2 from the balanced equation to equate and solve for the number of moles of Br_2 produced. Multiply this value by the molar mass of Br_2 to obtain the theoretical yield in grams.
(b) Divide the given actual yield by the theoretical yield calculated in (a) and multiply by 100%.

Act on Your Strategy

	$HBrO_{3(aq)}$	$+ 5HBr_{(aq)}$	$\rightarrow$	$3H_2O_{(l)}$	$+ 3Br_{2(aq)}$
mole ratio	1	5		3	3
molar mass	128.91 g/mol	80.91 g/mol		18.02 g/mol	159.8 g/mol
given	20.0 g				*m*

(a) Number of moles of $HBrO_3 = \frac{20.0\ \cancel{g}\ HBrO_3}{128.91\ \cancel{g}/mol} = 0.155$ mol

Therefore, the number of moles of Br_2 produced $= \frac{0.155\ mol \times 3}{1} = 0.465$ mol

Hence, $m = 0.465$ mol $Br_2 \times 159.8$ g/mol $= 74.38$ g

The theoretical yield of Br_2 is 74.38 g.

(b) Percentage yield of $Br_2 = \frac{47.3\ \cancel{g}}{74.38\ \cancel{g}} \times 100\% = 63.6\%$

Check Your Solution

Use whole numbers for a quick inspection: $\frac{47}{74} = 0.635$. This decimal result is consistent with the percentage result.

32. Problem

Barium sulfate forms as a precipitate in the following reaction:

$$Ba(NO_3)_{2(aq)} + NaSO_{4(aq)} \rightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$$

When 35.0 g of $Ba(NO_3)_2$ is reacted with excess $NaSO_4$, 29.8 g of $BaSO_4$ is recovered by the chemist.

(a) Calculate the theoretical yield of $BaSO_4$.

(b) Calculate the percentage yield of $BaSO_4$.

What Is Required?

(a) You need to find the theoretical yield of $BaSO_4$ in the reaction.

(b) You need to find the actual percentage yield of $BaSO_4$ from the mass recovered.

What Is Given?

(a) Mass of reactant $Ba(NO_3)_2 = 35.0$ g.

(b) Actual yield of $BaSO_4 = 29.8$ g.

The balanced equation is given.

Plan Your Strategy

(a) Divide the mass of $Ba(NO_3)_2$ by its molar mass to obtain the mole amount. Use this value and the mole ratio of $Ba(NO_3)_2$ to $BaSO_4$ from the balanced equation to equate and solve for the number of moles of $BaSO_4$ produced. Multiply this value by the molar mass of $BaSO_4$ to obtain the theoretical yield in grams.

(b) Divide the given actual yield by the theoretical yield calculated in (a) and multiply by 100%.

Act on Your Strategy

	$Ba(NO_3)_{2(aq)}$	$+ Na_2SO_{4(aq)}$	$\rightarrow$	$BaSO_{4(s)}$	$+ 2NaNO_{3(aq)}$
mole ratio	1	1		1	2
molar mass	261.35 g/mol	142.05 g/mol		233.4 g/mol	85 g/mol
given	35.0 g			*m*	

(a) Number of moles of $Ba(NO_3)_2 = \frac{35.0\ \cancel{g}\ Ba(NO_3)_2}{261.35\ \cancel{g}/mol} = 0.134$ mol

Therefore, the number of moles of $BaSO_4$ produced $= \frac{0.134\ mol \times 1}{1} = 0.134$ mol

Hence, $m = 0.134\ \cancel{mol}\ BaSO_4 \times 233.4\ g/\cancel{mol} = 31.28$ g

The theoretical yield of $BaSO_4$ is 31.28 g.

(b) Percentage yield of $BaSO_4 = \frac{29.8 \cancel{g}}{31.28 \cancel{g}} \times 100\% = 95.3\%$

Check Your Solution
Use whole numbers for a quick inspection: $\frac{30}{31} = 0.968$. This decimal result is consistent with the percentage result.

33. Problem
Yeasts can act on a sugar, such as glucose, $C_6H_{12}O_6$, to produce ethyl alcohol, C_2H_5OH, and carbon dioxide.

$$C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$$

If 223 g of ethyl alcohol are recovered after 1.63 kg of glucose react, what is the percentage yield of the reaction?

What Is Required?
You need to find the percentage yield of this reaction.

What Is Given?
Reactant glucose = 1.63 kg. Actual yield of ethyl alcohol = 223 g. The balanced equation is given.

Plan Your Strategy
First find the theoretical yield: Divide the mass of glucose by its molar mass to obtain the mole amount. Use this value and the mole ratio of glucose to ethyl alcohol from the balanced equation to equate and solve for the number of moles of ethyl alcohol produced. Multiply this value by the molar mass of ethyl alcohol to obtain the theoretical yield in grams. Finally, divide the given actual yield by the theoretical yield and multiply by 100%.

Act on Your Strategy

	$C_6H_{12}O_6$	$\rightarrow$	$2C_2H_5OH$	$+ 2CO_2$
mole ratio	1		2	2
molar mass	180.18 g/mol		46.08 g/mol	44.01 g/mol
given	1.63 kg		m	

Number of moles of $C_6H_{12}O_6 = \frac{(1.63 \times 10^3)\ \cancel{g}\ C_6H_{12}O_6}{180.18\ \cancel{g}/mol} = 9.045$ mol

Number of moles of C_2H_5OH produced $= \frac{9.05 \text{ mol} \times 2}{1} = 18.09$ mol

Therefore, $m = 18.09\ \cancel{mol}\ C_2H_5OH \times 46.08\ g/\cancel{mol} = 833.7\ g\ C_2H_5OH$ produced

Finally, the percentage yield of $C_2H_5OH = \frac{233 \cancel{g}}{833.7 \cancel{g}} \times 100\% = 26.7\%$

Check Your Solution
Use whole numbers for a quick inspection: $\frac{200}{800} = 0.25$. This decimal result is consistent with the percentage result.

34. Problem
The following reaction proceeds with a 70% yield.

$$C_6H_{6(\ell)} + HNO_{3(aq)} \rightarrow C_6H_5NO_{2(\ell)} + H_2O_{(\ell)}$$

Calculate the mass of $C_6H_5NO_2$ expected if 12.8 g of C_6H_6 expected if 12.8 g of C_6H_6 reacts with excess HNO_3.

What Is Required?
You have to find the actual yield of $C_6H_5NO_2$ in this reaction.

What Is Given?
Reactant C_6H_6 = 12.8 g. Percentage yield of $C_6H_5NO_2$ is 70%. The balanced equation is given.

Plan Your Strategy

First find the theoretical yield: Divide the mass of C_6H_6 by its molar mass to obtain the mole amount. Use this value and the mole ratio of C_6H_6 to $C_6H_5NO_2$ from the balanced equation to equate and solve for the number of moles of $C_6H_5NO_2$ produced. Multiply this value by the molar mass of $C_6H_5NO_2$ to obtain the theoretical yield in grams. Finally, multiply the theoretical yield by the percentage yield (converted to a decimal) to obtain the actual yield in grams.

Act on Your Strategy

	$C_6H_{6(l)}$	$+ HNO_{3(aq)}$	$\rightarrow$	$C_6H_5NO_{2(l)}$	$+ H_2O_{(l)}$
mole ratio	1	1		1	1
molar mass	78.12 g/mol	63.02 g/mol		123.12 g/mol	18.02 g/mol
given	12.8 g			*m*	

Number of moles of $C_6H_6 = \frac{128 \cancel{g}}{78.12 \cancel{g}/mol} = 0.164$ mol

Number of moles of $C_6H_5NO_2$ produced $= \frac{0.164 \text{ mol} \times 1}{1} = 0.164$ mol

Therefore, $m = 0.164 \cancel{mol}\ C_6H_5NO_2 \times 123.12\ g/\cancel{mol} = 20.19\ g\ C_6H_5NO_2$

Finally, the actual yield $= 20.19\ g\ C_6H_5NO_2 \times 0.70 = 14.1$ g

Check Your Solution

Use whole numbers for a quick inspection: $20 \times 0.70 = 14.0$. The results match.

35. Problem

The reaction of toluene, C_7H_8, with potassium permanganate, $KMnO_4$, gives less than a 100% yield.

$$C_7H_{8(\ell)} + 2KMnO_{4(aq)} \rightarrow KC_7H_5O_{2(aq)} + 2MNO_{2(s)} + KOH_{(aq)} + H_2O_{(\ell)}$$

(a) 8.60 g of C_7H_8 is reacted with excess $2KMnO_4$. What is the theoretical yield, in grams, of $KC_7H_5O_2$?

(b) If the percentage yield is 70.0%, what mass of $KC_7H_5O_2$ can be expected?

(c) What mass of C_7H_8 is needed to produce 13.4 g of $KC_7H_5O_2$, assuming a yield of 60%?

What Is Required?

(a) You need to find the theoretical yield of $KC_7H_5O_2$.

(b) You need to find the mass of $KC_7H_5O_2$ assuming the given percentage yield.

(c) You need to find the mass of C_7H_8 that yields the give percentage yield of product.

What Is Given?

(a) Reactant $C_7H_8 = 8.6$ g

(b) Percentage yield of product $KC_7H_5O_2 = 70\%$

(c) Product $KC_7H_5O_2 = 13.4$ g at a percentage yield of 60%

The balanced equation is given.

Plan Your Strategy

(a) Divide the mass of C_7H_8 by its molar mass to obtain the mole amount. Use this value and the mole ratio of C_7H_8 to $KC_7H_5O_2$ from the balanced equation to equate and solve for the number of moles of $KC_7H_5O_2$ produced. Multiply this value by the molar mass of $KC_7H_5O_2$ to obtain the theoretical yield in grams.

(b) Multiply the theoretical yield in (a) by the percentage yield (converted to a decimal) to obtain the actual yield in grams.

(c) Work backward with the given values. Divide the given theoretical yield by the percentage yield (converted to a decimal) to obtain the actual yield in grams. Dive this mass by the molar mass of $KC_7H_5O_2$ to obtain its number of moles. Use this value and the mole ratio of C_7H_8 to $KC_7H_5O_2$ from the balanced equation to equate and solve for the number of moles of C_7H_8. Multiply this value by the molar mass of C_7H_8 to obtain the mass.

Act on Your Strategy

	$C_7H_{8(l)}$	$+ 2KMnO_{4(aq)}$	$\rightarrow$	$KC_7H_5O_{2(aq)}$	$+ 2MnO_{2(s)}$	$+ KOH_{(aq)}$	$+ H_2O_{(l)}$
mole ratio	1	2		1	2	1	1
molar mass	92.15 g/mol	158.04 g/mol		160.22 g/mol	86.94 g/mol	56.11 g/mol	18.02 g/mol
given	8.60 g			*m*			

(a) Number of moles of $C_7H_8 = \frac{8.60 \cancel{g}\ C_7H_8}{92.15 \cancel{g}/mol} = 0.093$ mol

Number of moles of $KC_7H_5O_2$ produced $= \frac{0.093 \text{ mol} \times 1}{1} = 0.093$ mol

Therefore, $m = 0.093 \cancel{mol}\ KC_7H_5O_2 \times 160.22\ g/\cancel{mol} = 14.9$ g

The theoretical yield of $KC_7H_5O_2$ is 14.9 g.

(b) Expected yield = actual yield = 14.9 g $KC_7H_5O_2$ × 0.70 = 10.43 g $KC_7H_5O_2$

(c) Actual yield of $KC_7H_5O_2 = \frac{13.4 \text{ g } KC_7H_5O_2}{0.6} = 22.3$ g

Number of moles of $KC_7H_5O_2 = \frac{22.3 \cancel{g}\ KC_7H_5O_2}{160.22 \cancel{g}/mol} = 0.139$ mol

Number of moles C_7H_8 required $= \frac{0.139 \text{ mol} \times 1}{1} = 0.139$ mol

$m = 0.139 \cancel{mol}\ C_7H_8 \times 92.15\ g/\cancel{mol} = 12.8$ g

The mass of C_7H_8 required is 12.8 g.

Check Your Solution

(a) and **(b)**. The number of moles of actual $KC_7H_5O_2 = \frac{10.43 \text{ g}}{160.22 \text{ g/mol}} = 0.065$ mol.

The theoretical number of moles calculated = 0.093 mol.

Actual mole / theoretical mol $= \frac{0.065}{0.093} = 0.698 = 69.8\%$. This is consistent with the 70% percentage yield given.

(c) Likewise, the theoretical number of moles of 13.4 g $KC_7H_5O_2 = \frac{13.4 \cancel{g}}{160.22 \cancel{g}/mol} = 0.0836$ mol. The actual number of moles calculated = 0.139 mol.

Actual mole / theoretical mol $= \frac{0.0836}{0.139} = 0.601 = 60.1\%$. This is consistent with the 60% percentage yield given.

36. Problem

Marble is made primarily of calcium carbonate. When calcium carbonate reacts with hydrogen chloride, it reacts to form calcium chloride, carbon dioxide, and water. If this reaction occurs with 81.5% yield, what mass of carbon dioxide will be collected if 15.7 g of $CaCO_3$ is added to sufficient hydrogen chloride?

What Is Required?

You need to find the actual yield of CO_2 from the reaction.

What Is Given?

Reactant $CaCO_3$ = 15.7 g. Percentage yield of CO_2 = 81.5%. The reactants and products of the reaction are given.

Plan Your Strategy

Step 1 Write the full balanced equation for this reaction.

Step 2 Determine the theoretical yield of CO_2. Divide the mass of $CaCO_3$ by its molar mass to obtain the mole amount. Use this value and the mole ratio of $CaCO_3$ to CO_2 from the balanced equation to equate and solve for the number of moles of CO_2 produced. Multiply this value by the molar mass of CO_2 to obtain the theoretical yield in grams.

Step 3 Multiply the theoretical yield by the percentage yield (converted to a decimal) to obtain the actual yield in grams.

Act on Your Strategy

	$CaCO_{3(s)}$	$+ 2HCl_{(aq)}$	$\rightarrow$	$CaCl_{2(s)}$	$+ CO_{2(g)}$	$+ H_2O_{(l)}$
mole ratio	1	2		1	1	1
molar mass	100.09 g/mol	36.46 g/mol		110.98 g/mol	44.01 g/mol	18.02 g/mol
given	15.7 g				*m*	

Step 1 Balanced equation: $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(s)} + CO_{2(g)} + H_2O_{(\ell)}$

Step 2 Number of moles of $CaCO_3 = \frac{15.7 \cancel{g}\ CaCO_3}{100.09 \cancel{g}/mol} = 0.159$ mol $CaCO_3$

Number of moles of CO_2 produced $= \frac{0.159 \text{ mol} \times 1}{1} = 0.159$ mol CO_2

Theoretical yield of $CO_2 = 0.159\ \cancel{mol}\ CO_2 \times 44.01\ g/\cancel{mol} = 6.99$ g

Step 3 Actual yield = 6.99 g $CO_2 \times 0.815 = 5.69$ g CO_2

Check Your Solution

$\frac{5.69 \cancel{g}\ CO_2}{44.01 \cancel{g}/mol} = 0.129$ mol of actual CO_2 produced.

Actual mole / theoretical mol $= \frac{0.129}{0.159} = 0.811 = 81.1\%$. This is close to the 81.5% value given in the question. The answer is reasonable.

37. Problem

Mercury, in its elemental form or in a chemical compound is highly toxic. Water-soluble mercury compounds, such as mercury(II) nitrate, can be removed from industrial wastewater by adding sodium sulfide to the water, which forms a precipitate of mercury(II) sulfide, which can then be filtered out.

$$Hg(NO_3)_{2(aq)} + Na_2S_{(aq)} \rightarrow HgS_{(s)} + 2NaNO_{3(aq)}$$

If 3.45×10^{23} formula units of $Hg(NO_3)_2$ are reacted with excess Na_2S, what mass of Na_2S, what mass of HgS can be expected if this process occurs with 97.0% yield?

What Is Required?

You need to find the actual yield of HgS produced.

What Is Given?

Reactant $Hg(NO_3)_2 = 3.45 \times 10^{23}$ formula units.
The percentage yield of product HgS = 97.0%.
The balanced equation is given.
You know Avogadro's number = 6.02×10^{23} formula units/mol.

Plan Your Strategy

Step 1 Determine the theoretical yield of HgS. Divide the number of formula units of $Hg(NO_3)_2$ by Avogadro's number to obtain the mole amount. Use this value and the mole ratio of $Hg(NO_3)_2$ to HgS from the balanced equation to equate and solve for the number of moles of HgS produced. Multiply this value by the molar mass of HgS to obtain the theoretical yield in grams.

Step 2 Multiply the theoretical yield by the percentage yield (converted to a decimal) to obtain the actual yield in grams.

Act on Your Strategy

	$Hg(NO_3)_{2(aq)}$	$+ Na_2S_{(aq)}$	$\rightarrow$	$HgS_{(s)}$	$+ 2NaNO_{3(aq)}$
mole ratio	1	1		1	2
molar mass	324.61 g/mol	55.06 g/mol		232.66 g/mol	85 g/mol
given	3.45×10^{23} formula units			m	

Step 1 Number of moles of $Hg(NO_3)_2 = \dfrac{(3.45 \times 10^{23})\ \cancel{\text{formula units}}}{(6.02 \times 10^{23})\ \cancel{\text{formula units}}/\text{mol}}$

$= 0.573$ mol

Number of moles of HgS produced $= \dfrac{0.573 \text{ mol} \times 1}{1} = 0.573$ mol

Theoretical yield, $m = 0.573\ \cancel{\text{mol}}\ \text{HgS} \times 232.66\ \text{g}/\cancel{\text{mol}} = 133.3$ g HgS

Step 2 Actual yield = 133.3 g HgS × 0.97 = 129.3 g HgS

Check Your Solution

$\dfrac{129.3\ \cancel{\text{g}}\ CO_2}{232.66\ \cancel{\text{g}}/\text{mol}} = 0.556$ mol of actual HgS produced.

Actual mole / theoretical mol $= \dfrac{0.556}{0.573} = 0.970 = 97.0\%$. This is equal to the value given in the question.

38. Problem

An impure sample of silver nitrate, $AgNO_3$, has a mass of 0.340 g. It is dissolved in water and then treated with excess hydrogen chloride, $HCl_{(aq)}$. This results in the formation of a precipitate of silver chloride, AgCl.

$$AgNO_{3(aq)} + HCl_{(aq)} \times AgCl_{(s)} + HNO_{3(aq)}$$

The silver chloride is filtered, and any remaining hydrogen chloride is washed away. Then the silver chloride is dried. If the mass of the dry silver chloride is measured to be 0.213 g, what mass of silver nitrate was contained in the original (impure) sample?

What Is Required?

You need to find the total mass of the original (impure) silver nitrate.

What Is Given?

Reactant silver nitrate = 0.340 g.
Actual yield of the silver chloride = 0.213 g.
The balanced equation is given. You can assume the reaction proceeds to completion under the excess HCl.

Plan Your Strategy

Divide the actual yield of silver chloride by its molar mass to obtain the number of moles. Use this value and the mole ratio of $AgNO_3$ and AgCl from the balanced equation to equate and solve for the number of moles of $AgNO_3$. Multiply this value by the molar mass of silver nitrate to obtain the mass.

Act on Your Strategy

	$AgNO_{3(aq)}$	$+ HCl_{(aq)}$	$\rightarrow$	$AgCl_{(s)}$	$+ HNO_{3(aq)}$
mole ratio	1	1		1	1
molar mass	169.88 g/mol	36.46 g/mol		143.32 g/mol	63.02 g/mol
given	m			0.213 g	

Number of moles of of AgCl produced $= \dfrac{0.123\ \cancel{\text{g}}\ \text{AgCl}}{143.32\ \cancel{\text{g}}/\text{mol}} = 0.0015$ mol

Number of moles of $AgNO_3$ required $= \dfrac{0.0015 \text{ mol} \times 1}{1} = 0.0015$ mol

Therefore, $m = 0.0015\ \cancel{\text{mol}}\ AgNO_3 \times 169.88\ \text{g}/\cancel{\text{mol}} = 0.254$ g.

Therefore, 0.254 g of silver nitrate in the original sample.

Check Your Solution

[Note that the mass of the original sample was not needed for the calculation.] Work backwards. The mass of silver nitrate calculated divided by the number of moles should give the molar mass. $\frac{0.254 \text{ g}}{0.0015 \text{ mol}} = 169.33 \text{ g/mol}$. This is close to the value in the calculations. Your answer is reasonable.

39. Problem

Copper metal is mined as one of the several copper-containing ores. One of these ores contains copper in the form of malachite. Malachite exists as a double salt, $Cu(OH)_2 \bullet CuCO_3$. It can be thermally decomposed as 200°C to yield copper(II) oxide, carbon dioxide gas, and water vapour.

$$Cu(OH)_2 \bullet CuCO_{3(s)} \rightarrow 2CuO_{(s)} + CO_{2(g)} + H_2O_{(g)}$$

(a) 5.000 kg of malachite ore, containing 5.20% malachite, $Cu(OH)_2 \bullet CuCO_3$, is thermally decomposed. Calculate the mass of copper(II) oxide that is formed. Assume 100% reaction.

(b) Suppose that the reaction has a 78.0% yield, due to incomplete decomposition. How many grams of CuO would be produced?

What Is Required?

(a) You need to find the theoretical yield of CuO formed in the reaction.

(b) You need to find the actual yield of CuO at the given percentage yield.

What Is Given?

(a) Malachite ore = 5.000 kg; malachite composition in ore = 5.20%. Reaction proceeds to 100%.

(b) Percentage yield of CuO is 78.0%.

Plan Your Strategy

(a) Multiply the mass of malachite ore by the given percentage (expressed as a decimal) to get the mass of the malachite. Convert the kg quantity to gram. Divide this mass in grams by the molar mass of malachite to obtain the number of moles. Use this value and the mole ratio of $Cu(OH)_2 \bullet CuCO_3$ and CuO from the balanced equation to equate and solve for the number of moles of CuO. Multiply this value by the molar mass of CuO to obtain the theoretical yield in grams.

(b) Multiply the theoretical yield obtained in (a) by the percentage yield (expressed as a decimal) to get the actual yield in grams.

Act on Your Strategy

	$Cu(OH)_2 \bullet CuCO_{3(s)}$	$\rightarrow$	$2CuO_{(s)}$	$+ CO_{2(g)}$	$+ H_2O_{(g)}$
mole ratio	1		2	1	1
molar mass	221.13 g/mol		79.55 g/mol	44.01 g/mol	18.02 g/mol
given	5.000 kg				

(a) Mass of $Cu(OH)_2 \bullet CuCO_3$ in malachite ore = $5.0 \times 10^3 \text{ g} \times 0.052 = 260 \text{ g}$

Number of moles of $Cu(OH)_2 \bullet CuCO_3 = \frac{260 \cancel{\text{g}}\ Cu(OH)_2 \bullet CuCO_3}{221.13 \cancel{\text{g}}/\text{mol}} = 1.18 \text{ mol}$

Number of moles of CuO produced = $\frac{1.18 \text{ mol} \times 2}{1} = 2.36 \text{ mol}$

Mass CuO produced = $2.36 \cancel{\text{mol}} \times 79.55 \text{ g}/\cancel{\text{mol}} = 187.1 \text{ g}$

The theoretical yield of CuO is 187.1 g.

(b) Actual yield = 187.1 g CuO × 0.78 = 145.9 g

Check Your Solution

The actual yield in mole amount is $\frac{145.9\ \cancel{g}}{79.55\ \cancel{g}/mol} = 1.834$ mol.

$\frac{\text{Actual mole}}{\text{theoretical mol}} = \frac{1.834}{2.36} = 0.777 = 77.7\%$. This closely matches the percentage yield of 78% given in the question. The result is reasonable.

40. Problem

Ethylene oxide, C_2H_4O, is a multi-purpose industrial chemical used, among other things, as a rocket propellant. It can be prepared by reacting ethylene bromohydrin, C_2H_5OBr, with sodium hydroxide.

$$C_2H_5OBr + NaOH \rightarrow C_2H_4O + NaBr + H_2O$$

If this reaction proceeds with an 89% yield, what mass of C_2H_4O can be obtained when 3.61×10^{23} molecules of C_2H_5OBr react with excess sodium hydroxide?

What Is Required?

You need to find the actual yield of the C_2H_4O produced.

What Is Given?

Reactant $C_2H_5OBr = 3.61 \times 10^{23}$ molecules
Reactant NaOH is in excess
Percentage yield of $C_2H_4O = 89\%$
You are given the balanced equation.
You know Avogadro's number $= 6.02 \times 10^{23}$ molecules/mol.

Plan Your Strategy

First determine the theoretical yield. Divide the given number of molecules of C_2H_5OBr by the Avogadro number to obtain its number of moles. Use this value and the mole ratio of C_2H_5OBr and C_2H_4O from the balanced equation to equate and solve for the number of moles of C_2H_4O. Multiply this value by the molar mass of C_2H_4O to obtain the theoretical yield in grams. Multiply the theoretical yield by the percentage yield (expressed as a decimal) to get the actual yield in grams.

Act on Your Strategy

	C_2H_5OBr	+NaOH	$\rightarrow$	C_2H_4O	+ NaBr	+ H_2O
mole ratio	1	1		1	1	1
molar mass	124.97 g/mol	40 g/mol		44.06 g/mol	102.89 g/mol	18.02 g/mol
given	3.61×10^{23} molecules			*m*		

Number of moles of $C_2H_5OBr = \frac{3.61 \times 10^{23}\ \cancel{\text{molecules}}\ C_2H_5OBr}{6.02 \times 10^{23}\ \cancel{\text{molecules}}/\text{mol}} = 0.6$ mol

Number of moles of C_2H_4O produced $= \frac{0.6\ \text{mol} \times 1}{1} = 0.6$ mol

Hence, *m* of $C_2H_4O = 0.6\ \cancel{mol}\ C_2H_4O \times 44.06\ g/\cancel{mol} = 26.41$ g

Actual yield $= 26.41$ g $C_2H_4O \times 0.89 = 23.5$ g C_2H_4O

Check Your Solution

The actual yield in mole amount is $\frac{23.5\ \cancel{g}}{44.06\ \cancel{g}/mol} = 0.533$ mol.

$\frac{\text{Actual mole}}{\text{theoretical mol}} = \frac{0.533}{0.6} = 0.889 = 88.9\%$. This closely matches the percentage yield of 89% given in the question. The result is reasonable.