

## Chapter 7 Study Guide Answers

7.1

1a. The SI unit that measures the amount of substance is the mole. A mole of any substance is composed of Avogadro's number ( $6.02 \times 10^{23}$ ) of representative particles. The representative particle of most element is the atom. A molecule is the representative particle of diatomic elements and molecular compounds. The representative particle of ionic compounds is a formula unit.

2a. The molar mass of a substance is the mass in grams of one mole of that substance.

a.  $22.99 \text{ g/mol}$

b.  $58.44 \text{ g/mol}$

c.  $18.02 \frac{\text{g}}{\text{mol}}$

7.2

3a.  $45.98 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 2 \text{ mol Na}$

b.  $3 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 54.06 \text{ g H}_2\text{O}$

4a.  $9.3 \times 10^{15} \text{ atoms Pb} \times \frac{1 \text{ mol Pb}}{6.02 \times 10^{23} \text{ atoms Pb}} = 1.5 \times 10^{-8} \text{ mol Pb}$

4b.  $9.3 \times 10^{15} \text{ atoms Pb} \times \frac{207.2 \text{ g Pb}}{6.02 \times 10^{23} \text{ atoms Pb}} = 3.2 \times 10^{-6} \text{ g Pb}$

c.  $2.5 \text{ mol O}_2 \times \frac{22.4 \text{ L O}_2}{1 \text{ mol O}_2} = 56 \text{ L O}_2$

$$4d \quad 2.5 \text{ mol } O_2 \times \frac{6.02 \times 10^{23} \text{ molecules } O_2}{1 \text{ mol } O_2} = 1.5 \times 10^{24} \text{ molecules } O_2$$

1.3

5a  $C_6H_{12}O_6$  % Composition

1st calculate molar mass

$$C: 6 \times 12.01 = 72.06$$

$$H: 12 \times 1.01 = 12.12$$

$$O: 6 \times 16.00 = \frac{96.00}{180.18 \frac{g}{mol}}$$

$$\% C = 72.06 \div 180.18 = 0.3999 \times 100\% = \boxed{40\%}$$

$$\% H = 12.12 \div 180.18 = 0.067 \times 100\% = \boxed{7\%}$$

$$\% O = 96.00 \div 180.18 = 0.53 \times 100\% = \boxed{53\%}$$

$$6a \quad 94.1 \text{ g } O \times \frac{1 \text{ mol } O}{16 \text{ g } O} = 5.88 \text{ mol } O \quad \div 5.84 = 1$$

$$5.9 \text{ g } H \times \frac{1 \text{ mol } H}{1.01 \text{ g } H} = 5.8416 \text{ mol } H \quad \div 5.84 = 1$$

Empirical formula is: HO

Calculate <sup>empirical</sup> molar mass of HO = 17.01

Compare to molecular mass of HO @ 34.00

$$\frac{34.00}{17.01} = 2 \quad \text{so the molecular formula is } 2(HO) = \boxed{H_2O_2} = \text{hydrogen peroxide}$$

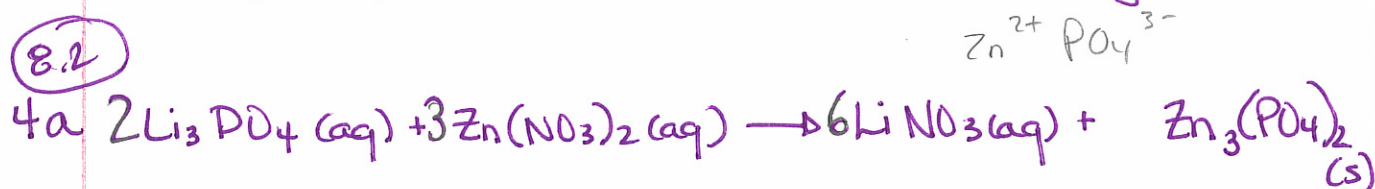
Note: H and O were not calculated as diatomic (HO-Br-Cl twins) because it is a compound.

## Chapter 8 Study Guide Answers

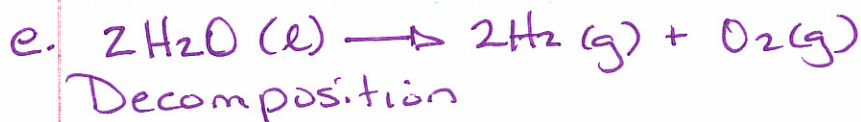
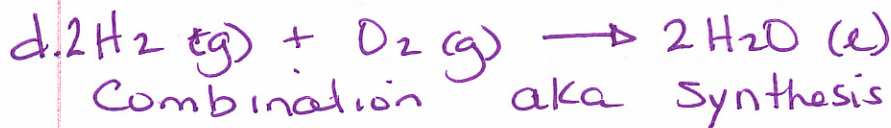
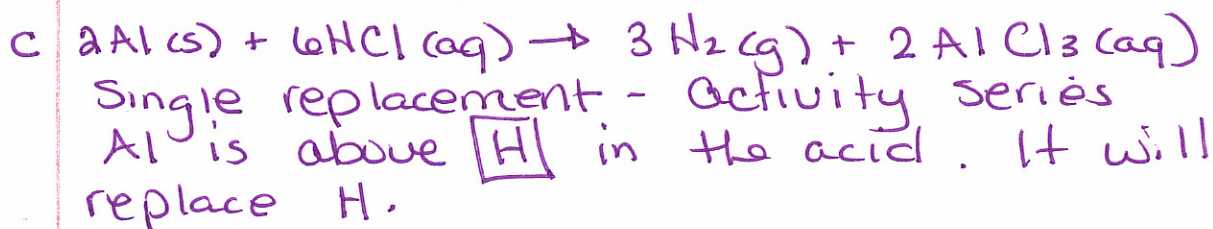
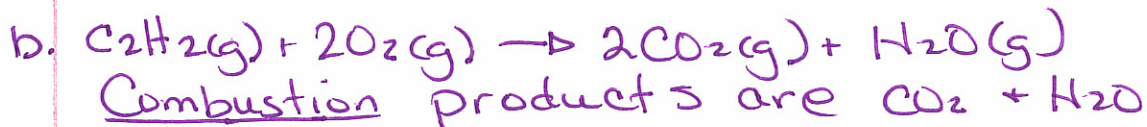
8.1



8.2

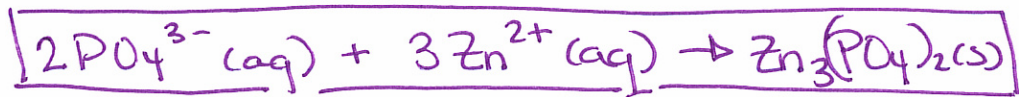
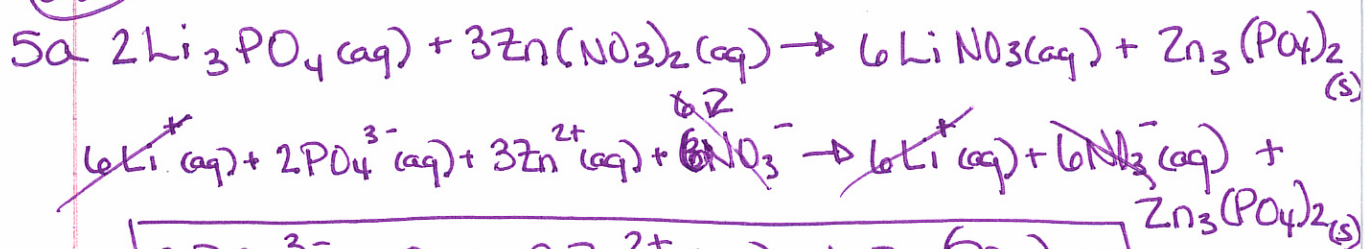


One of the products of a double replacement reaction must be either a solid (which zinc phosphate is) OR water - usually formed in double replacement rxns which are acid-base reactions or a gas.





Ch 8 Continued  
8.3



- 1: Complete the equation
- 2: Balance equation
- 3: Add States (aq) (l) (s) etc...
- 4: Take (aq) apart into ions
- 5: Cross off spectator ions
- 6: Reform Net Ionic Eq.
- 7: Check charges

## 9.1 The Arithmetic of Equations

1a.  $288 \text{ tricycles} \times \frac{3 \text{ wheels}}{1 \text{ tricycle}} = 864 \text{ wheels}$

$288 \text{ tricycles} \times \frac{1 \text{ seat}}{1 \text{ tricycle}} = 288 \text{ seats}$

$288 \text{ tricycles} \times \frac{2 \text{ pedals}}{1 \text{ tricycle}} = 576 \text{ pedals}$

2a.

<u>moles</u>	<u>molecules</u>	<u>volumes</u>
$N_2 = 2$	2	$2 \times 22.4 = 44.8 \text{ L}$
$O_2 = 3$	3	$3 \times 22.4 = 67.2 \text{ L}$
$N_2O_3 = 2$	2	$2 \times 22.4 = 44.8 \text{ L}$

mass Reactants :  $2 \text{ mol } N_2 = 2(14.01 \times 2) = 56.04 \text{ g}$   
 $3 \text{ mol } O_2 = 3(16.00 \times 2) = 96.00 \text{ g}$   
152.04 g

Products :  $2 \text{ mol } N_2O_3 = 2[(2 \times 14.01) + (3 \times 16.00)]$   
= 152.04 g

2b.  $14 \text{ mol } FeCl_3 \times \frac{3 \text{ mol } Cl_2}{2 \text{ mol } FeCl_3} = 21 \text{ mol } Cl_2$

## 9.2 Chemical Calculations

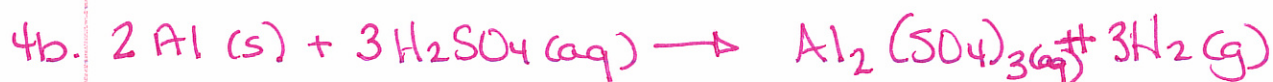


$$4.8\text{ g O}_2 \times \frac{1\text{ mol O}_2}{32.00\text{ g O}_2} \times \frac{2\text{ mol CO}}{1\text{ mol O}_2} \times \frac{22.4\text{ L CO}}{1\text{ mol CO}} = 6.72\text{ L CO}$$



$$2.1 \times 10^{24} \text{ molecules O}_2 \times \frac{1\text{ mol O}_2}{6.02 \times 10^{23} \text{ molecules O}_2} \times \frac{4\text{ mol NH}_3}{7\text{ mol O}_2}$$

$$\times \frac{17.04\text{ g NH}_3}{1\text{ mol NH}_3} = \boxed{33.97\text{ g NH}_3}$$



$$250\text{ g H}_2\text{SO}_4 \times \frac{1\text{ mol H}_2\text{SO}_4}{98.09\text{ g H}_2\text{SO}_4} \times \frac{1\text{ mol Al}_2(\text{SO}_4)_3}{3\text{ mol H}_2\text{SO}_4} \times \frac{342.17\text{ g Al}_2(\text{SO}_4)_3}{1\text{ mol Al}_2(\text{SO}_4)_3} =$$

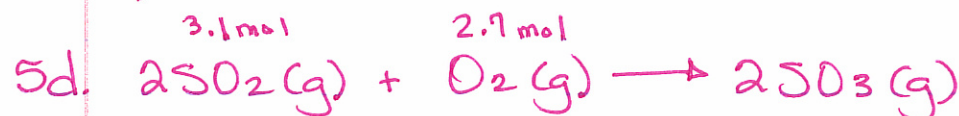
## 9.3 LR + % Yield

$$= \boxed{290.69\text{ g Al}_2(\text{SO}_4)_3}$$

5a. limiting reagent

5b. used up.

5c. product



$$3.1\text{ mol SO}_2 \times \frac{1\text{ mol O}_2}{2\text{ mol SO}_2} = \boxed{1.55\text{ mol O}_2 \text{ needed, so SO}_2 \text{ is LR}}$$

5e.  $3.1\text{ mol SO}_2 \times \frac{2\text{ mol SO}_3}{2\text{ mol SO}_2} = 3.10\text{ mol SO}_3 \text{ can be formed (248.22g)}$

$$2.7\text{ mol O}_2 - 1.55\text{ mol O}_2 \text{ used} = 1.15\text{ mol O}_2 \text{ in excess (36.80g)}$$

9.3 Continued



$$36 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} \times \frac{1 \text{ mol CS}_2}{4 \text{ mol H}_2} \times \frac{22.4 \text{ L CS}_2}{1 \text{ mol CS}_2} = 9 \text{ L CS}_2$$

Since only 9 L CS<sub>2</sub> were required, 3 L CS<sub>2</sub> is left over. (making CS<sub>2</sub> the excess reagent)



$$54.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol Cu}}{2 \text{ mol Al}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 190.79 \text{ g Cu}$$

$$319 \text{ g CuSO}_4 \times \frac{1 \text{ mol CuSO}_4}{159.62 \text{ g CuSO}_4} \times \frac{3 \text{ mol Cu}}{3 \text{ mol CuSO}_4} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}}$$

$$= 127.00 \text{ g Cu}$$

6a. theoretical yield

b. maximum

c. actual yield

1.87g

(produced)  
4.65g



Theoretical Yield:

$$1.87 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol Cu}}{2 \text{ mol Al}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 6.61 \text{ g Cu}$$

Actual yield:

4.65g Cu

$$\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\% = \frac{4.65 \text{ g}}{6.61 \text{ g}} \times 100\% = 70.4\% \text{ Yield}$$