CH 222 Chapter Seven Part II Concept Guide

1. Balancing Chemical Equations

Description

When chlorine gas, Cl_2, is added to solid phosphorus, P_4, a reaction occurs to produce liquid phosphorus trichloride, PCl_3, and heat.

$P_4(s) + 6\ Cl_2(g) \rightarrow 4\ PCl_3(l)$

Question

If you want to make 150.0 grams of PCl_3, how many moles of chlorine gas must you begin with?

Approach

The balanced chemical equation relates the number of moles of each species involved in the reaction. If we can determine the number of moles of PCl_3 we want to make, we can use the stoichiometric coefficients in the balanced equation to determine the quantity in moles of each reactant that is necessary.

Solution

Step 1. To determine the quantity of PCl_3 in units of moles, we must make use of the compound's molar mass, which is 137.33 g/mol. We convert units using the molar mass as a unit conversion factor, remembering to place the unit we are converting to in the numerator of the ratio.

Convert PCl_3 from grams to moles.

$(150.0\ g\ PCl_3)\ (1\ mol\ PCl_3/137.33\ g\ PCl_3) = 1.092\ mol\ PCl_3$

Step 2. We can determine the quantity in moles of Cl_2 needed to make 1.092 mol of PCl_3 using the fact that 6 mol of Cl_2 are used to form 4 mol of PCl_3. These are used in the form of a ratio to convert from one to the other.

Determine the number of moles of Cl_2 needed.

$(1.092\ mol\ PCl_3)(6\ molCl_2/4\ mol\ PCl_3) = 1.638\ mol\ Cl_2$

We must begin with 1.638 moles of Cl_2 gas to make 150.0 grams of PCl_3.

2. Balancing Chemical Equations

Description

Butane reacts with oxygen gas to produce carbon dioxide and water vapor.

$C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Question

What is the balanced form of this reaction equation?

Approach

Balance the numbers of elements on each side of the equation one compound at a time.

Solution

Step 1. Start with carbon. There are four carbon atoms on the left side of the arrow, so we need to put a 4 in front of the CO_2 on the right.

$C_4H_{10}(g) + O_2(g) \rightarrow 4\ CO_2(g) + H_2O(g)$

Step 2. Increase the number of hydrogen atoms on the right to match the number found on the left.

$C_4H_{10}(g) + O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$

Step 3. Oxygen is the last element to be balanced. There are a total of thirteen oxygen atoms on the right side of the equation. To balance this, we must put 13/2 in front of the O_2 on the left.

$C_4H_{10}(g) + 13/2\ O_2(g) \rightarrow 4\ CO_2(g) + 5\ H_2O(g)$

Step 4. Although the equation is now balanced, we are not finished. Because we are looking at this reaction at the molecular scale, we cannot talk about half molecules. Therefore, the whole reaction equation must be doubled.

$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$

3. Balancing Chemical Equations

Description
Xenon Tetrafluoride gas and water react to give xenon, oxygen, and hydrogen fluoride gases.

Question
What is the balanced form of this reaction?

Approach
Write out the reaction. Then balance the numbers of atoms on each side of the equation one element at a time.

Solution
Step 1. Write out the reaction. Indicate the physical state of each reactant and product.

$XeF_4(g) + H_2O(l) \rightarrow Xe(g) + O_2(g) + HF(g)$

Step 2. It is best to start with an element that appears in only one species on each side of the equation. Start by writing a coefficient of 4 for HF, thus obtaining 4 fluorine atoms on each side.

$XeF_4(g) + H_2O(l) \rightarrow Xe(g) + O_2(g) + 4\ HF(g)$

Step 3. Now consider the xenon atoms. There is one xenon atom on each side, therefore, the xenon atoms are balanced.

$XeF_4(g) + H_2O(l) \rightarrow Xe(g) + O_2(g) + HF(g)$

Step 4. There are 2 hydrogen atoms on the left side of the reaction, and 4 hydrogen atoms on the right side. To obtain 4 hydrogen atoms on the left, write a coefficient of 2 for H_2O.

$XeF_4(g) + 2\ H_2O(l) \rightarrow Xe(g) + O_2(g) + 4\ HF(g)$

Step 5. Finally, consider oxygen, the last element to be balanced. There are now 2 oxygen atoms on the left, and 2 oxygen atoms on the right, thus the oxygen atoms are balanced as is. This is the final balanced equation for the reaction of xenon tetrafluoride gas and water to give xenon, oxygen, and hydrogen fluoride gases.

$XeF_4(g) + 2\ H_2O(l) \rightarrow Xe(g) + O_2(g) + 4\ HF(g)$

4. Mole Ratios

Problem
Calcium carbide is produced by the reaction of calcium oxide with carbon at high temperatures:

$CaO(s) + 3\ C(s) \rightarrow CaC_2(s) + CO(g)$

What are the mole ratios that give
(a) the amount of calcium carbide produced by each mole of calcium oxide that reacts,
(b) the amount of carbon required by each mole of calcium oxide that reacts, and
(c) the amount of calcium carbide produced by each mole of carbon that reacts.

Approach
Use the balanced chemical equation to determine each mole ratio.

Solution
(a) 1 mol CaC_2 / 1 mol CaO
(b) 3 mol C / 1 mol CaO
(c) 1 mol CaC_2 / 3 mol C

5. Mole Ratios

Problem
How many moles of silicon dioxide would be required to produce 4.5 mol of P_4?

$2\ Ca_3(PO_4)_2(s) + 6\ SiO_2(s) + 10\ C(s) \rightarrow 6\ CaSiO_3(l) + 10\ CO(g) + P_4(g)$

Solution
(4.5 mol P_4)(6 mol SiO_2 / 1 mol P_4) = 27 mol SiO_2

6. Stoichiometry

Question
What mass of hydrogen fluoride can be produced by the reaction of 15.0 g of calcium fluoride with excess sulfuric acid?

$CaF_2(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + 2\ HF(aq)$

Approach
Convert known masses to moles. Use mole ratios to find the unknown in number of moles. Finally, convert from moles to the desired quantity, which in this case is mass.

Solution
Step 1. Convert the mass of calcium fluoride into moles using molecular mass.

(15.0 g CaF_2)(1 mol CaF_2 / 78.1 g CaF_2) = 0.192 mol CaF_2

Step 2. Multiply the moles of calcium fluoride times the mole ratio of calcium fluoride to hydrogen fluoride.

(0.192 mol CaF_2)(2 mol HF / 1 mol CaF_2) = 0.384 mol HF

Step 3. Now, multiply the moles of hydrogen fluoride times its molecular weight to convert from moles HF to grams HF.

(0.384 mol HF)(20.0 g HF / 1 mol HF) = 7.68 g HF

7.68 g hydrogen fluoride can be produced using 15.0 g calcium fluoride and excess sulfuric acid.

7. Limiting Reactants

Question
Suppose 378 g of CO are mixed with 60.0 g of H_2 to form CH_3OH. Which is the limiting reactant?

Approach
Write the balanced equation. Then, convert known quantities to number of moles. Use mole ratios to find the number of moles of CO required for all the H_2 to react, and the number of moles of H_2 required for all the CO to react.

Solution
Step 1. Write the balanced chemical equation.

$CO(g) + 2\ H_2(g) \rightarrow CH_3OH(l)$

Step 2. Convert mass of CO to moles of CO, and mass of H_2 to moles of H_2.

(378 g CO)(1 mol CO / 28.01 g CO) = 13.5 mol CO

(60.0 g H_2)(1 mol H_2 / 2.02 g H_2) = 29.7 mol H_2

Step 3. Use two mole ratios to determine (a) the number of moles of CO necessary for all the H_2 to react, and (b) the number of moles of H_2 for all the CO to react.

(13.5 mol CO)(2 mol H_2 / 1 mol CO) = 27.0 mol H_2

(29.7 mol H_2)(1 mol CO / 2 mol H_2) = 14.9 mol CO

The amount of CO available, 13.5 mol, is less than the required amount, 14.9 mol. The amount of H_2 available, however, is sufficient. Therefore, CO is the limiting reactant.

8. Percent Yield

Problem

When 50.0 g of calcium carbide reacted with an excess of water, 14.2 g of ethyne (acetylene) were produced. What is the percent yield of ethyne for the reaction?

$CaC_2(s) + 2\ H_2O(l) \rightarrow Ca(OH)_2(aq) + C_2H_2(g)$

Approach

First, calculate the theoretical yield, the stoichiometric amount that could be produced. Then, calculate the percent yield using the following equation:

% yield = actual yield / theoretical yield * (100%)

Solution

Step 1. Calculate the theoretical yield of ethyne.

(50.0 g CaC_2)(1 mol CaC_2 / 64.10 g CaC_2)(1 mol C_2H_2 / 1 mol CaC_2)(26.04 g C_2H_2 / 1 mol C_2H_2) = 20.3 g C_2H_2

Step 2. Divide the actual yield by the theoretical yield and multiply by 100 to calculate the percent yield.

% yield = (14.2 g ethyne / 20.3 g ethyne)(100) = 70.0%

9. Empirical Formula

Problem

0.225 g of magnesium burns in nitrogen to form 0.312 g of magnesium nitride. What is the empirical formula of magnesium nitride?

Approach

We would expect Mg^{2+} and N^{3-} ions to react to form Mg_3N_2, yet we need to determine if this is in fact correct given the experimental data. We need to calculate the number of moles of magnesium and nitrogen.

Solution

Step 1. Calculate the moles of magnesium.

(0.225 g Mg)(1 mol Mg / 24.32 g Mg) = 0.00925 mol Mg

Step 2. Calculate the mass of nitrogen.

0.312 g magnesium nitride - 0.225 Mg = 0.0870 g N

Step 3. The number of moles of nitrogen can be calculated from its mass.

(0.0870 g N)(1 mol N / 14.01 g N) = 0.00621 mol N

Step 4. Divide all moles by the smallest number of moles.

0.00925 mol Mg / 0.00621 mol N = 1.49 mol Mg / N

0.00621 mol N / 0.00621 mol N = 1.00 mol N

The mole ratio is, therefore, 1.49 mol Mg : 1.00 mol N.
The empirical formula for magnesium nitride is $Mg_{1.5}N_{1}$. To obtain all whole numbers, multiply by a factor of 2: Mg_3N_2. This result suggests that the empirical formula of magnesium nitride is what was expected: Mg_3N_2.

10. Empirical Formula

Problem
Potassium dichromate contains three elements: potassium, chromium, and oxygen. A chemical analysis of a sample of potassium chromate resulted in 13.3 g K, 17.7 g Cr, and 19.0 g O. Calculate the empirical formula for potassium dichromate.

Approach
Calculate the number of moles of potassium, chromate, and oxygen. Then, divide all moles by the smallest number of moles to calculate the mole ratio.

Solution
Step 1. Calculate the number of moles of potassium, chromium, and oxygen.
(13.3 g K)(1 mol K / 39.10 g K) = 0.340 mol K
(17.7 g Cr)(1 mol Cr / 52.00 g Cr) = 0.340 mol Cr
(19.0 g O)(1 mol O / 16.00 g O) = 1.19 mol O

Step 2. Divide all moles by the smallest number of moles.
0.340 mol K / 0.340 mol K = 1.00 mol K
0.340 mol Cr / 0.340 mol K = 1.00 mol Cr/K
1.19 mol O / 0.340 mol K = 3.50 mol O/K

The mole ratio is, therefore, 1.00 mol K:1.00 mol Cr: 3.50 mol O

Step 3. Multiply by a factor of 2 to convert all moles to whole numbers. This is the empirical formula for potassium dichromate: $K_2Cr_2O_7$.

11. Stoichiometry of Reacting Gases

Problem
Calculate the volume of sulfur dioxide produced at 25 ˚C and 1.00 atm by the combustion of 15.6 g of sulfur.
$S_8(s) + 8\ O_2(g) \rightarrow 8\ SO_2(g)$ *Note that* 1 mol of gas at 25 ˚C occupies 24.47 L

Approach
The first step is to convert to moles of S_8. Second, use a mole ratio to convert moles S_8 to SO_2. Then, use molar volume and the number of moles of SO_2 to calculate the volume of SO_2.

Solution
Step 1. Convert grams to moles of S_8.
(15.6 g S_8)(1 mol S_8 / 256.48 grams S_8) = 0.0608 mol S_8

Step 2. Use a mole ratio to convert from moles of sulfur to moles of sulfur dioxide.
(0.0608 mol S_8)(8 mol SO_2 / 1 mol S_8) = 0.486 mol SO_2

Step 3. Finally, use molar volume to convert moles of SO_2 to liters of SO_2.
0.486 mol SO_2 * 24.47 L/mol = 11.9 L SO_2

12. Limiting Reactants and Molarity

Question
What mass of hydrogen will be produced from the reaction of 8.0 g of zinc with 20 mL of 5.0 M hydrochloric acid?

Approach
Write the balanced equation. Then, convert known quantities to number of moles. Use mole ratios to find the unknown in number of moles. Finally, convert from moles to the desired quantity, which is in this case, mass.

Solution
Step 1. Write the balanced chemical equation.

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Step 2. Convert mass of Zn to moles of Zn. Also, convert the molarity of HCl to moles of HCl.

(8.0 g Zn)(1 mol Zn / 65.38 g Zn) = 0.122 mol Zn
(20 mL)(1 L / 1000 mL)(5.0 mol HCl / 1 L) = 0.10 mol HCl

Step 3. Use two mole ratios to determine (a) the number of moles of HCl necessary for all the Zn to react, and (b) the number of moles of Zn for all the HCl to react.

(0.122 mol Zn)(2 mol HCl / 1 mol Zn) = 0.244 mol HCl
(0.10 mol HCl)(1 mol Zn / 2 mol HCl) = 0.050 mol Zn

The amount of HCl available, 0.10 mol, is less than the required amount, 0.244 mol HCl. The amount of Zn available, however, is sufficient. Therefore, HCl is the limiting reactant, and the amount of H_2 formed is determined by the amount of HCl:

(0.10 mol HCl)(1 mol H_2 / 2 mol HCl) = 0.050 mol H_2

Step 4. To calculate the mass of H_2 produced in this reaction, multiply the moles of H_2 times its molecular weight.

(0.050 mol H_2)(2.02 g H_2 / 1 mol H_2) = 0.10 g H_2

13. Empirical Formula

Question
A 1.27 g sample of an oxide contains 0.55 g phosphorus and 0.72 g oxygen. What this oxide's empirical formula?

Approach
First, it is necessary to determine from the experimental data the number of moles of atoms of each element present. The simplest ratio is then found by dividing the numbers of moles of each element by the number of moles of the element present in the smallest amount.

Solution
Step 1. Calculate the number of moles of phosphorus and oxygen.

(0.55 g P)(1 mol P/30.97 g P) = 0.018 mol P
(0.72 g O)(1 mol O/16.00 g O) = 0.045 mol O

Step 2. Divide the numbers of moles by the number of moles of the element present in the smallest amount: P.

0.018 mol P/0.018 mol P = 1.0
0.045 mol O/0.018 mol P = 2.5

Thus, the empirical formula is $P_{1.0}O_{2.5}$.

Step 3. Double all numbers in the formula to convert the fraction to a whole number.

$2(P_{1.0}O_{2.5}) = P_2O_5$

The empirical formula is P_2O_5.

14. Percent Composition

Question

A 2.91 g sample of potassium metal when burned in oxygen formed a compound weighing 6.11 g and containing only potassium and oxygen. What is the percent composition of each element in this compound?

Approach

The percent composition is the percent by mass of each element in the compound, which is given by the mass of that element divided by the total mass of the compound, times 100.

Solution

The percent potassium in this compound is:

%K = (2.91 g / 6.11 g compound)(100%) = 47.6 %

The percent oxygen in this compound is:

6.11 g compound - 2.91 g K = 3.20 g O

%O = (3.20 g / 6.11 g compound)(100%) = 52.4 %

The percent composition of the compound is 47.6% potassium and 52.4% oxygen.