# CH 221 Chapter Four Part I 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: $K_2Cr_2O_7$.