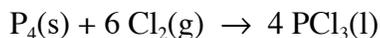


# CH 221 Chapter Four Part I Concept Guide

## 1. Balancing Chemical Equations

### Description

When chlorine gas,  $\text{Cl}_2$ , is added to solid phosphorus,  $\text{P}_4$ , a reaction occurs to produce liquid phosphorus trichloride,  $\text{PCl}_3$ , and heat.



### Question

If you want to make 150.0 grams of  $\text{PCl}_3$ , 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  $\text{PCl}_3$  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  $\text{PCl}_3$  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  $\text{PCl}_3$  from grams to moles.

$$(150.0 \text{ g PCl}_3) (1 \text{ mol PCl}_3/137.33 \text{ g PCl}_3) = 1.092 \text{ mol PCl}_3$$

**Step 2.** We can determine the quantity in moles of  $\text{Cl}_2$  needed to make 1.092 mol of  $\text{PCl}_3$  using the fact that 6 mol of  $\text{Cl}_2$  are used to form 4 mol of  $\text{PCl}_3$ . These are used in the form of a ratio to convert from one to the other.

Determine the number of moles of  $\text{Cl}_2$  needed.

$$(1.092 \text{ mol PCl}_3)(6 \text{ mol Cl}_2/4 \text{ mol PCl}_3) = 1.638 \text{ mol Cl}_2$$

We must begin with 1.638 moles of  $\text{Cl}_2$  gas to make 150.0 grams of  $\text{PCl}_3$ .

## 2. Balancing Chemical Equations

### Description

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



### Question

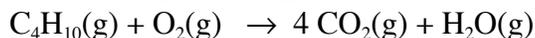
What is the balanced form of this reaction equation?

### Approach

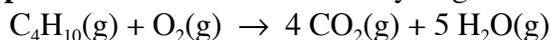
Balance the numbers of elements on each side of the equation one compound at a time.

## Solution

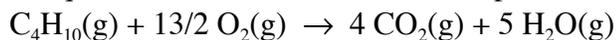
**Step 1.** Start with carbon. There are four carbon atoms on the left side of the arrow, so we need to put a 4 in front of the CO<sub>2</sub> on the right.



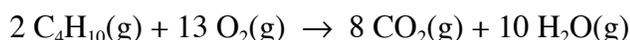
**Step 2.** Increase the number of hydrogen atoms on the right to match the number found on the left.



**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<sub>2</sub> on the left.



**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.



## 3. Balancing Chemical Equations

### Description

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

### Question

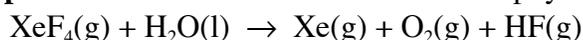
What is the balanced form of this reaction?

### Approach

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

### Solution

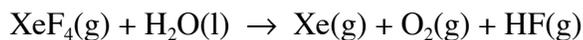
**Step 1.** Write out the reaction. Indicate the physical state of each reactant and product.



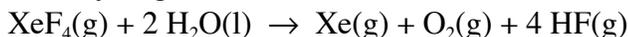
**Step 2.** It is best to start with an element that appears in only one species on each side of the equation. Start by writing a coefficient of 4 for HF, thus obtaining 4 fluorine atoms on each side.



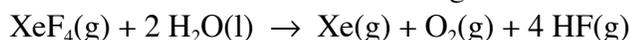
**Step 3.** Now consider the xenon atoms. There is one xenon atom on each side, therefore, the xenon atoms are balanced.



**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<sub>2</sub>O.



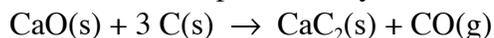
**Step 5.** Finally, consider oxygen, the last element to be balanced. There are now 2 oxygen atoms on the left, and 2 oxygen atoms on the right, thus the oxygen atoms are balanced as is. This is the final balanced equation for the reaction of xenon tetrafluoride gas and water to give xenon, oxygen,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



## 4. Mole Ratios

### Problem

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



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.

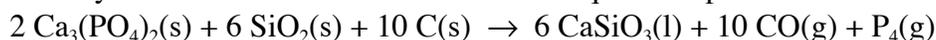
### Solution

- (a) 1 mol  $\text{CaC}_2$  / 1 mol  $\text{CaO}$
- (b) 3 mol  $\text{C}$  / 1 mol  $\text{CaO}$
- (c) 1 mol  $\text{CaC}_2$  / 3 mol  $\text{C}$

## 5. Mole Ratios

### Problem

How many moles of silicon dioxide would be required to produce 4.5 mol of  $\text{P}_4$ ?



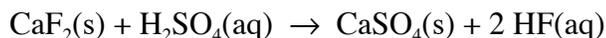
### Solution

$$(4.5 \text{ mol P}_4)(6 \text{ mol SiO}_2 / 1 \text{ mol P}_4) = 27 \text{ mol SiO}_2$$

## 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?



### 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.

$$(15.0 \text{ g CaF}_2)(1 \text{ mol CaF}_2 / 78.1 \text{ g CaF}_2) = 0.192 \text{ mol CaF}_2$$

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

$$(0.192 \text{ mol CaF}_2)(2 \text{ mol HF} / 1 \text{ mol CaF}_2) = 0.384 \text{ mol HF}$$

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

$$(0.384 \text{ mol HF})(20.0 \text{ g HF} / 1 \text{ mol HF}) = 7.68 \text{ g HF}$$

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

## 7. Limiting Reactants

### Question

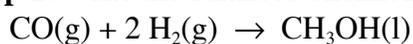
Suppose 378 g of CO are mixed with 60.0 g of H<sub>2</sub> to form CH<sub>3</sub>OH. Which is the limiting reactant?

### 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<sub>2</sub> to react, and the number of moles of H<sub>2</sub> required for all the CO to react.

### Solution

**Step 1.** Write the balanced chemical equation.



**Step 2.** Convert mass of CO to moles of CO, and mass of H<sub>2</sub> to moles of H<sub>2</sub>.

$$(378 \text{ g CO})(1 \text{ mol CO} / 28.01 \text{ g CO}) = 13.5 \text{ mol CO}$$

$$(60.0 \text{ g H}_2)(1 \text{ mol H}_2 / 2.02 \text{ g H}_2) = 29.7 \text{ mol H}_2$$

**Step 3.** Use two mole ratios to determine (a) the number of moles of CO necessary for all the H<sub>2</sub> to react, and (b) the number of moles of H<sub>2</sub> for all the CO to react.

$$(13.5 \text{ mol CO})(2 \text{ mol H}_2 / 1 \text{ mol CO}) = 27.0 \text{ mol H}_2$$

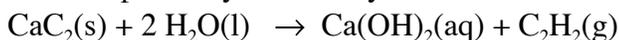
$$(29.7 \text{ mol H}_2)(1 \text{ mol CO} / 2 \text{ mol H}_2) = 14.9 \text{ mol CO}$$

The amount of CO available, 13.5 mol, is less than the required amount, 14.9 mol. The amount of H<sub>2</sub> 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?



### Approach

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

$$\% \text{ yield} = \text{actual yield} / \text{theoretical yield} * (100\%)$$

### Solution

**Step 1.** Calculate the theoretical yield of ethyne.

$$(50.0 \text{ g CaC}_2)(1 \text{ mol CaC}_2 / 64.10 \text{ g CaC}_2)(1 \text{ mol C}_2\text{H}_2 / 1 \text{ mol CaC}_2)(26.04 \text{ g C}_2\text{H}_2 / 1 \text{ mol C}_2\text{H}_2) = 20.3 \text{ g C}_2\text{H}_2$$

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

$$\% \text{ yield} = (14.2 \text{ g ethyne} / 20.3 \text{ g ethyne})(100) = 70.0\%$$

## 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  $\text{Mg}^{2+}$  and  $\text{N}^{3-}$  ions to react to form  $\text{Mg}_3\text{N}_2$ , 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.

$$(0.225 \text{ g Mg})(1 \text{ mol Mg} / 24.32 \text{ g Mg}) = 0.00925 \text{ mol Mg}$$

**Step 2.** Calculate the mass of nitrogen.

$$0.312 \text{ g magnesium nitride} - 0.225 \text{ Mg} = 0.0870 \text{ g N}$$

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

$$(0.0870 \text{ g N})(1 \text{ mol N} / 14.01 \text{ g N}) = 0.00621 \text{ mol N}$$

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

$$0.00925 \text{ mol Mg} / 0.00621 \text{ mol N} = 1.49 \text{ mol Mg} / \text{N}$$

$$0.00621 \text{ mol N} / 0.00621 \text{ mol N} = 1.00 \text{ mol N}$$

The mole ratio is, therefore, 1.49 mol Mg : 1.00 mol N.

The empirical formula for magnesium nitride is  $\text{Mg}_{1.5}\text{N}_1$ . To obtain all whole numbers, multiply by a factor of 2:  $\text{Mg}_3\text{N}_2$ . This result suggests that the empirical formula of magnesium nitride is what was expected:  $\text{Mg}_3\text{N}_2$ .

## 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.

$$(13.3 \text{ g K})(1 \text{ mol K} / 39.10 \text{ g K}) = 0.340 \text{ mol K}$$

$$(17.7 \text{ g Cr})(1 \text{ mol Cr} / 52.00 \text{ g Cr}) = 0.340 \text{ mol Cr}$$

$$(19.0 \text{ g O})(1 \text{ mol O} / 16.00 \text{ g O}) = 1.19 \text{ mol O}$$

**Step 2.** Divide all moles by the smallest number of moles.

$$0.340 \text{ mol K} / 0.340 \text{ mol K} = 1.00 \text{ mol K}$$

$$0.340 \text{ mol Cr} / 0.340 \text{ mol K} = 1.00 \text{ mol Cr/K}$$

$$1.19 \text{ mol O} / 0.340 \text{ mol K} = 3.50 \text{ mol O/K}$$

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:  $\text{K}_2\text{Cr}_2\text{O}_7$ .