CH 221 Chapter Seven 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. Net Ionic Equations

Problem
When aqueous solutions of aluminum nitrate and sodium carbonate are mixed, a precipitate of aluminum carbonate forms. Write the net ionic equation for the reaction between aluminum nitrate and sodium carbonate.

Approach
The general approach to writing net ionic equations is to write the complete, balanced equation, then write the equation in terms of the individual ions that are in solution. Finally, eliminate any spectator ions, taking care to cross out an equal number of ions on each side of the equation.

Solution

Step 1. Write out the formulas of the reactants and determine the products.

$Al(NO_3)_3(aq) + Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + NaNO_3(aq)$

Step 2. Balance the equation.

$2\ Al(NO_3)_3(aq) + 3\ Na_2CO_3(aq) \rightarrow Al_2(CO_3)_3(s) + 6\ NaNO_3(aq)$

Step 3. Write out the ions in solution for the aqueous compounds.

$2\ Al^{3+}(aq) + 6\ NO_3^{-}(aq) + 6\ Na^{+}(aq) + 3\ CO_3^{2-}(aq) \rightarrow$

$Al_2(CO_3)_3(s) + 6\ NO_3^{-}(aq) + 6\ Na^{+}(aq)$

Step 4. Cross out equal numbers of spectator ions-ions that appear on both sides of the equation that do not participate in the reaction:

$2\ Al^{3+}(aq) + \cancel{6\ NO_3^{-}(aq)} + \cancel{6\ Na^{+}(aq)} + 3\ CO_3^{2-}(aq) \rightarrow$

$Al_2(CO_3)_3(s) + \cancel{6\ NO_3^{-}(aq)} + \cancel{6\ Na^{+}(aq)}$

Step 5. The balanced net ionic equation is:

$2\ Al^{3+}(aq) + 3\ CO_3^{2-}(aq) \rightarrow Al_2(CO_3)_3(s)$

5. Reaction of Iron (III) Nitrate and Sodium Hydroxide

Question

Iron (III) nitrate reacts with sodium hydroxide to make iron (III) hydroxide and sodium nitrate. Does this reaction form a precipitate?

$Fe(NO_3)_3(aq) + 3\ NaOH(aq) \rightarrow Fe(OH)_3(?) + 3\ NaNO_3(?)$

Solution

Notice that we are considering the compounds that we'd have if the pairs of cations and anions were exchanged. We know that the original compounds are soluble, so an insoluble product would have to be either iron ions paired with hydroxide ions or sodium ions paired with nitrate ions.

If either of the products is insoluble, the reaction will produce a precipitate. Sodium salts are soluble, and nitrate salts are soluble. Therefore, sodium nitrate is soluble, and will not precipitate. All hydroxide ionic compounds are insoluble, with the exception of the alkali metal cations. Iron is not an alkali metal, so we expect iron (III) hydroxide to be insoluble. Thus, we expect a precipitate to form.

$Fe(NO_3)_3(aq) + 3\ NaOH(aq) \rightarrow Fe(OH)_3(s) + 3\ NaNO_3(aq)$

When the reaction is conducted, a brown precipitate of iron (III) hydroxide forms. Our prediction was correct. The net ionic equation is

$Fe^{3+}(aq) + 3\ OH^{-}(aq) \rightarrow Fe(OH)_3(s)$

6. Synthesis of Iron (II) Carbonate.

Problem

Propose a reaction that yields iron (II) carbonate as a product. This compound is insoluble.

Solution

In our desired product, iron (II) carbonate, the iron (II) ion serves as the cation and the carbonate ion is the anion. The reactants, therefore, must include an iron (II) cation and a carbonate anion, which when reacted, produce this insoluble compound. The other product must be soluble in order for pure iron (II) carbonate to be obtained.

Step 1. Start by writing the product.

$FeCO_3(s)$

Step 2. Write the net ionic equation formation of this product.

$Fe^{2+}(aq) + CO_3^{2-}(aq) \rightarrow FeCO_3(s)$

Step 3. Choose a soluble compound that will release Fe^{2+} ions when dissolved, and another compound that will release carbonate ion. Choices here are iron (II) nitrate and sodium carbonate, because all nitrate and sodium salts are soluble. (Note: there are several other compounds that would work here. Sodium and nitrate salts are just one possibility.)

$Fe(NO_3)_2(aq) + Na_2CO_3(aq)$

Step 4. Write out the balanced equation and indicate the physical state of each reactant and product. This is a final reaction for the synthesis of iron (II) carbonate using iron (II) nitrate and sodium carbonate.

$Fe(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow FeCO_3(s) + 2\ NaNO_3(aq)$

7. Acid-Base Reactivity: Determining Net Ionic Equations

Question

What is the net ionic equation for the acid-base reaction between HNO_3 and NaF?

Solution

Step 1. Write out the complete reaction.

$HNO_3(aq) + NaF(aq) \rightarrow HF(aq) + NaNO_3(aq)$

Step 2. Write out the reaction in terms of ions in solution.

$H^+(aq) + NO_3^-(aq) + Na^+(aq) + F^-(aq) \rightarrow HF(aq) + Na^+(aq) + NO_3^-(aq)$

HF is a weak acid, so it exists in solution mainly in the undissociated form. We therefore write it as HF instead of dissociated ions.

Step 3 . Cross out spectator ions.

$H^+(aq) + \cancel{NO_3}^-\ (aq) + \cancel{Na}^+(aq) + F^-(aq) \rightarrow HF(aq) + \cancel{Na}^+(aq) + \cancel{NO_3}^-(aq)$

Step 4. Write the net ionic equation.

$H^+(aq) + F^-(aq) \rightarrow HF(aq)$

8. Redox Reaction of NaCl

Question

When solid sodium and gaseous chlorine are combined in a flask, a vigorous reaction occurs to produce sodium chloride. Which species is oxidized and which is reduced?

Approach

To answer this question, we must examine the oxidation numbers of the elements before and after the reaction.

Solution

The charges on both sodium metal and chlorine gas are zero. Once the reaction occurs, and sodium chloride is produced, the charge of sodium is plus one, and the charge of chlorine is minus one.

	$2\ Na(s)$ +	$Cl_2(g) \rightarrow$	$2\ NaCl(s)$
charge per atom =	0	0	+1 -1

Na loses 1e$^-$ per atom. Na is oxidized.
Cl gains 1e$^-$ per Cl atom. Cl is reduced.

The sodium atom becomes more positively charged, because it loses one electron to each chlorine atom. Sodium is oxidized. Each chlorine atom gains one electron from the sodium atom and becomes more negatively charged. Chlorine is reduced.

Chlorine is the oxidizing agent because it is the substance that gains or accepts electrons from sodium. The sodium metal is the reducing agent because it is losing or donating electrons to chlorine.

9. Oxidation Numbers

Question
Which species is oxidized and which species is reduced in the combustion reaction of methane and oxygen?

Approach
Examine the oxidation numbers of the elements before and after the reaction.

Solution
Step 1. Write the balanced reaction of methane and oxygen, and indicate the physical state of each species.

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

Step 2. Assign oxidation numbers to each element.

$CH_4(g)$ +	2 $O_2(g)$	$\rightarrow$	$CO_2(g)$ +	2 $H_2O(g)$
-4 +1	0		+4 -2	+1 -2

Step 3. Using oxidation numbers, determine which element gains electrons (reduction) and which element loses electrons (oxidation).

Oxidation state of carbon: -4 $\rightarrow$ +4
Oxidation state of oxygen: 0 $\rightarrow$ -2

In the reaction of methane and oxygen, carbon is oxidized and oxygen is reduced. Hydrogen is neither oxidized nor reduced, as its oxidation number remains +1 after the reaction has occurred.

10. Oxidation Numbers

Question
Which species is oxidized and which species is reduced in the reaction of potassium nitrite, potassium iodide, and sulfuric acid?

Approach
Examine the oxidation numbers of the elements before and after the reaction.

Solution
Step 1. Write the balanced reaction of potassium nitrite, potassium iodide, and sulfuric acid, and indicate the physical state of each species.

$2\ KNO_2(aq) + 2\ KI(aq) + 2\ H_2SO_4(aq) \rightarrow$
$2\ NO(aq) + 2\ K_2SO_4(aq) + 2\ H_2O(aq) + I_2(aq)$

Step 2. Assign oxidation numbers to each element.

2 KNO_2(aq) +	2 KI(aq) +	2 H_2SO_4(aq)	→
+1 +3 -2	+1 -1	+1 +6 -2	
2 NO(aq) +	2 K_2SO_4(aq) +	2 H_2O(aq) +	I_2(aq)
+2 -2	+1 +6 -2	+1 -2	0

Step 3. Using oxidation numbers, determine which element gains electrons (reduction) and which element loses electrons (oxidation).

Oxidation state of iodine: -1 → 0
Oxidation state of nitrogen: +3 → +2

In the reaction of potassium nitrite, potassium iodide, and sulfuric acid, iodine is oxidized and nitrogen is reduced. Hydrogen, potassium, sulfur, and oxygen are neither oxidized nor reduced, as their oxidation numbers remain the same after the reaction has occurred. Moreover, SO_4^{2-} stays intact, thus there is no need to consider it in terms of being oxidized or reduced.

11. Oxidation Numbers

Question
Which species is oxidized and which species is reduced in the reaction of chlorine and hydrogen sulfide?

Approach
Examine the oxidation numbers of the elements before and after the reaction.

Solution
Step 1. Write the balanced reaction of chlorine and hydrogen sulfide, and indicate the physical state of each species.

8 Cl_2(g) + 8 H_2S(aq) → S_8(s) + 16 HCl(aq)

Step 2. Assign oxidation numbers to each element.

8 Cl_2(g) +	8 H_2S(aq)	→	S_8(s) +	16 HCl(aq)
0	+1-2		0	+1 -1

Step 3. Using oxidation numbers, determine which element gains electrons (reduction) and which element loses electrons (oxidation).

Oxidation state of chlorine: 0 → -1
Oxidation state of sulfur: -2 → 0

In the reaction of chlorine and hydrogen sulfide, sulfur is oxidized and chlorine is reduced. Hydrogen is neither oxidized nor reduced, as its oxidation number remains the same after the reaction has occurred.