CH 221 Chapter Four Part II Concept Guide

1. Solubility

Why are some compounds soluble and others insoluble? In solid potassium permanganate, $KMnO_4$, the potassium ions, which have a charge of +1, are attracted to permanganate ions, which have a charge of -1. There's an electrostatic force between them that locks them into the lattice structure. When $KMnO_4$ dissolves, ionic bonds are broken as water molecules surround the ions. The partial positive charges on water molecules are attracted to the negatively charged permanganate ion.

When an ionic compound dissolves, forces are broken between ions of the ionic compounds and between the water molecules.

Water molecules that surround, or hydrate, an ion are less able to interact with other water molecules. So when an ionic compound dissolves, two different electrostatic forces are broken: the forces between positive and negative ions of the ionic compound, and the forces between water molecules that break when a water molecule hydrates an ion.

Two kinds of forces also form when an ionic compound dissolves: forces between cations and water molecules, and forces between anions and water molecules.

The relative strengths of the forces being made and broken are one factor that determines solubility. If the forces between the ions in an ionic compound are fairly weak compared to the forces between the ions and the water molecules, then the compound should be fairly soluble because stronger forces occur when the compound dissolves. However, if the forces within the lattice are stronger than the forces between the ions and the water molecules, then a compound will tend to be insoluble.

The relative strength of forces is only one of several different aspects of solubility. Another relates to the degree of order or disorder involved in the dissolution process. Entry of ions into solution causes an ordering of water molecules and ions. This effect deters solubility. Solubility is difficult to predict because the force that causes an ion to be attracted to a water molecule also causes it to be attracted to ions of the opposite charge. An ion that is strongly attracted to its neighboring ions will also be strongly attracted to water molecules.

2. Electrolytes

Compounds that form ions in aqueous solution are called electrolytes. The name reflects the fact that aqueous solutions containing ions can conduct electricity through the movement of those ions. Ions can be formed through dissolution of ionic compounds or by other reactions, including reactions of neutral molecular compounds with water molecules.

A strong electrolyte is a compound that ionizes completely when it dissolves. That is, nearly all of the compound forms cations and anions. Aqueous solutions of strong electrolytes are good conductors of electricity.

- **Strong electrolyte:** a compound that breaks up completely to form ions in aqueous solution.

Hydrogen chloride, HCl, is an example of a strong electrolyte. When hydrogen chloride is in aqueous solution, the molecules break up into hydrogen cations and chloride anions. Each molecule of HCl is neutral. When it dissolves, chlorine gains an electron to become a chloride ion with a negative charge. The hydrogen atom has lost its bonding electron, so it is now a hydrogen ion, which is really just a proton. Because this dissociation process happens to almost every hydrogen chloride molecule in aqueous solution, hydrogen chloride is called a strong electrolyte. Weak electrolytes conduct electricity only moderately well because only a small percentage of the compound reacts to form ions in aqueous solution.

Acetic acid, the main dissolved component in vinegar, is a weak electrolyte. In water, acetic acid partially dissociates into a proton and a polyatomic acetate anion. But the molecule doesn't remain dissociated for long. The acetate anion eventually links up with another proton. At the same time, other acetic acid molecules are dissociating into ions and then re-forming. About 0.42 percent of the acetic acid molecules in vinegar are dissociated.

A nonelectrolyte forms no ions and therefore cannot conduct an electric charge.

- **Nonelectrolyte:** a compound that dissolves but does not ionize in aqueous solution.

For example, the compound ethanol, which is the alcohol in alcoholic beverages, is a nonelectrolyte. Ethanol dissolves easily in water because, like water, it is a polar molecule.

The -OH group has a partial negative charge on the oxygen atom and a partial positive charge on the hydrogen atom. When ethanol dissolves in water, water molecules surround the polar ethanol molecules just as they surround ions, but aqueous ethanol is not ionic, and so cannot conduct electricity.

- **Examples of Electrolytes**

Strong Electrolytes	Weak Electrolytes	Nonelectrolytes
NaCl	HF	H_2O
$MgBr_2$	HNO_2	CH_3CH_2OH (ethanol)
HCl	NH_3	CH_3COCH_3 (acetone)
NaOH	CH_3CO_2H (acetic acid)	$C_2H_4(OH)_2$

3. 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:

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

4. 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

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

5. 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.

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

6. 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.

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

7. Polyprotic Acids

Question

What are the three ionization steps for the polyprotic acid, H_3PO_4, reacting with H_2O?

Approach

The ionizations of polyprotic acids occur stepwise, one at a time. Phosphoric acid has three acidic protons, thus there are three ionization steps.

Solution

Step 1. Write out the reaction for phosphoric acid and water.

$H_3PO_4(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + H_2PO_4^-(aq)$

Step 2. Now, write the reaction for second ionization step, i.e., the removal of a total of 2 protons from phosphoric acid. Use $H_2PO_4^-$ from the first step and react with water.

$H_2PO_4^-(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + HPO_4^{2-}(aq)$

Step 3. Finally, the third ionization step is the reaction of HPO_4^{2-} with water. This reaction yields the tribasic form of phosphate. There are no additional ionization steps because the product of this step, PO_4^{3-}, does not have any protons.

$HPO_4^{2-}(aq) + H_2O(aq) \rightarrow H_3O^+(aq) + PO_4^{3-}(aq)$

8. Acids Without Hydrogen Atoms

Question

What is the reaction for the synthesis of HNO_3?

Approach

Oxides of the non-metals are usually acids. Nonmetal oxides react with water molecules to increase the acidity of the solution. The reaction, therefore, should be that of NO_2 and H_2O. $H_2(g)$ is also generated through the oxidation of NO_2.

Solution

Step 1. Write the synthesis reaction of HNO_3 using NO_2 and H_2O as reactants.

$NO_2(aq) + H_2O(aq) \rightarrow HNO_3(aq) + {}^{1}/_{2}\, H_2(aq)$

9. Acid-base Synthesis Reaction

Question

What is an appropriate example of an acid-base reaction for the synthesis of K_2SO_4?

Solution

Step 1. Choose a strong acid to give SO_4^{2-} ions.

$H_2SO_4(aq)$

Step 2. Choose a strong base to give K^+ ions.

$KOH(aq)$

Step 3. Write the balanced reaction for the synthesis for the dissolved salt, K_2SO_4.

$H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(aq)$

Step 4. The net ionic equation for the reaction is:

$H^+(aq) + OH^-(aq) \rightarrow H_2O(aq)$

10. Acids without Hydrogen Atoms

An aqueous solution of dissolved carbon dioxide gas is acidic: it has an H^+ concentration greater than that of pure water. How can CO_2 act as an acid if it has no hydrogen to donate?

Carbon dioxide reacts with the hydroxide ions in solution. When CO_2 and OH^- come together, the oxygen of the OH^- bonds to carbon, forming hydrogen carbonate ion, also known as bicarbonate ion. This leaves an excess of hydronium ions in solution, which agrees with the Arrhenius definition of an acid.

When dissolved in water, oxides of the nonmetals yield acids, and are sometimes called acid anhydrides. Nonmetal oxides can also react with intact water molecules, rather than with free hydroxide ions, to increase the acidity of the solution. For example, when coal is burned, sulfur compounds can be released into the atmosphere. There they react to form SO_3. which then reacts with water to make sulfuric acid:

$$SO_3(aq) + H_2O(l) \rightarrow H_2SO_4\ (aq)$$

The reaction takes place in droplets of water in the atmosphere, and the product, sulfuric acid, falls to Earth dissolved in raindrops. Burning coal without trapping the sulfur generated during combustion can increase the acidity of raindrops.

Although most oxides of non-metals yield acidic aqueous solutions through reactions such as those of CO_2 and SO_3, these compounds are not usually listed in tables of acids**.** It's important to recognize that they act as acids in aqueous solution.

11. Metal Oxides are Bases

We know that oxides of nonmetal elements act as acids. Oxides of the metallic elements act as bases and are sometimes called basic anhydrides. A metal oxide can react in water to make a basic solution.

For example, solid calcium oxide reacts with water to form calcium hydroxide:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

Some of the product, $Ca(OH)_2$, dissociates, adding hydroxide ions to solution:

$$Ca(OH)_2(s) \rightarrow Ca^{2+}\ (aq) + 2\ OH^-\ (aq)$$

12. 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.

13. 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.

14. 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.

15. 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.

16. Molarity

Problem
A student prepared a solution by dissolving 1.455 g of potassium nitrate in enough water to make 20.00 mL of solution. Calculate the molarity of the solution.

Approach
Molarity is calculated by dividing the number of moles of solute by the volume of solution (in liters).

Solution
Step 1. Calculate the number of moles of solute, KNO_3.
Number of moles of KNO_3 = (1.455 g KNO_3)(1 mol / 101.11 g KNO_3) = 0.01439 moles KNO_3

Step 2. Calculate the molarity of the solution.
Molarity = 0.01439 moles KNO_3 / 0.02000 L solution = 0.7195 M KNO_3

17. Molarity

Question
How many moles of NaCl are present in 15.00 mL of a 1.60 M NaCl?

Approach
To calculate the number of moles, multiply the molarity of the solution with the volume of solution (in liters).

Solution
Number of moles NaCl = 1.60 mol/L NaCl * 0.01500 L solution = 0.0240 mol NaCl

18. Molarity

Question
What volume of a 1.35×10^{-3} M $C_6H_{12}O_6$ solution should you transfer to obtain a solution that contains 2.44×10^{-6} moles of glucose?

Approach
Use the molarity of the glucose solution to convert from units of moles to units of liters.

Solution
Volume of solution to transfer = 2.44×10^{-6} moles glucose / 1.35×10^{-3} M $C_6H_{12}O_6$
= 0.00181 L
= 1.81 mL of 1.35×10^{-3} M $C_6H_{12}O_6$

19. Dilution

Problem
Calculate the volume of 0.0111 M HCl that we should use to prepare 100 mL of a 4.23×10^{-4} M HCl solution.

Approach
Use the following equation to calculate the final volume of the solution. The other variables, $M_{initial}$, $V_{initial}$, and M_{final} are all known values.

$$M_{initial} \times V_{initial} = M_{final} \times V_{final}$$

Solution

$$M_{initial} \times V_{initial} = M_{final} \times V_{final}$$
$$(0.0111 \text{ M HCl})(V_{initial}) = (4.23 \times 10^{-4} \text{ M HCl})(100 \text{ mL})$$
$$V_{initial} = 3.81 \text{ mL}$$

3.81 mL of a 0.0111 M HCl(aq) solution are needed to prepare 100 mL of a 4.23×10^{-4} M HCl(aq) solution.

20. Titrations

Question
If 20.00 mL of a solution of oxalic acid, $H_2C_2O_4$, was titrated with 0.500 M NaOH(aq) and the endpoint was reached when 30.0 mL of the solution of base had been added, what is the molarity of the oxalic acid solution?

Approach
First, we need to write the balanced chemical equation. Calculate the number of moles of NaOH that were added. Then calculate the number of moles of $H_2C_2O_4$ consumed using the mole ratios indicated in the balanced equation.

Solution
Step 1. Write the balanced chemical equation.

$$H_2C_2O_4(aq) + 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(l)$$

Step 2. Calculate the number of moles of NaOH added.

0.03000 L NaOH x 0.500 mo/L NaOH = 0.0150 mol NaOH

Step 3. Calculate the number of moles of $H_2C_2O_4$.
The stoichiometric relation we need is 2 mol NaOH: 1 mol $H_2C_2O_4$.
Number of moles of $H_2C_2O_4$
= 0.0150 mol NaOH x (1 mol $H_2C_2O_4$ / 2 mol NaOH)
= 0.00750 mol $H_2C_2O_4$

Step 4. The molarity of the oxalic acid solution can be calculated by dividing the moles of $H_2C_2O_4$ by the volume of base added in liters.

0.00750 mol $H_2C_2O_4$ / 0.0200 L solution = 0.375 M $H_2C_2O_4$

21. Titrations

Question
A student prepared a sample of aqueous HCl that contained 0.45 g of HCl in 250 mL of solution. This solution was used to titrate 20.0 mL of a solution of $Ca(OH)_2$ and the equivalence point was reached when 14.4 mL of acid solution had been added. What was the molarity of the $Ca(OH)_2$ solution?

Approach
First, we need to write the balanced chemical equation. Then the molarity of the HCl solution must be calculated. From this value, the number of moles of HCl can be determined, and the number of moles of $Ca(OH)_2$ can be calculated using a mole ratio. Then, the molarity of the $Ca(OH)_2$ solution can be calculated.

Solution
Step 1. Write the balanced chemical equation.
$2\ HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$

Step 2. Calculate the molarity of HCl.
(0.45 g HCl)(1 mol HCl / 36.46 g HCl) / 0.250 L solution = 0.049 mol/L HCl

Step 3. Calculate the number of moles of HCl.
0.049 mol/L HCl x 0.0144 L HCl = 7.1 x 10^{-4} mol HCl

Step 4. Calculate the number of moles of $Ca(OH)_2$.
The stoichiometric relation we need is 2 mol HCl : 1 mol $Ca(OH)_2$.
Number of moles of $Ca(OH)_2$
= 7.1 x 10^{-4} mol HCl * (1 mol $Ca(OH)_2$ / 2 mol HCl)
= 3.6 x 10^{-4} mol $Ca(OH)_2$.

Step 5. The molarity of the $Ca(OH)_2$ solution can be calculated by dividing the moles of $Ca(OH)_2$ by the volume of $Ca(OH)_2$ solution, in liters.
3.6 x 10^{-4} mol $Ca(OH)_2$ / 0.020 L $Ca(OH)_2$ = 0.018 mol/L $Ca(OH)_2$

22. 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

23. 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

24. Synthesis Using Acid-Base Reactions

Compounds can be synthesized using acid-base reactions. Consider the compound calcium chloride, $CaCl_2$. Calcium chloride is soluble in water. Thus, given an aqueous $CaCl_2$ solution, we could isolate the $CaCl_2$ by evaporating the water.

When a strong acid and a strong base react, they form a dissolved salt and water. To form the salt calcium chloride, we need to react a strong acid that will donate a chloride ion to water with a strong base that will donate a calcium ion to water.

One calcium chloride producing reaction uses hydrochloric acid, HCl, for the acid and calcium hydroxide, $Ca(OH)_2$, for the base:

$Ca(OH)_2\ (aq) + 2HCl(aq) \rightarrow CaCl_2\ (aq) + 2H_2O(l)$

The water can be evaporated to isolate the pure calcium chloride compound.

Another reaction we could use to generate calcium chloride begins with HCl and calcium carbonate, $CaCO_3$, as reactants. Metal carbonates react with strong acids to generate a salt, water, and carbon dioxide:

$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

The CO_2 gas escapes from solution, leaving water and calcium chloride. Again, evaporating the water could isolate the pure calcium chloride compound.