8

BIG IDEA 3: STOICHIOMETRY

Big Idea 3

Changes in matter involve the rearrangement and/or reorganization of atoms and/or the transfer of electrons.

In this chapter you learn about standard analytical chemistry methods that allow determination of percent composition, molarity, empirical formula, and the quantities of materials consumed and produced in chemical reactions (stoichiometry). You may need to review the mole conversions and molar mass calculations in Chapter 2.

You should be able to

- Describe how to prepare solutions and to use titration data to determine the molarity (concentration) of solutions.
- Write balanced chemical equations and predict the amount of product formed from a given mass of reactant or the amount of reactant required to produce a desired amount of product.
- Identify limiting reactants, and calculate the amount of product formed when given the amounts of all of the reactants present.
- Calculate the percent yield of a reaction.
- For reactions in solution, given the molarity and the volume of the reactants, calculate the amount of product produced or the amount of reactant required to react.
- Use data from a titration, or describe an experiment using titration, to find the concentration of a solution.

MOLARITY

(*Chemistry* 8th ed. pages 136–144/9th ed. pages 145–153)

The number of moles of solute in 1 liter of solution is a measure of its molarity or solution concentration.

EXAMPLE: Prepare 2.00 L of 0.250 *M* NaOH from solid NaOH.

SOLUTION:

$$2.00\ \text{L} \times \frac{0.250\ \text{mol NaOH}}{\text{L}} \times \frac{40.00\ \text{g NaOH}}{1\ \text{mol}} = \mathbf{20.0\ g\ NaOH}$$

Place 20.0 g NaOH in a 2-L volumetric flask; add water to dissolve the NaOH, and fill to the mark with water, mixing several times along the way.

Another way to prepare a solution of a specific molarity is to dilute a more concentrated solution, using the relationship:

$$\boxed{M_1V_1 = M_2V_2}$$

EXAMPLE: Prepare 2.00 L of 0.250 *M* NaOH from 1.00 *M* NaOH.

SOLUTION:

$$1.00\ M\ \ V_1 = 0.250\ M \times\ 2.00\ \text{L}\quad V_1 = \mathbf{0.500\ L}$$

Add 500. mL of 1.00 *M* NaOH stock solution to a 2-L volumetric flask; add deionized water in several increments with mixing until the flask is filled to the mark on the neck of the flask.

DETERMINATION OF EMPIRICAL FORMULA BY COMBUSTION ANALYSIS

(*Chemistry* 8th ed. pages 90–94/9th ed. pages 96–100)

You learned about the empirical formula in Chapter 2. A standard way of determining this is by combustion analysis. In this procedure, a small sample of the substance is burned in oxygen and the amount of resulting products, which may be solid or gaseous oxides, measured.

(The following problem involves two steps. First, the percent composition of the compound is determined from combustion data. Second, the empirical formula of the compound is determined from the percent composition.)

EXAMPLE: A compound contains only carbon, hydrogen, and oxygen. Combustion of 10.68 mg of the compound yields 16.01 mg CO_2 and 4.37 mg of H_2O. What is the percent composition of the compound?

SOLUTION: Assume that all the carbon in the compound is converted to CO_2 and determine the mass of carbon present in the 10.68-mg sample. The % C in CO_2 is calculated as described in percent composition, above. It can then be multiplied times the mass of CO_2 to find the mass of C.

$$16.01\ \text{mg}\ CO_2 \times \frac{12.01\ \text{g C}}{44.01\ \text{g}\ CO_2} = 4.369\ \text{mg C}$$

The mass of C can be divided by the mass of the original compound to find the mass percent of C in this compound:

$$\frac{4.369\ \text{mg C}}{10.68\ \text{mg}} \times 100\% = 40.91\%\ \text{C}$$

The same procedure can be used to find the mass percent of hydrogen in the unknown compound. We assume that all the hydrogen present in 10.68 mg of compound was converted to H_2O.

$$0.0437\ \text{g}\ H_2O \times \frac{2.016\ \text{g H}}{18.02\ \text{g}\ H_2O} = 0.489\ \text{mg H}$$

The mass percent of H in the compound is

$$\frac{0.489\ \text{mg}}{10.68\ \text{mg}} \times 100\% = 4.58\%\ \text{H}$$

The unknown compound contains only carbon, hydrogen, and oxygen. The remainder must be oxygen.

$$100.00\% - (40.91\%\ \text{C} + 4.58\%\ \text{H}) = 54.51\%\ \text{O}$$

CONVERSIONS INVOLVED IN STOICHIOMETRIC CALCULATIONS

Type of Conversion	Example	Reference
g → mol (using molar mass)	$\frac{18\ \text{g}}{1\ \text{mol}\ H_2O}$	*Chemistry* 8th ed. page 85/9th ed. page 91
mol → mol (using mole ratio – coefficients from balanced chemical equation)	$\frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O}$	*Chemistry* 8th ed. pages 102–104/9th ed. pages 108–110
mol → L (using molarity which is a solution's concentration equal to the number of mols of solute in 1 liter of solution)	$\frac{6\ \text{mol HCl}}{1\ \text{L}}$	*Chemistry* 8th ed. pages 136–144/9th ed. pages 145–153

SOLUTION STOICHIOMETRY

(*Chemistry* 8th ed. pages 156–160/9th ed. pages 165–170)

Solution stoichiometry involves calculations for reactions that occur in aqueous solution. The moles of product or reactant are determined by multiplying concentration by volume, being careful to keep units consistent.

EXAMPLE: Mercury(II) ion can be removed from water by precipitation with sodium sulfide. What volume of 0.100 *M* Na_2S must be added to 100.0 mL of a 0.150 *M* $HgCl_2$ solution to remove all the Hg^{2+} as HgS?

SOLUTION: First, write a balanced molecular equation.

$$HgCl_2 + Na_2S \rightarrow HgS + 2\,NaCl$$

Find the moles of Na_2S required from the volume and molarity of $HgCl_2$.

$$0.1000\text{ L HgCl}_2 \times \frac{0.150\text{ mol HgCl}_2}{1\text{ L}} \times \frac{1\text{ mol Na}_2\text{S}}{1\text{ mol HgCl}_2} = 0.0150\text{ mol Na}_2\text{S}$$

The volume of Na_2S solution required is:

$$0.0150\text{ mol Na}_2\text{S} \times \frac{1\text{ L}}{0.100\text{ mol Na}_2\text{S}} = 0.150\text{ L}$$

TITRATION

(*Chemistry* 8th ed. pages 713–728/9th ed. pages 727–742)

A titration can be used to determine the concentration of an unknown solution. The calculations involve stoichiometric principles. The example here is a redox titration. Acid–base titrations are discussed in Big Idea 6.

Changing the oxidation state of a substance often results in a color change that can be used to identify the endpoint of a titration. For example, the following reaction can be used to find the concentration of Fe^{2+} in solution:

$$\underset{}{8\,H^+(aq) + 5\,Fe^{2+}(aq)} + \underset{\text{purple}}{MnO_4^-(aq)} \longrightarrow 5\,Fe^{3+}(aq) + \underset{\text{colorless}}{Mn^{2+}(aq)} + 4H_2O(l)$$

Permanganate solution of known concentration is added dropwise, with a buret, to the Fe^{2+} solution. The solution will initially be colorless because the ions are reacting. When the solution stays purple, there is no Fe^{2+} remaining to react with MnO_4^-. This is the equivalence point of the reaction, when enough moles of titrant have been added to react exactly with the moles of the substance being titrated. For this reaction, at equivalence point:

$$\text{mol } MnO_4^- = 5 \times \text{mol } Fe^{2+}.$$

EXAMPLE: A 50.0-mL solution of Fe^{2+} is titrated with 0.0020 *M* $KMnO_4$. Find the concentration of Fe^{2+} if 20.0 mL of $KMnO_4$ is required.

SOLUTION: First, write the balanced equation for the reaction, which in this example is given above. Then use the volume and molarity of the $KMnO_4$ and the mole ratio from the balanced equation to calculate the moles of Fe^{2+} in solution.

$$0.0200\text{ L KMnO}_4 \times \frac{0.0020\text{ mol KMnO}_4}{1\text{ L}} \times \frac{5\text{ mol Fe}^{2+}}{1\text{ mol KMnO}_4}$$

$$= 2.0 \times 10^{-4}\text{ mol Fe}^{2+}$$

The concentration of Fe^{2+} is:

$$\frac{2.0\times10^{-4}\ \text{mol Fe}^{2+}}{0.050\ \text{L}} = 4.0\times10^{-3}\ M\ \text{Fe}^{2+}.$$

Remember to use the initial volume of solution when determining the molarity!

LIMITING REACTANT

(*Chemistry* 8th ed. pages 107–113/9th ed. pages 114–121)

In a limiting reactant problem, the amounts of all reactants are given and the amount of product is to be determined. The limiting reactant is completely consumed when the reaction goes to completion. It determines how much product is formed.

EXAMPLE: What mass of precipitate can be produced when 50.0 mL of 0.200 M $Al(NO_3)_3$ is added to 200.0 mL of 0.100 M KOH?

SOLUTION:

Always begin with a balanced molecular equation.

$$\text{Al(NO}_3)_3(aq) + 3\ \text{KOH}(aq) \rightarrow \text{Al(OH)}_3(s) + 3\ \text{KNO}_3(aq)$$

For each reactant, determine the amount of a designated product produced (in mol or g).

$$0.0500\ \text{L} \times \frac{0.200\ \text{mol Al(NO}_3)_3}{\text{L}} \times \frac{1\ \text{mol Al(OH)}_3}{1\ \text{mol Al(NO}_3)_3}$$

$$= 0.0100\ \text{mol Al(OH)}_3$$

$$0.2000\ \text{L} \times \frac{0.100\ \text{mol KOH}}{\text{L}} \times \frac{1\ \text{mol Al(OH)}_3}{3\ \text{mol KOH}}$$

$$= 0.00667\ \text{mol Al(OH)}_3$$

The limiting reactant produces the least amount of product. The limiting reactant in this case is KOH. It produces only 0.00667 mol of product which is less than 0.0100 mol.

The mass of precipitate produced is:

$$0.00667\ \text{mol Al(OH)}_3 \times \frac{78.00\ \text{g Al(OH)}_3}{1\ \text{mol Al(OH)}_3} = 0.520\ \text{g Al(OH)}_3$$

PERCENT YIELD

(*Chemistry* 8th ed. pages 113–115/9th ed. pages 121–123)

The *percent yield* of a reaction is the actual yield of a product as a percentage of the theoretical yield. The *actual yield* is the amount of product obtained in an experiment. The *theoretical yield* is the amount of product calculated from the amounts of reactants used. The percent yield is generally less than 100% because the reaction doesn't go to completion, because of side reactions that occur without generating desired product, and because of difficulties in collecting the entire product. Yields greater than 100% are evidence of some contamination or error in calculation or measurement.

EXAMPLE: If the reaction above has an 85.3% yield of precipitate, how much aluminum hydroxide is produced?

SOLUTION:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

$$\text{Actual Yield} = \frac{85.3\% \times 0.520 \text{ g}}{100\%} = 0.444 \text{ g } Al(OH)_3$$

MULTIPLE-CHOICE QUESTIONS

No calculators are to be used in this section.

1. What volume of 10.0 *M* NaOH must be used to prepare 500. mL of a 2.50 *M* solution?
(A) 125. mL
(B) 200. mL
(C) 250. mL
(D) 12.5 mL

2. Analysis of a sample of an oxide of chromium is reported as 26 g of chromium and 12 g of oxygen. From these data determine the empirical formula of this compound.
(A) CrO
(B) Cr_2O_3
(C) CrO_3
(D) Cr_4O_6

3. Aluminum reacts with sulfuric acid, H_2SO_4, to form aluminum sulfate, $Al_2(SO_4)_3$, and hydrogen gas. Give the sum of all coefficients (all reactants and all products) for this balanced chemical expression.
(A) 5
(B) 6
(C) 9
(D) 12

4. Methane reacts with oxygen to form carbon dioxide and water as shown in the following chemical equation: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$. If 72 g of water form, how much methane must have reacted?
(A) 16 g
(B) 32 g
(C) 36 g
(D) 84 g

5. Hydrogen reacts with oxygen to form only water. If 16 grams of hydrogen is mixed with 16 grams of oxygen, how much water can form?
(A) 0.50 grams
(B) 8.0 grams
(C) 18 grams
(D) 72 grams

6. How many moles of barium sulfide, BaS, will form when 60.0 mL of 1.0 *M* $Ba(NO_3)_2$ are mixed with 25.0 mL of 0.80 *M* K_2S solution to form barium sulfide solid?
(A) 0.020 mol
(B) 0.040 mol
(C) 0.060 mol
(D) 0.10 mol

7. Consider the neutralization reaction $Be(OH)_2 + 2\ HCl \rightarrow BeCl_2 + 2\ H_2O$.
What volume of 5.00 *M* HCl is required to react completely with 4.30 g of $Be(OH)_2$? (Molar mass of $Be(OH)_2$ = 43.0 g/mol.)
(A) 10.0 mL
(B) 30.0 mL
(C) 40.0 mL
(D) 50.0 mL

8. Sucrose, $C_{12}H_{22}O_{11}$, has a molar mass of 342 g/mol. How many atoms of carbon are there in 684 g of sucrose?
(A) 6.02×10^{23} atoms
(B) 1.45×10^{25} atoms
(C) 1.20×10^{24} atoms
(D) 3.01×10^{23} atoms

9. Nitric acid reacts with silver metal: $4\ HNO_3 + 3\ Ag \rightarrow NO + 2\ H_2O + 3\ AgNO_3$. Calculate the number of grams of NO formed when 10.8 g of Ag reacts with 12.6 g of HNO_3.
(A) 0.999 g
(B) 9.00 g
(C) 12.0 g
(D) 18.0 g

10. How many grams of chromium are in 58.5 g of $K_2Cr_2O_7$? (Molar mass = 294 g/mol.)
(A) 10.4 g of Cr
(B) 15.6 g of Cr
(C) 20.8 g of Cr
(D) 208 g of Cr

11. If 7.0 moles of sulfur atoms and 10 moles of oxygen molecules are combined to form the maximum amount of sulfur trioxide, how many moles of which reactant remain unused at the end?

$$2\ S + 3\ O_2 \rightarrow 2\ SO_3$$

(A) 0.25 mol O_2
(B) 0.33 mol O_2
(C) 0.33 mol S
(D) 0.67 mol S

12. What is the percent mass nitrogen in sodium cyanide, NaCN?
(A) 14.00%
(B) 24.41%
(C) 28.57%
(D) 49.10%

13. A compound of nitrogen and oxygen is 63.64% by mass nitrogen. What is the empirical formula of this compound?
(A) NO
(B) NO_2
(C) NO_3
(D) N_2O

14. When 16 g of methane (CH_4) and 32 g of oxygen (O_2) reacted to produce carbon dioxide and water, 11 g of carbon dioxide was produced. Calculate the percent yield of carbon dioxide in this reaction.
(A) 5.0%
(B) 10%
(C) 25%
(D) 50%

15. Zinc sulfide reacts with oxygen to yield zinc oxide and sulfur dioxide as follows:

$$2\ ZnS(s) + 3\ O_2(g) \rightarrow 2\ ZnO(s) + 2\ SO_2(g)$$

How many moles of ZnO are produced when 32 g of oxygen is allowed to react with an excess of ZnS?
(A) 0.67
(B) 1.0
(C) 1.3
(D) 2.0

FREE-RESPONSE QUESTIONS

Calculators may be used for this section.

1. A 10.0-g sample of an oxide of copper is heated in a stream of pure hydrogen, forming 1.26 g of water.
(a) Determine the percentage of copper in the compound.
(b) Determine the empirical formula of the copper oxide. Name it.

2. Nonprescription antacids may contain MgO, $Mg(OH)_2$, or $Al(OH)_3$.
   (a) Write a balanced equation for the neutralization of hydrochloric acid by each of these substances.
   (b) Which of these substances will neutralize the greatest amount of 0.1 *M* HCl per gram?

Answers

MULTIPLE-CHOICE QUESTIONS

1. **A** Rearrange the relationship $M_1V_1 = M_2V_2$ to solve for V_1. Then $V_1 = (2.5\ M/10.0\ M) \times 500.$ mL (*Chemistry* 8th ed. pages 145–153/9th ed. pages 153–162). LO 3.4

2. **B** You have 0.5 mol of Cr (26 g × 1 mol/52 g) and 0.75 mol of O (12 g × 1 mol/16 g). The ratio of mol Cr: mol O is 1/2:3/4 = (1/2 × 4/3) = 2/3 or 2 mol Cr: 3 mol O. While answer D gives the same ratio, it is not the smallest whole number ratio (*Chemistry* 8th ed. pages 93–97/9th ed. pages 99–103). LO 1.3

3. **C** The equation is $2\ Al + 3\ H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3\ H_2$. Note that hydrogen is diatomic. Be sure to count the coefficients for the whole equation; it is understood that the coefficient in front of the aluminum sulfate is one (*Chemistry* 8th ed. pages 97–103/9th ed. pages 103–109). LO 3.2

4. **B** The equation indicates a mole ratio of water to methane of 2:1. The 72 grams of water is 4.0 moles of water (72 g/18 g/mol = 4.0 mol), so 2.0 moles of methane are required. Two moles of methane have a mass of 32 grams (16 g/mol × 2) (*Chemistry* 8th ed. pages 102–107/9th ed. pages 108–114). LO 3.2

5. **C** Write an equation for the reaction: $2\ H_2 + O_2 \rightarrow 2\ H_2O$. The mole ratio can then be seen as 2 mol H_2 for 1 mol O_2. Determine the number of moles of each reactant: 16 g H_2 × 1 mol H_2/2.0 g = 8.0 mol H_2 and 16 g O_2 × 1 mol O_2/32 g = 0.50 mol O_2.

   Since O_2 is the limiting reagent, it determines the amount of water formed. From the mole ratio in the equation, it follows that 0.50 mol O_2 forms two times that amount of water or 1 mol of water. The molar mass of water is 18 g/mol (*Chemistry* 8th ed. pages 107–115/9th ed. pages 114–123). LO 3.4

6. **A** The equation, $Ba(NO_3)_2 + K_2S \rightarrow BaS + 2\ KNO_3$, indicates that with 0.060 mole $Ba(NO_3)_2$ and 0.020 mole K_2S, the potassium sulfide is the limiting reactant (they react in a one-to-one ratio). The equation also shows that for one mole of potassium sulfide, one mole of barium sulfide forms, hence 0.020 mol of potassium sulfide will allow for the formation of 0.020 mol of the solid BaS (*Chemistry* 8th ed. pages 144–149, 151–152/9th ed. pages 153–157, 160–161). LO 3.4

7. C The mass of 4.30 g of $Be(OH)_2$ represents 0.100 mol of this base; the equation indicates that twice that amount (0.200 mol) of acid is needed for neutralization to be complete. Since both the concentration and number of moles needed are known for the acid, the volume of the acid can be calculated (0.200 mol/5.00 mol/L = 0.0400 L = 40.0 mL) (*Chemistry* 8th ed. pages 158–160/9th ed. pages 167–170). LO 3.4

8. B It takes 342 g to make one mole of the compound and you have twice that amount. Therefore you have 2.00 mole of this substance (684 g/342 g/mol). Each mole of the substance contains 12 moles of carbon, for a total of 24 moles of carbon. Multiply 24 times Avogadro's number to obtain the atoms of carbon (*Chemistry* 8th ed. pages 81–87/9th ed. pages 85–92). LO 1.4

9. A Find the number of moles of each reactant to determine which is in excess and which is the limiting reactant: 10.8 g Ag/108 g/mol = 0.100 mol of Ag; 12.6 g HNO_3/63.0 g/mol = 0.400 mol of HNO_3. Since the Ag/HNO_3 ratio (from the equation) is 3/4, Ag is the limiting reactant. 0.100 mol of Ag will form one-third as much NO (0.100 × 1 NO/3 Ag) or 0.0333 mol of NO. Finally, 0.0333 mol NO × 30.0 g NO/mol = 0.999 g NO (*Chemistry* 8th ed. pages 107–115/9th ed. pages 114–123). LO 3.4

10. C First, determine the number of moles of the salt: 58.5 g/294 g/mol = 0.200 mol $K_2Cr_2O_7$. Since Cr has a molar mass of 52.0 g/mol and there are two Cr per each unit of $K_2Cr_2O_7$, 0.200 mol × 2 Cr/$K_2Cr_2O_7$ × 52.0 g/mol = 20.8 g of Cr (*Chemistry* 8th ed. pages 82–84/9th ed. pages 86–90). LO 1.4

11. C According to the balanced equation, for complete reaction the required ratio of moles of oxygen to sulfur is 3/2 or 1.5. In the problem we have a ratio of 10/7 or 1.43. Since this ratio is less than the required ratio, this means the oxygen is limiting; 10 moles of oxygen require 10 × (2/3) mol S = 6.67 mol S used. Therefore, 7.0 – 6.67 = 0.33 mol S remains unused (*Chemistry* 8th ed. pages 107–113/9th ed. pages 114–121). LO 3.4

12. C For NaCN the molar mass is 22.99 + 12.01 + 14.01 = 49.01 g/mol. % mass of N = (14.01/49.01) × 100 = 28.57%. Since you won't have a calculator, a quick estimate would be 15/50 = 30%, which is closest to 28.57% (*Chemistry* 8th ed. pages 88–90/9th ed. pages 94–96). LO 1.2

13. D The % oxygen is 100 – 63.64 = 36.36%.

The steps to solving this problem are detailed below:

	N	O
Ratio by mass	63.64	36.36
Divide by molar masses	63.64/14.01	36.36/16.00
Mole ratio	4.54	2.27
Divide by smaller number	2	1

(*Chemistry* 8th ed. pages 90–97/9th ed. pages 96–103). LO 1.2

14. D Start with the balanced equation: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

16 grams of methane = 1.0 mole

32 grams of oxygen = 1.0 mole. This is the limiting reactant, since by the equation 1 mole of methane requires 2 moles of oxygen, and since there is only 1 mole of oxygen, it will be used up first leaving 0.5 mole of methane unreacted.

Thus, if one mole of oxygen is used up, according to the equation, 0.5 mol CO_2 would be produced. The molar mass of CO_2 is 44.0 g/mol, so the expected yield would be 22 g. The actual yield was 11 g so (11 g/22 g) × 100 = 50.% (*Chemistry* 8th ed. pages 107–113/9th ed. pages 114–121). LO 3.4

15. A According to the equation, 1.0 mol O_2 produces 1.0 × (2/3) mol ZnO. Thus 32 grams of oxygen = 1.0 mol, 2/3 or 0.67 mol ZnO would be produced (*Chemistry* 8th ed. pages 102–107/9th ed. pages 108–114). LO 3.4

FREE-RESPONSE QUESTIONS

1. (a) 88% copper

$$1.26\text{ g H}_2\text{O} \times \frac{1\text{ mol}}{18.0\text{ g H}_2\text{O}} = 0.0700\text{ mol H}_2\text{O} \times \frac{1\text{ mol O}}{1\text{ mol H}_2\text{O}}$$
$$= 0.0700\text{ mol O atoms}$$

0.0700 mol × 16.0 g/mol = 1.12 g of oxygen in the oxide

10.0 g – 1.12 g = 8.8 g of copper; 8.8 g/10.0 g × 100% = 88% copper

(b) Cu_2O copper(I) oxide

Consider the mole ratio of Cu to O:

[8.8 g/63.55 g/mol = 0.14 mol Cu]

$$\frac{0.14\text{ mol Cu}}{0.070\text{ mol O}} = \frac{2\text{ mol Cu}}{1\text{ mol O}}$$ Cu_2O is the empirical formula.

This is called copper(I) oxide since the Cu has a charge or oxidation state of +1 in this compound (*Chemistry* 8th ed. pages 87–91, 93–97, 102–107/9th ed. pages 93–97, 99–103, 108–114). LO 1.4

2. (a) $MgO(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$

$Mg(OH)_2(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + 2\ H_2O(l)$

$Al(OH)_3(s) + 3\ HCl(aq) \rightarrow AlCl_3(aq) + 3\ H_2O(l)$

(*Chemistry* 8th ed. pages 154–155/9th ed. pages 163–164). LO 3.7

(b)

	molar mass	mol HCl/mol	mol HCl/g
MgO	40.31 g	2/1	4.96×10^{-2}
$Mg(OH)_2$	58.33 g	2/1	3.43×10^{-2}
$Al(OH)_3$	78.00 g	3/1	3.84×10^{-2}

Of the three, MgO neutralizes the most acid per gram.

(*Chemistry* 8th ed. pages 85–87/9th ed. pages 90–92). LO 3.4