CHY 113 | Stoichiometry & Limiting Reactants — ANSWER KEY

Guided activity • Work with classmates or independently • Show calculations and explain your reasoning

The core activity (Models 1–4) is designed for approximately 45 minutes. Additional practice follows.

Model 1 — Start with the Mole Ratio

$$\mathbf{2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)}$$

A container initially contains 6 molecules of H_2 and 2 molecules of O_2. Assume the reaction goes to completion.

1. How many H_2 molecules are needed to react with both O_2 molecules? How many H_2O molecules form?

2 O_2 molecules require 4 H_2 molecules (2 H_2 per O_2). They produce 4 H_2O molecules.

2. Which reactant is used up first? How many molecules of the other reactant remain?

O_2 is consumed first (limiting reactant). H_2 remaining = 6 − 4 = 2 molecules; O_2 remaining = 0.

3. Now suppose the container holds 6.00 mol H_2 and 2.00 mol O_2. Give the moles of water formed and the moles of each reactant remaining.

2.00 mol O_2 × (2 mol H_2O / 1 mol O_2) = 4.00 mol H_2O. H_2 used = 4.00 mol; H_2 remaining = 6.00 − 4.00 = 2.00 mol; O_2 remaining = 0.

4. A student argues that H_2 must always be limiting because its coefficient is 2. Explain what the student overlooked.

The limiting reactant depends on the amounts available relative to the 2:1 requirement—not on which reactant has the larger coefficient. With sufficient H_2, O_2 can be limiting.

Model 2 — Review: One Reactant in Excess

$$\mathbf{CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)}$$

A student reacts 2.22 g $CaCl_2$ with more than enough Na_2CO_3.

5. What does “more than enough” tell you? Can you determine the initial mass of Na_2CO_3?

Na_2CO_3 is present in excess: there is enough to consume all the $CaCl_2$. The exact initial mass of Na_2CO_3 cannot be determined from this information.

6. Calculate the theoretical mass of $CaCO_3$. Show grams → moles → moles → grams.

$M(CaCl_2)$ = 40.08 + 2(35.45) = 110.98 g/mol; $M(CaCO_3)$ = 40.08 + 12.01 + 3(16.00) = 100.09 g/mol.
2.22 g $CaCl_2$ × (1 mol $CaCl_2$ / 110.98 g $CaCl_2$) × (1 mol $CaCO_3$ / 1 mol $CaCl_2$) × (100.09 g $CaCO_3$ / 1 mol $CaCO_3$) = 2.00 g $CaCO_3$.

7. Which conversion used molar mass? Which used coefficients from the balanced equation? Explain the distinction.

Grams $CaCl_2$ → moles $CaCl_2$ and moles $CaCO_3$ → grams $CaCO_3$ use the respective molar masses. Moles $CaCl_2$ → moles $CaCO_3$ uses the 1:1 coefficients. Molar mass relates mass and moles of the SAME substance; coefficients relate moles of DIFFERENT substances.

Model 3 — When Both Reactant Amounts Are Known

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

A reaction begins with 4.00 mol Mg and 3.00 mol O_2.

Substance	Initially	After reaction
Mg	*4.00 mol*	0 mol
O_2	3.00 mol	*1.00 mol*
MgO	0 mol	*4.00 mol*

8. If all 4.00 mol Mg reacts, how many moles of O_2 are required? Is that much available?

4.00 mol Mg × (1 mol O_2 / 2 mol Mg) = 2.00 mol O_2 required. Yes, 3.00 mol O_2 is available.

9. Identify the limiting reactant. Calculate the moles of MgO formed and excess reactant remaining.

Mg is limiting. MgO produced = 4.00 mol Mg × (2 mol MgO / 2 mol Mg) = 4.00 mol MgO. O_2 left = 3.00 − 2.00 = 1.00 mol.

10. Complete the "After reaction" column. Why can reactant remain even when the reaction is complete?

After: Mg = 0 mol; O_2 = 1.00 mol; MgO = 4.00 mol. The reaction stops when Mg is used up; O_2 remains because there is no Mg left with which to react.

11. Would adding more O_2 increase the MgO yield? Would adding more Mg increase it? Explain each.

Adding O_2 alone: no increase, because Mg is already limiting. Adding Mg: yes, initially, because the remaining O_2 can react with additional Mg (up to 2.00 more mol Mg with the 1.00 mol O_2 left).

Model 4 — Three Methods, One Answer

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

Now begin with 6.00 g Mg and 2.00 g O_2. Each method must identify the same limiting reactant.

Method A — Compare possible product yields

Often the simplest approach when you only need one product amount.

12. Assuming Mg reacts completely, calculate the mass of MgO it could produce.

M(Mg) = 24.31 g/mol; M(MgO) = 40.31 g/mol.
6.00 g Mg × (1 mol Mg / 24.31 g Mg) × (2 mol MgO / 2 mol Mg) × (40.31 g MgO / 1 mol MgO) = 9.95 g MgO (if enough O_2 were available).

13. Assuming O_2 reacts completely, calculate the mass of MgO it could produce.

$M(O_2)$ = 32.00 g/mol.
2.00 g O_2 × (1 mol O_2 / 32.00 g O_2) × (2 mol MgO / 1 mol O_2) × (40.31 g MgO / 1 mol MgO) = 5.04 g MgO (if enough Mg were available).

14. Which yield is physically possible? Identify the limiting reactant and explain why.

The smaller possible yield, 5.04 g MgO, is the actual theoretical yield. O_2 is limiting. There is not enough O_2 to produce the 9.95 g MgO predicted from all the Mg.

Method B — Compare available and required mole ratios

This is the approach used in your textbook. Some students find it intuitive; others prefer another method.

15. Convert the starting masses of Mg and O_2 into moles.

n(Mg) = 6.00 g / 24.31 g/mol = 0.247 mol Mg. n(O_2) = 2.00 g / 32.00 g/mol = 0.0625 mol O_2.

16. Calculate the available ratio mol Mg / mol O_2. What ratio is required by the balanced equation?

Available mol Mg / mol O_2 = 0.2468 / 0.0625 = 3.95 (about 3.95:1). Required by the equation = 2 mol Mg / 1 mol O_2 = 2.00 (2:1).

17. Compare the ratios to identify the limiting reactant. Explain the logic in words.

The available Mg:O_2 ratio (3.95:1) is greater than the required ratio (2:1): there is too much Mg relative to O_2. Therefore O_2 is limiting and Mg is in excess.

Method C — Before–Change–After (BCA) table

Especially useful when you need the amounts of everything, including excess reactant remaining. Use moles throughout the table.

18. First, deliberately assume Mg is the limiting reactant—even if your earlier work suggests otherwise. Convert the initial masses to moles, and complete the BCA table below under the assumption that ALL the Mg reacts.

Assume Mg is completely consumed. Initial amounts: Mg = 0.247 mol, O_2 = 0.0625 mol, MgO = 0. Changes based on Mg: -0.247 mol Mg, -0.123 mol O_2, +0.247 mol MgO. This gives a physically impossible negative amount of O_2 remaining.

First attempt — assume Mg is limiting

	Mg	O_2	MgO
Before	*0.247 mol*	*0.0625 mol*	0 mol
Change	*-0.247 mol*	*-0.123 mol*	*+0.247 mol*
After	0 mol	−0.0609 mol	*0.247 mol*

18a. Does your "After" row contain a negative amount of O_2? Can a reaction leave a negative number of moles of a substance? Explain what this tells you about your assumption.

Yes. O_2 after = 0.0625 - 0.1234 = -0.0609 mol. A physical quantity of reactant cannot be negative. Consuming all Mg would require more O_2 than exists, so Mg cannot be limiting.

18b. Now assume O_2 is limiting. Complete the corrected BCA table below. Use the coefficients to determine the changes, then find the final amounts.

If O_2 is limiting, all 0.0625 mol O_2 reacts, consuming 0.125 mol Mg and forming 0.125 mol MgO. Mg remaining = 0.2468 - 0.1250 = 0.1218 mol; O_2 remaining = 0.

Corrected BCA table — use the actual limiting reactant

	Mg	O_2	MgO

Before	*0.247 mol*	*0.0625 mol*	0 mol
Change	*−0.125 mol*	*−0.0625 mol*	+0.125 mol
After	*0.122 mol*	0 mol	0.125 mol

19. Calculate the mass of MgO produced and the mass of excess reactant remaining.

MgO: 0.125 mol × 40.31 g/mol = 5.04 g MgO. Excess Mg: 0.1218 mol × 24.31 g/mol = 2.96 g Mg remaining. Check: 5.04 + 2.96 = 8.00 g total, matching 6.00 + 2.00 g initially.

20. All three methods should agree. Which found the product amount most directly? Which made the amount remaining easiest to find?

Method A gives the single product amount most directly (compare the two possible yields). Method C makes the amounts remaining easiest to track because the BCA table accounts for all species.

Choose an Efficient Method

21. For each goal, choose a method and explain: (a) find only one product amount; (b) compare available and required mole ratios; (c) find final amounts of every reactant and product.

(a) Method A: calculate possible yields and take the smaller. (b) Method B: compare available and required mole ratios. (c) Method C: a BCA table tracks initial, consumed/formed, and remaining amounts of every species. All three methods are valid.

Additional Practice — Choose Your Strategy

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

A reaction begins with 14.0 g N_2 and 2.00 g H_2.

22. Why can't you identify the limiting reactant by comparing the starting masses directly?

The reaction requires 3 mol H_2 per 1 mol N_2. Different molar masses mean grams do not directly show whether the available mole amounts meet that ratio.

23. Identify the limiting reactant and calculate the theoretical mass of NH_3.

$M(N_2)$ = 28.02 g/mol; $M(H_2)$ = 2.016 g/mol; $M(NH_3)$ = 17.03 g/mol.
$n(N_2)$ = 14.0/28.02 = 0.500 mol; $n(H_2)$ = 2.00/2.016 = 0.992 mol. To consume all N_2 requires 1.50 mol H_2, so H_2 is limiting.
2.00 g H_2 × (1 mol H_2 / 2.016 g) × (2 mol NH_3 / 3 mol H_2) × (17.03 g NH_3 / 1 mol NH_3) = 11.3 g NH_3.

24. Calculate the mass of excess reactant remaining. Would a BCA table be useful?

N_2 consumed = 0.992 mol H_2 × (1 mol N_2 / 3 mol H_2) = 0.331 mol N_2. N_2 left = 0.500 − 0.331 = 0.169 mol = 4.73 g N_2. Yes; a BCA table organizes reactants consumed, product formed, and excess remaining.

$$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$$

A student mixes 5.00 g $CaCl_2$ and 4.00 g Na_2CO_3.

25. Identify the limiting reactant and calculate the theoretical mass of $CaCO_3$.

$M(CaCl_2)$ = 110.98 g/mol; $M(Na_2CO_3)$ = 105.99 g/mol; $M(CaCO_3)$ = 100.09 g/mol.
$n(CaCl_2)$ = 5.00/110.98 = 0.0451 mol; $n(Na_2CO_3)$ = 4.00/105.99 = 0.0377 mol. The required ratio is 1:1, so Na_2CO_3 is limiting. $CaCO_3$ = 0.0377 mol × 100.09 g/mol = 3.78 g.

26. Determine the mass of excess reactant remaining.

$CaCl_2$ consumed = 0.0377 mol × 110.98 g/mol = 4.19 g. $CaCl_2$ remaining = 5.00 − 4.19 = 0.812 g (≈ 0.81 g).

27. If 3.20 g dry $CaCO_3$ is recovered, calculate the percent yield.

Percent yield = (actual yield / theoretical yield) × 100% = (3.20 g / 3.78 g) × 100% = 84.7%.

Further Practice — Beyond the Core Activity

$$\mathbf{2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)}$$

A container holds 8.10 g Al and 14.2 g Cl_2.

28. Determine the limiting reactant and theoretical mass of $AlCl_3$.

$M(Al) = 26.98$ g/mol; $M(Cl_2) = 70.90$ g/mol; $M(AlCl_3) = 133.33$ g/mol.
$n(Al) = 8.10/26.98 = 0.300$ mol; $n(Cl_2) = 14.2/70.90 = 0.200$ mol. All 0.300 mol Al would require 0.450 mol Cl_2, so Cl_2 is limiting.
$0.200\ mol\ Cl_2 \times (2\ mol\ AlCl_3 / 3\ mol\ Cl_2) \times 133.33\ g/mol = 17.8\ g\ AlCl_3$.

29. Find the mass of excess reactant remaining. Check your results using conservation of mass.

Al consumed = $0.200\ mol\ Cl_2 \times (2\ mol\ Al / 3\ mol\ Cl_2) \times 26.98\ g/mol = 3.60$ g Al. Al remaining = 8.10 − 3.60 = 4.50 g. Mass check: initial = 8.10 + 14.2 = 22.3 g; final = 17.8 g $AlCl_3$ + 4.50 g Al = 22.3 g.

$$\mathbf{2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)}$$

A 12.3 g sample of $KClO_3$ decomposes completely.

30. Calculate the mass of O_2 produced. Why is this a stoichiometry problem but not a limiting-reactant problem?

$M(KClO_3) = 39.10 + 35.45 + 3(16.00) = 122.55$ g/mol.
$12.3\ g\ KClO_3 \times (1\ mol\ KClO_3 / 122.55\ g) \times (3\ mol\ O_2 / 2\ mol\ KClO_3) \times (32.00\ g\ O_2 / 1\ mol\ O_2) = 4.82\ g\ O_2$. Only one reactant is specified, so there is no competing reactant whose supply could limit the reaction.

Reflection

31. What extra decision is needed when both reactant amounts are given, compared with a problem stating that one reactant is in excess?

When one reactant is stated to be in excess, use the specified reactant for the stoichiometric conversion. When both amounts are given, first identify which reactant is limiting; that reactant determines the maximum product amount.

32. Complete and explain: “The balanced equation gives required _____ ratios, but the available reactant amounts determine _____.”

"The balanced equation gives required MOLE ratios, but the available reactant amounts determine WHICH REACTANT IS LIMITING." Example: 2 Mg + O_2 → 2 MgO requires 2:1, but 0.247 mol Mg and 0.0625 mol O_2 have an available ratio of about 3.95:1, so O_2 runs out first.

33. Which method do you prefer now? Describe a situation in which you would deliberately choose a different one.

Sample response: Method A is efficient when I need one product yield. I would use Method C instead if I need the amount of excess reactant remaining or final amounts of all species. Method B may be preferred when the available-versus-required mole ratio is especially clear. (Reasoned preferences may vary.)