Solutions to Practice Problems in Chapter 6 Rates of Reaction

Calculating an Average Reaction Rate
Determining Reaction Rates in Terms of Products and Reactants
(Student textbook pages 360-1)

1. In the reaction A + 2B → 3C + 4D, the initial concentration of A was 0.0415 mol/L, and after 14.7 min the concentration of A was 0.0206 mol/L. What is the average rate of consumption in moles per litre per second of reactant B?

What Is Required?
You need to determine the average rate of consumption of reactant B over the first 14.7 minutes of the reaction.

What Is Given?
You know the initial concentration of A: $[A]_{initial} = 0.0415$ mol/L
You know the final concentration of A: $[A]_{final} = 0.0206$ mol/L
You know the balanced chemical equation for the reaction between A and B.

Plan Your Strategy	Act on Your Strategy
Convert the units for time elapsed, Δt, from minutes to seconds.	$\Delta t = 14.7 \text{ min}$ $= 14.7 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 882 \text{ s}$
Use the formula $\text{average rate} = \frac{\Delta[A]}{\Delta t}$ to determine the rate of consumption of reactant A.	$\frac{\Delta A}{\Delta t} = \frac{0.0206 \text{ mol/L} - 0.0415 \text{ mol/L}}{882 \text{ s}}$ $= -2.3696 \times 10^{-5} \text{ mol/L} \cdot \text{s})$
The rate of consumption of reactants is in proportion to the mole ratio given in the balanced chemical equation. Use the mole ratio in the balanced equation A:B to determine the rate of consumption of reactant B.	$n_B : n_A = 2:1$ $\frac{\Delta[B]}{\Delta t} = 2\left(\frac{\Delta[A]}{\Delta t}\right)$ $= 2(-2.3696 \times 10^{-5} \text{ mol/L} \cdot \text{s})$ $= -4.7392 \times 10^{-5} \text{ mol/L} \cdot \text{s}$ $= -4.74 \times 10^{-5} \text{ mol/L} \cdot \text{s}$ The average rate of reaction of B is -4.74×10^{-5} mol/L•s. The value is negative because the concentration is decreasing. The rate of consumption is the opposite of the rate at which it is decreasing so the rate of consumption is $+4.74 \times 10^{-5}$ mol/L·s.

Check Your Solution
The answer is correctly expressed in the proper units and to the correct number of significant digits. The rate of consumption of reactant B is twice that of reactant A as shown by the mol ratio in the balanced equation.

2. Calculate the average rate of a reaction, given the following data:

Reaction Data

Time (s)	Concentration (mol/L)
60	5.00×10^{-2}
85	3.25×10^{-2}

What Is Required?
You need to calculate the average rate of reaction from the given data.

What Is Given?
You know the concentration at two different times.
You know the initial concentration: $c_{\text{initial}} = 5.00 \times 10^{-2}$ mol/L
You know the final concentration: $c_{\text{final}} = 3.25 \times 10^{-2}$ mol/L
You know the initial time: $t_{\text{initial}} = 60$ s
You know the final time: $t_{\text{final}} = 85$ s

Plan Your Strategy	Act on Your Strategy
Determine the time elapsed, Δt.	$\Delta t = t_{\text{final}} - t_{\text{initial}}$ $= 85\text{ s} - 60\text{ s}$ $= 25\text{ s}$
Use the formula $\text{rate} = \frac{\Delta\text{concentration}}{\Delta t}$ to calculate the average rate of reaction.	$\text{rate} = \frac{\Delta\text{concentration}}{\Delta t}$ $= \frac{c_{\text{final}} - c_{\text{initial}}}{\Delta t}$ $= \frac{3.25 \times 10^{-2}\text{ mol/L} - 5.00 \times 10^{-2}\text{ mol/L}}{25\text{ s}}$ $= -7.0 \times 10^{-4}\text{ mol/L}\cdot\text{s}$ The average rate of the reaction is -7.0×10^{-4} mol/L•s

Check Your Solution
The reaction rate is negative since the concentration is decreasing. The answer has the correct number of significant digits and seems reasonable.

3. For the reaction shown below, the average rate of formation of $CO_2(g)$ is 5.50×10^{-4} mol/s.

$$Br_2(aq) + HCOOH(aq) \rightarrow 2Br^-(aq) + 2H^+(aq) + CO_2(g)$$

a. What amount in moles of $CO_2(g)$ is formed in 5.00 min?

What Is Required?
You need to calculate the amount in moles of $CO_2(g)$ produced in 5.00 min.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the average rate of formation of $CO_2(g)$: rate = 5.50×10^{-4} mol/s
You know the time elapsed: $t = 5.00$ min

Plan Your Strategy	Act on Your Strategy
Convert the units for time elapsed, t, from minutes to seconds.	$t = 5.00 \text{ min}$ $= 5.00 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 300 \text{ s}$
Use the expression $n = \frac{\Delta\text{concentration}}{\Delta t} \times t$ to calculate the amount in moles, n, of $CO_2(g)$.	$n_{CO_2} = \frac{\Delta\text{concentration}}{\Delta t} \times t$ $= \frac{5.50 \times 10^{-4} \text{ mol}}{1 \cancel{\text{s}}} \times 300 \cancel{\text{s}}$ $= 0.165 \text{ mol}$

b. How does this compare with the amount of $Br_2(aq)$ that reacts in the same time?

What Is Required?
You need to compare the amount of $CO_2(g)$ produced to the amount of $Br_2(aq)$ consumed.

What Is Given?
You know the amount of $CO_2(g)$ produced.
You know the balanced chemical equation.

Plan Your Strategy	Act on Your Strategy
The rate of consumption of reactants is in proportion to the mole ratio given in the balanced equation. Determine the mole ratio from the balanced equation.	$\frac{n_{Br_2}}{n_{CO_2}} = \frac{1}{1}$

Determine the rate at which $Br_2(aq)$ reacts.	$\frac{\Delta[Br_2]}{\Delta t} = \frac{\Delta[CO_2]}{\Delta t}$ $Br_2(aq)$ will be consumed at the same rate as CO_2 is produced. Therefore, 0.165 mol of $Br_2(aq)$ reacts in this same time interval.

Check Your Solution

The answer has the correct number of significant digits. The equation shows a 1:1 mol ratio between $Br_2(aq)$ and $CO_2(g)$. The answer is reasonable.

4. The concentration of a reactant is 4.0×10^{-2} mol/L at $t = 2.0$ min. If the average rate of consumption of the reactant from $t = 1.5$ min to $t = 2.0$ min was 0.045 mol/L•s, what was the concentration of this reactant at $t = 1.5$ min?

What Is Required?
You need to determine the concentration of a reactant after 1.5 minutes have elapsed in the reaction.

What Is Given?
You know the average rate of reaction: rate = 0.045 mol/L•s
You know the final concentration: $c_{final} = 4.0 \times 10^{-2}$ mol/L
You know the initial time: $t_{initial} = 1.5$ min
You know the final time: $t_{final} = 2.0$ min

Plan Your Strategy	Act on Your Strategy
Convert the units for $t_{initial}$ and t_{final} from minutes to seconds.	$t_{initial} = 1.5 \text{ min}$ $= 1.5 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 90 \text{ s}$ $t_{final} = 2.0 \text{ min}$ $= 2.0 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 120 \text{ s}$
Determine the time elapsed, Δt.	$\Delta t = t_{final} - t_{initial}$ $= 120 \text{ s} - 90 \text{ s}$ $= 30 \text{ s}$
Use the formula $\text{average rate} = -\frac{\Delta\text{concentration}}{\Delta t}$ to calculate the initial concentration, $c_{initial}$. The rate is negative because the amount of reactant is decreasing.	$\text{average rate} = -\frac{\Delta\text{concentration}}{\Delta t}$ $\text{average rate} = -\frac{(c_{final} - c_{initial})}{\Delta t}$ $0.045 \text{mol/L} \bullet \text{s} = \frac{-(4.0 \times 10^{-3} \text{ mol/L} - c_{initial})}{30 \text{ s}}$ $c_{initial} = 1.346 \text{ mol/L}$ $= 1.4 \text{ mol/L}$ The initial concentration of the reactant was 1.4 mol/L.

Check Your Solution
The initial concentration is greater than the final concentration. The answer has the correct number of significant digits and seems reasonable.

5. A zinc electrode is immersed in dilute sulfuric acid at 35.0°C and the following reaction occurs:

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

The volume of $H_2(g)$ present at $t = 1.0$ min is 30.0 mL and at $t = 1.4$ min is 42.0 mL. What is the average rate of formation of $H_2(g)$ over this period of time measured in litres per second?

What Is Required?
You must determine the average rate of formation of the product, $H_2(g)$.

What Is Given?
You know the initial volume: $V_{initial} = 30.0$ mL
You know the final volume: $V_{final} = 42.0$ mL
You know the initial time: $t_{initial} = 1.0$ min
You know the final time: $t_{final} = 1.4$ min

Plan Your Strategy	Act on Your Strategy
Calculate □t and convert the units from minutes to seconds.	$\Delta t = t_{final} - t_{initial}$ $= 1.4 \text{ min} - 1.0 \text{ min}$ $= 0.4 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 24 \text{ s}$
Use the formula $\text{rate} = \frac{\Delta\text{quantity}}{\Delta t}$ to determine the average rate of formation of $H_2(g)$.	$\text{rate of formation} = \frac{\Delta\text{quantity}}{\Delta t}$ $= \frac{\Delta V}{\Delta t}$ $= \frac{V_{final} - V_{initial}}{\Delta t}$ $= \frac{(42.0 \text{ mL} - 30.0 \text{ mL})}{24 \text{ s}}$ $= \frac{12 \text{ mL}}{24 \text{ s}}$ $= 0.50 \text{ mL/s}$ $= 5.0 \times 10^{-4} \text{ L/s}$

Check Your Solution
The rate is positive since $H_2(g)$ is produced. The correct number of significant digits is shown.

6. For the reaction shown below, the instantaneous rate of formation of $Br^-(aq)$ is 0.12 mol/L•s at $t = 2.0$ min.

$$3BrO^-(aq) \rightarrow BrO_3^-(aq) + 2Br^-(aq)$$

What are the instantaneous rates of formation of $BrO_3^-(aq)$ and consumption of $BrO^-(aq)$?

What Is Required?
You need to calculate the instantaneous rate formation of $BrO_3^-(aq)$ and rate of consumption of $BrO^-(aq)$?

What Is Given?
You know the balanced chemical equation for the reaction.
You know the instantaneous rate of formation of $Br^-(aq)$ is 0.12 mol/ L•s.

Plan Your Strategy	Act on Your Strategy
The rate of formation of products and the rate of consumption of reactants is in proportion to the mole ratios given in the balanced equation. Determine the mole ratios $BrO_3^-:Br^-$ and $BrO^-:Br^-$ from the balanced equation.	$\frac{\text{rate of production of } BrO_3^-(aq)}{\text{rate of production of } Br^-(aq)} = \frac{n_{BrO_3^-}}{n_{Br^-}}$ $\frac{-\text{rate of disappearance of } BrO^-(aq)}{\text{rate of production of } Br^-(aq)} = \frac{n_{BrO^-}}{n_{Br^-}}$
Use the first ratio to calculate the instantaneous rate of formation of $BrO_3^-(aq)$.	For every 2 mol of $Br^-(aq)$ produced, 1 mol of $BrO_3^-(aq)$ is produced. $\frac{\text{rate of production of } BrO_3^-(aq)}{0.12 \text{ mol/L} \bullet \text{s}} = \frac{1 \text{ mol } BrO_3^-(aq)}{2 \text{ mol } Br^-(aq)}$ $\text{rate of production of } BrO_3^-(aq) = \frac{1}{2} \times 0.12 \text{ mol/L} \bullet \text{s}$ $= 0.06 \text{ mol/L} \bullet \text{s}$ $BrO_3^-(aq)$ is produced half as fast as $Br^-(aq)$.

Use the second ratio to calculate the instantaneous rate of consumption of BrO^-(aq).	For every 2 mol Br^-(aq) that is produced, 3 mol BrO^-(aq) is consumed. $\frac{-\text{rate of disappearance of } BrO^-(aq)}{0.12 \text{ mol/L} \bullet \text{s}} = \frac{3 \text{ mol } BrO^-(aq)}{2 \text{ mol } Br^-(aq)}$ $\text{rate of disappearance of } BrO^-(aq) = \frac{3}{2} \times 0.12 \text{ mol/L} \bullet \text{s}$ $= 0.18 \text{ mol/L} \bullet \text{s}$ BrO^-(aq) disappears 1.5 times faster than Br^-(aq) is produced.

Check Your Solution

The answers are reasonable based upon the balanced equation. The correct number of significant digits is shown.

7. The data in the table below were obtained for the following reaction:

$$4HBr(g) + O_2(g) \rightarrow 2Br_2(g) + 2H_2O(g)$$

Reaction Data

Time (s)	[HBr] (mol/L)	[Br_2] (mol/L)
0.00	0.42	0.00
50.0	0.26	?

a. What is the average rate of consumption of HBr(g) over 50.0 s?
b. What is the molar concentration of $Br_2(g)$ at t = 50.0 s?

What Is Required?
You need to calculate rate of consumption of HBr(g) and the final molar concentration of $Br_2(g)$.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the initial concentration of HBr(g): $c_{initial}$ = 0.42 mol/L
You know the final concentration of HBr(g): c_{final} = 0.26 mol/L
You know the initial concentration of $Br_2(g)$: $c_{initial}$ = 0.00 mol/L
You know the initial time: $t_{initial}$ = 0.00 s
You know the final time: t_{final} = 50.0 s

Plan Your Strategy	Act on Your Strategy
a. Use the formula: $\text{rate} = \frac{\Delta\text{concentration}}{\Delta t}$ to determine the rate of consumption of HBr(g).	**a.** rate of consumption of HBr(g): $\text{rate} = \frac{\Delta\text{concentration}}{\Delta t}$ $= \frac{c_{final} - c_{initial}}{\Delta t}$ $= \frac{0.26\text{ mol/L} - 0.42\text{ mol/L}}{50\text{ s}}$ $= -0.0032\text{mol/L}\bullet\text{s}$ The average rate of consumption of HBr(g) over 50.0 s is 0.0032 mol/L•s.
b. The rate of formation of products is in proportion to the mole ratio given in the balanced chemical equation. Determine the mole ratio HBr:Br_2 from the balanced equation.	**b.** $\frac{n_{HBr}}{n_{Br_2}} = \frac{4}{2}$

Use the mole ratio to calculate the rate of formation of Br_2.	For every 4 mol HBr(g) consumed, 2 mol Br_2(g) is produced. $\frac{-\Delta[\text{HBr}]}{\Delta[\text{Br}_2]} = \frac{4}{2}$ $\Delta[\text{Br}_2] = \left(\frac{2}{4}\right)\left(-\Delta[\text{HBr}]\right)$ $= (0.5)\left[-(-0.0032\ \text{mol/L} \bullet \text{s})\right]$ $= 0.0016\ \text{mol/L} \bullet \text{s}$
Use the formula: $\text{rate} = \frac{\Delta\text{concentration}}{\Delta t}$ to determine the final concentration, c_{final}, of Br_2(g) at 50 s.	$\text{rate} = \frac{\Delta\text{concentration}}{\Delta t}$ $\text{rate} = \frac{(c_{\text{final}} - c_{\text{initial}})}{\Delta t}$ $0.0016\ \text{mol/L} \bullet \text{s} = \frac{(c_{\text{final}} - 0.00\ \text{mol/L})}{50\ \text{s}}$ $c_{\text{final}} = 0.080\ \text{mol/L}$ The final concentration of Br_2(g) is 0.080 mol/L.

Check Your Solution

The average rate of reaction has the correct number of significant digits and seems reasonable. The final concentration of Br_2(g) is greater than the initial concentration.

8. Consider the following reaction: $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

Over a period of 1.80 min, the average rate of formation of NO(g) is 1.04 mol/ L•s. What was the amount in moles of $O_2(g)$ consumed over this period of time?

What Is Required?
You need to determine the amount in moles of $O_2(g)$ consumed.

What Is Given?
You know the balanced chemical equation for the reaction.

Plan Your Strategy	Act on Your Strategy
Convert the unit of time from minutes to seconds.	$\Delta t = 1.8 \text{ min}$ $= 1.8 \cancel{\text{min}} \times \frac{60 \text{ s}}{\cancel{\text{min}}}$ $= 108 \text{ s}$
The rate of formation of products is in proportion to the mole ratio given in the balanced chemical equation. Determine the mole ratio $O_2(g)$:NO(g)	$\frac{n_{O2}}{n_{NO_2}} = \frac{5}{4}$
Calculate the rate of consumption of $O_2(g)$.	For every 4 mol NO(g) produced, 5 mol $O_2(g)$ is consumed. $\frac{\text{rate of consumption of } O_2(g)}{\text{rate of production of } NO_2(g)} = \frac{5 \text{ mol } O_2(g)}{4 \text{ mol } NO_2(g)}$ $\frac{\text{rate of consumption of } O_2(g)}{1.04 \text{ mol/L} \bullet \text{s}} = \frac{5 \text{ mol } O_2(g)}{4 \text{ mol } NO_2(g)}$ $\text{rate of consumption of } O_2(g) = 1.04 \text{ mol/L} \bullet \text{s} \times 1.25$ $= 1.30 \text{ mol/L} \bullet \text{s}$ The rate of consumption of O_2 is 1.30 mol/L•s.
Calculate the amount of $O_2(g)$ consumed after 108 s.	$\Delta O_2 = 1.30 \text{ mol/L} \bullet \cancel{\text{s}} \times 108 \cancel{\text{s}}$ $= 140.4 \text{ mol/L}$ $= 1.40 \times 10^2 \text{ mol/L}$ 1.40×10^2 mol of $O_2(g)$ was consumed.

Check Your Solution
The concentration of the reactant $O_2(g)$ decreased in proportion to the mole ratio in the balanced equation. The answer is reasonable, and the correct number of significant digits is shown.

9. A mass of 0.50 g of sodium metal reacts with water in 90.0 s.

$$2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g) + \text{heat}$$

a. Express the rate of consumption of Na(s) in moles per second.
b. Calculate the rate at which $H_2(g)$ is generated, in litres per second, at 30.0°C and 102.4 kPa. Assume that hydrogen is an ideal gas.

What Is Required?
a. You need to determine the rate of consumption of Na(s) at a given temperature and pressure.
b. You need to determine the rate of formation of $H_2(g)$ at a given temperature and pressure.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the mass of Na(s): $m = 0.50$ g
You know the temperature: $T = 30.0°C$
You know the pressure: $P = 102.4$ kPa
You know the elapsed time: $\Delta t = 90.0$ s

Plan Your Strategy	Act on Your Strategy
a. Using the mass and molar mass of sodium, use the formula $n = \frac{m}{M}$ to determine the amount in moles, n, of Na(s).	$n = \frac{m}{M}$ $n_{Na} = \frac{0.50\ \cancel{g}}{22.99\ \cancel{g}/\text{mol}}$ $= 0.021\ 75$ mol
Use the formula $\text{rate} = \frac{\Delta\text{quantity}}{\Delta t}$ to calculate the rate at which Na(s) is consumed in the reaction.	$\text{rate of consumption of Na(s)} = \frac{\Delta\text{quantity}}{\Delta t}$ $= \frac{0.021\ 75\ \text{mol}}{90.0\ \text{s}}$ $= 2.42 \times 10^{-4}$ mol/s Na(s) is consumed at the rate of 2.42×10^{-4} mol/s.
The rate of formation of products and the rate of consumption of reactants is in proportion to the mol ratio given in the balanced equation. Determine the mole ratio of Na(s):H_2(g) from the balanced equation.	$\frac{n_{Na}}{n_{H_2}} = \frac{2}{1}$

Calculate the rate of formation of $H_2(g)$ in mol/s.	For every 1 mol $H_2(g)$ produced, 2 mol Na(s) is consumed. $\frac{\text{rate of formation of } H_2(g)}{\text{rate of consumption of Na(s)}} = \frac{n_{H_2}}{n_{Na}}$ $\frac{\text{rate of formation of } H_2(g)}{2.42 \times 10^{-4} \text{ mol/s}} = \frac{1}{2}$ $\text{rate of formation of } H_2(g) = \frac{2.42 \times 10^{-4} \text{ mol/s}}{2}$ $= 1.21 \times 10^{-4} \text{ mol/s}$
Use the ideal gas equation, $PV = nRT$, to convert the amount in moles of $H_2(g)$ to volume of $H_2(g)$. Express the rate of formation of $H_2(g)$ in L/s.	$V = \frac{nRT}{P}$ $= \frac{(1.21 \times 10^{-4} \cancel{\text{mol}}\left(8.314 \frac{\cancel{\text{kPa}} \cdot \text{L}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)(303.15 \cancel{\text{K}})}{102.4 \cancel{\text{kPa}}}$ $= 2.98 \times 10^{-3} \text{ L}$ The rate of formation of $H_2(g)$ is 2.98×10^{-3} L/s.

Check Your Solution

The rate of formation of $H_2(g)$ is positive. The number of significant digits is shown correctly and the answer seems reasonable.

10. For the reaction shown below, the instantaneous change in concentration of NO(g) is 1.4 mol/L•s.

$$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$$

What is the rate at which $Cl_2(g)$ is consumed and the rate at which NOCl(g) is formed at this time?

What Is Required?
You need to determine the rate at which $Cl_2(g)$ is consumed
You need to determine the rate at which NOCl(g) is produced.

What Is Given?
You know the balanced chemical equation.
The instantaneous change in concentration of NO(g) is –1.4 mol/L•s.

Plan Your Strategy	**Act on Your Strategy**
The rate of formation of products and the rate of consumption of reactants is in proportion to the mole ratios given in the balanced chemical equation. Determine the mole ratios Cl_2:NO and NOCl:NO from the balanced equation.	$\frac{n_{Cl_2}}{n_{NO}} = \frac{1}{2}$ $\frac{n_{NOCl}}{n_{NO}} = \frac{2}{2}$
Calculate the rate of consumption of $Cl_2(g)$.	For every 1 mol $Cl_2(g)$ consumed, 2 mol NOCl(g) is produced. $\frac{\text{rate of consumption } Cl_2(g)}{\text{rate of consumption NO(g)}} = \frac{n_{Cl_2}}{n_{NO}}$ $\frac{\text{rate of consumption } Cl_2(g)}{-1.4 \text{ mol/L} \bullet \text{s}} = \frac{1}{2}$ $\text{rate of consumption } Cl_2(g) = \frac{-1.4 \text{ mol/L} \bullet \text{s}}{2}$ $= -0.7 \text{ mol/L} \bullet \text{s}$ $Cl_2(g)$ is consumed at a rate of 0.70 mol/L•s.
Calculate the rate of formation of NOCl(g).	Since NO(g) and NOCl(g) have a molar ratio 2:2, NOCl(g) is produced at the same rate that NO(g) is consumed. NOCl(g) is therefore produced at a rate of 1.4 mol/L•s.

Check Your Solution
The rates of consumption and formation agree with the mol ratio in the balanced equation. The answers are expressed to the correct number of significant digits.