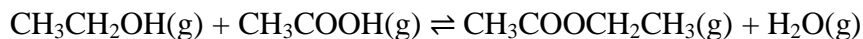# Chemistry 12

Solutions Manual Part B

Unit 4 Chemical Systems and Equilibrium

Solutions to Chapter 7 Chemical Equilibrium Practice Problems

Equilibrium Constant Expressions for Homogeneous Chemical Systems
(Student textbook page 428)

1. Write the equilibrium constant expression for the reaction at 200°C between ethanol and ethanoic acid to form ethyl ethanoate and water:

$$CH_3CH_2OH(g) + CH_3COOH(g) \rightleftharpoons CH_3COOCH_2CH_3(g) + H_2O(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
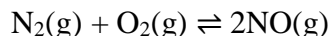
What Is Given?
You know the balanced chemical equation.

Plan Your Strategy	Act on Your Strategy
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced chemical equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[CH_3COOCH_2CH_3][H_2O]}{[CH_3CH_2OH][CH_3COOH]}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The products are in the numerator and the reactants are in the denominator and each is raised to the power of 1 which is each coefficient in the balanced equation.

2. Write the equilibrium constant expression for the reaction between nitrogen gas and oxygen gas at high temperatures is given by the following balanced chemical reaction:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
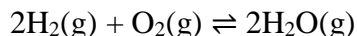
What Is Given?
You know the balanced chemical equation.

Plan Your Strategy	Act on Your Strategy
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced chemical equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[NO]^2}{[N_2][O_2]}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The product is in the numerator, the reactants are in the denominator and each is raised to power of its coefficient in the balanced chemical equation.

3. Write the equilibrium constant expression for the reaction between hydrogen gas and oxygen gas to form water vapour:

$$2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
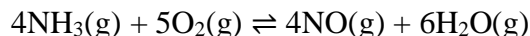
What Is Given?
You know the balanced chemical equation.

Plan Your Strategy	Act on Your Strategy
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced chemical equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[H_2O]^2}{[H_2]^2[O_2]}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The product is in the numerator and the reactants are in the denominator and each is raised to power of its coefficient in the balanced chemical equation.

4. Write the equilibrium constant expression for the oxidation of ammonia (one of the reactions in the manufacture of nitric acid):

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
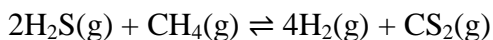
What Is Given?
You know the balanced equation.

Plan Your Strategy	Act on Your Strategy
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The products are in the numerator, the reactants are in the denominator, and each is raised to power of its coefficient in the balanced chemical equation.

5. Write the equilibrium constant expression for the reaction in which hydrogen sulfide reacts with methane and forms hydrogen gas and carbon disulfide as represented by the following reaction:

$$2H_2S(g) + CH_4(g) \rightleftharpoons 4H_2(g) + CS_2(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the balanced equation.

Plan Your Strategy	Act on Your Strategy
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[H_2]^4[CS_2]}{[H_2S]^2[CH_4]}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The products are in the numerator, the reactants are in the denominator, and each is raised to power of its coefficient in the balanced equation.

6. Write the equilibrium constant expression for the the reaction in which antimony pentachloride, $SbCl_5$ (g), decomposes into antimony trichloride, $SbCl_3$(g), and chlorine gas, Cl_2 (g), in an equilibrium reaction.

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
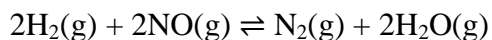
What Is Given?
You know the reactant is antimony pentachloride, $SbCl_5$ (g) and the products are antimony trichloride, $SbCl_3$(g), and chlorine gas, Cl_2(g).

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction.	$SbCl_5(g) \rightleftharpoons SbCl_3(g) + Cl_2(g)$
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[SbCl_3][Cl_2]}{[SbCl_5]}$

Check Your Solution
The equation is balanced. The square brackets indicate the concentration in mol/L. The products are in the numerator, the reactant is in the denominator, and each is raised to power of its coefficient in the balanced equation.

7. Write the equilibrium constant expression for the reaction in which hydrogen gas displaces nitrogen from nitrogen monoxide gas to form nitrogen gas and water vapour according to the following reaction:

$$2H_2(g) + 2NO(g) \rightleftharpoons N_2(g) + 2H_2O(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the balanced equation.

Plan Your Strategy	Act on Your Strategy
The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[N_2][H_2O]^2}{[H_2]^2[NO]^2}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The products are in the numerator, the reactants are in the denominator, and each is raised to power of its coefficient in the balanced equation.

8. Write the equilibrium constant expression for the reaction in which sulfur trioxide gas, $SO_3(g)$, decomposes into sulfur dioxide gas, $SO_2(g)$, and oxygen gas, $O_2(g)$, in an equilibrium reaction.
a. Write the balanced chemical equation for the equilibrium reaction.
b. Write the equilibrium constant expression.

What Is Required?
a. You need to write the balanced equation.
b. You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the reactant is sulfur trioxide gas, $SO_3(g)$, and the products are sulfur dioxide gas, $SO_2(g)$, and oxygen gas, $O_2(g)$.

Plan Your Strategy	Act on Your Strategy
a. Use the given reactant and products to write the balanced chemical equation for the reaction.	$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$
b. The expression for the equilibrium constant follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$$K_{eq} = \frac{[SO_2]^2[O_2]}{[SO_3]^2}$$

Check Your Solution
The equation is balanced. The square brackets indicate the concentration in mol/L. The products are in the numerator, the reactant is in the denominator, and each is raised to power of its coefficient in the balanced equation.

9. Write the equilibrium constant expression for the reaction in which nitrogen monoxide gas reacts with oxygen gas in an equilibrium reaction where nitrogen dioxide gas forms.
a. Write the balanced chemical equation for the equilibrium reaction.
b. Write the equilibrium constant expression.

What Is Required?
a. You need to write the formulas for the reactants and product and write the balanced equation for the reaction.
b. You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the reactants are nitrogen monoxide gas and oxygen gas and the product is nitrogen dioxide gas.

Plan Your Strategy	Act on Your Strategy
a. Write the chemical formulas for the given reactants and product and write the balanced equation for the reaction.	$2NO(g) + O_2(g) \rightleftharpoons 2NO_2 (g)$
b. The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[NO_2]^2}{[NO]^2[O_2]}$

Check Your Solution
The equation is balanced. The square brackets indicate the concentration in mol/L. The product is in the numerator, the reactants are in the denominator, and each is raised to power of its coefficient in the balanced equation.

10. Write the equilibrium constant expression for the reaction in which oxygen gas displaces chlorine from hydrogen chloride gas in an equilibrium reaction where chlorine gas and water vapour form.
a. Write the balanced chemical equation for the equilibrium reaction.
b. Write the equilibrium constant expression.

What Is Required?
a. You need to write the formulas for the reactants and product and write the balanced equation for the reaction.
b. You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know that oxygen gas reacts with hydrogen chloride gas to produce chlorine gas and water vapour.

Plan Your Strategy	Act on Your Strategy
a. Write the chemical formulas for the given reactants and products and write the balanced chemical equation for the reaction.	$4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g)$
b. The expression for the equilibrium constant follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Follow this format using the reactants, products, and coefficients in the given equation.	$K_{eq} = \dfrac{[H_2O]^2[Cl_2]^2}{[HCl]^4[O_2]}$

Check Your Solution
The equation is balanced. The square brackets indicate the concentration in mol/L. The products are in the numerator, the reactants are in the denominator, and each is raised to power of its coefficient in the balanced equation.

Equilibrium Constant Expressions for Heterogeneous Chemical Systems
(Student textbook page 430)

11. Write the equilibrium constant expression. In the decomposition of solid ammonium chloride, $NH_4Cl(s)$, the products ammonia gas, $NH_3(g)$, and hydrogen chloride gas, $HCl(g)$, form:

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$

What Is Required?
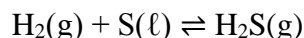You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the balanced chemical equation and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{\left[NH_3(g)\right]\left[HCl(g)\right]}{\left[NH_4Cl(s)\right]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression	$$K\left[NH_4Cl(s)\right] = \left[NH_3(g)\right]\left[HCl(g)\right]$$
The terms on the left side of the equation are combined and become the K_{eq}.	$K_{eq} = [NH_3][HCl]$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The products are in the numerator, and each is raised to power of its coefficient in the balanced equation.
There are no reactants in the denominator.

12. Write the equilibrium constant expression. Hydrogen gas and liquid sulfur react and form hydrogen sulfide gas:

$$H_2(g) + S(\ell) \rightleftharpoons H_2S(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
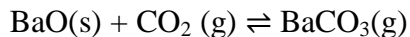
What Is Given?
You know the balanced chemical equation and the state of each component.

Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[H_2S(g)]}{[H_2(g)][S(\ell)]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K[S(\ell)] = \frac{[H_2S(g)]}{[H_2(g)]}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{[H_2S(g)]}{[H_2(g)]}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The product is in the numerator, the reactant is in the denominator, and each is raised to power of its coefficient in the balanced equation.

13. Write the equilibrium constant expression. Barium oxide and carbon dioxide react:

$BaO(s) + CO_2 (g) \rightleftharpoons BaCO_3(g)$
(NOTE: The state of $BaCO_3$ is given as a gas in this problem. However, this is not the case. It is a solid at room temperature. The problem will be shown for the solid state.
$BaO(s) + CO_2 (g) \rightleftharpoons BaCO_3(s)$)

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the balanced equation and the physical state of each component.

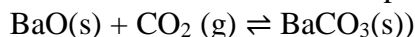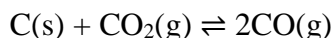| Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[BaCO_3(s)]}{[BaO(s)][CO_2(g)]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K\frac{[BaO(s)]}{[BaCO_3(s)]} = \frac{1}{[CO_2(g)]}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{1}{[CO_2]}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. There are no products in the numerator. The reactant, $CO_2(g)$, is in the denominator raised to the power of its coefficient in the balanced equation.

14. Write the equilibrium constant expression. Solid carbon and carbon dioxide gas react and form carbon monoxide:

$$C(s) + CO_2(g) \rightleftharpoons 2CO(g)$$

What Is Required?
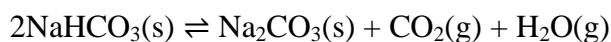You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the balanced chemical equation and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $K_{eq} = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$K = \dfrac{[CO(g)]^2}{[C(s)][CO_2(g)]}$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$K[C(s)] = \dfrac{[CO(g)]^2}{[CO_2(g)]}$
The terms on the left side of the equation are combined and become the K_{eq}.	$K_{eq} = \dfrac{[CO(g)]^2}{[CO_2(g)]}$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The product is in the numerator, the reactant is in the denominator, and each is raised to the power of its coefficient in the balanced equation.

15. Write the equilibrium constant expression. Solid sodium hydrogen carbonate, $NaHCO_3(s)$, decomposes into solid sodium carbonate, $NaHCO_3(s)$, carbon dioxide gas, $CO_2(g)$, and water vapour, $H_2O(g)$:

$$2NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + CO_2(g) + H_2O(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
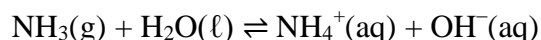
What Is Given?
You know the balanced chemical equation and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[Na_2CO_3(s)][CO_2(g)][H_2O(g)]}{[NaHCO_3(s)]^2}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K\frac{[NaHCO_3(s)]^2}{[Na_2CO_3(s)]} = [CO_2(g)][H_2O(g)]$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = [CO_2(g)][H_2O(g)]$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The products that change in concentration are in the numerator and the denominator is 1 because no reactants remain in the denominator. There is no denominator in the final equation.

16. When ammonia gas reacts with water, ammonium ions, NH_4^+, and hydroxide ions, OH^-, form according to the following reaction:

$$NH_3(g) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the balanced chemical equation and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(g)][H_2O(\ell)]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K[H_2O(\ell)] = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(g)]}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(g)]}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The products are in the numerator, the reactant that changes in concentration is in the denominator, and each is raised to the power of its coefficient in the balanced equation.

17. Solid iron reacts with water vapour as represented by the following reaction:

$$3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$$

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.
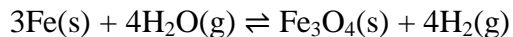
What Is Given?
You know the balanced chemical equation and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[Fe_3O_4(s)][H_2(g)]^4}{[Fe(s)]^3[H_2O(g)]^4}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K\frac{[Fe(s)]^3}{[Fe_3O_4(s)]} = \frac{[H_2(g)]^4}{[H_2O(g)]^4}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{[H_2(g)]^4}{[H_2O(g)]^4}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The product that changes in concentration is in the numerator, the reactant that changes in concentration is in the denominator, and each is raised to the power of its coefficient in the balanced equation.

18. In a closed container, solid magnesium and oxygen gas react to form magnesium oxide on the surface of the magnesium. This process is an equilibrium process. Write the equilibrium constant expression for this process.

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the reactants and the product and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
Write the chemical formulas of the reactants and the product and write the balanced chemical equation for the reaction.	$2Mg(s) + O_2(g) \rightleftharpoons 2MgO(s)$
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[MgO(s)]^2}{[Mg(s)]^2[O_2(g)]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K\,\frac{[Mg(s)]^2}{[MgO(s)]^2} = \frac{1}{[O_2(g)]}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{1}{[O_2(g)]}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. There are no products in the numerator. The reactant $O_2(g)$ is in the denominator raised to the power of its coefficient in the balanced equation.

19. In the synthesis of hydrogen iodide gas, HI(g), hydrogen gas reacts with solid iodine in an equilibrium process. Write the equilibrium constant expression.

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the reactants and the product and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
Write the chemical formulas of the reactants and the product and write the balanced chemical equation for the reaction.	$H_2(g) + I_2(s) \rightleftharpoons 2HI(g)$
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C], and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[HI(g)]^2}{[H_2(g)][I_2(s)]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K[I_2(s)] = \frac{[HI(g)]^2}{[H_2(g)]}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{[HI(g)]^2}{[H_2(g)]}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The product is in the numerator, the reactant that changes in concentration is in the denominator, and each is raised to the power of its coefficient in the balanced equation.

20. Solid carbon reacts with oxygen gas to form carbon dioxide gas in an equilibrium reaction. Write the equilibrium constant expression for this process.

What Is Required?
You need to write the expression for the equilibrium constant, K_{eq}.

What Is Given?
You know the reactants and the product and the physical state of each component.

Plan Your Strategy	Act on Your Strategy
Write the chemical formulas of the reactants and the product and write the balanced chemical equation for the reaction.	$C(s) + O_2(g) \rightleftharpoons CO_2(g)$
The equilibrium constant expression follows the form: $$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$ where [A], [B], [C] and [D] represent the concentrations of the reactants and products at equilibrium. The exponents a, b, c, and d are the stoichiometric coefficients from the balanced equation. Use this general form for K_{eq} to write a temporary equilibrium expression, K, for the reaction.	$$K = \frac{[CO_2(g)]}{[C(s)][O_2(g)]}$$
The expression for the equilibrium constant is affected only by the concentrations that change. Since the concentrations of pure substances in the liquid or solid state do not change, their constant concentration is incorporated into the temporary equilibrium expression.	$$K[C(s)] = \frac{[CO_2(g)]}{[O_2(g)]}$$
The terms on the left side of the equation are combined and become the K_{eq}.	$$K_{eq} = \frac{[CO_2(g)]}{[O_2(g)]}$$

Check Your Solution
The square brackets indicate the concentration in mol/L. The expression for K_{eq} contains only components of the system that undergo a change in concentration. The product is in the numerator, the reactant that changes is in concentration in the denominator, and each is raised to the power of its coefficient in the balanced equation.

Using Le Châtelier's Principle
(Student textbook page 439)

21. In which direction does the following reaction shift if the temperature increases? Explain why.
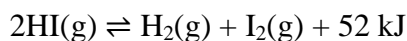
$$2HI(g) \rightleftharpoons H_2(g) + I_2(g) + 52 \text{ kJ}$$

What Is Required?
You have to determine the direction a system at equilibrium will shift when the temperature increases.

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

Plan Your Strategy	Act on Your Strategy
Identify the heat term and in which direction the endothermic reaction occurs.	The reverse reaction is endothermic.
Identify how the system will temporarily shift before equilibrium is restored.	The position of equilibrium will shift to use some of the heat that is introduced to the system with an increase in temperature. The reaction will shift to the left, in the direction of the endothermic reaction.

Check Your Solution
The shift in the position of equilibrium agrees with the prediction based on Le Châtelier's principle.

22. In the gaseous equilibrium system shown below, the volume of the container is increased, causing a decrease in pressure. In which direction does the reaction shift? Explain why.

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$$

What Is Required?
You have to determine the direction a system at equilibrium will shift when the volume of the container is increased causing a decrease in the pressure.

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

Plan Your Strategy	Act on Your Strategy
Determine on which side of the equation there are more particles.	There is the same number of particles on both sides of the reaction.
Identify how the system will temporarily shift before equilibrium is restored.	As the reaction proceeds, there is no change in the number of gas molecules on each side of the equation. Therefore, increasing the volume of the container has no effect on the position of equilibrium.

Check Your Solution
The position of equilibrium changes only when the rate of either the forward or the reverse reaction is favoured. There will be no change in the position of equilibrium.

23. In the gaseous equilibrium system shown below, the volume of the container is increased, causing a decrease in pressure. In which direction does the reaction shift? Explain why.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

What Is Required?
You have to determine the direction a system at equilibrium will shift when the volume of the container is increased causing a decrease in the pressure.

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

Plan Your Strategy	Act on Your Strategy
Determine on which side of the equation there are more particles.	There are more gas molecules on the left side of the equation.
Identify how the system will temporarily shift before equilibrium is restored.	The position of equilibrium will shift to favour the formation of more particles. Therefore, increasing the volume of the container causes the reaction to shift to the left.

Check Your Solution
The shift in the position of equilibrium agrees with the prediction based on Le Châtelier's principle.

24. In the gaseous equilibrium system shown below, the volume of the container is increased, causing a decrease in pressure. In which direction does the reaction shift? Explain why.

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$$

What Is Required?
You have to determine the direction a system at equilibrium will shift when the volume of the container is increased causing a decrease in the pressure.
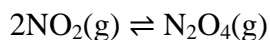
What Is Given?
You know the balanced equation for the reaction.

Plan Your Strategy	Act on Your Strategy
Determine on which side of the equation there are more particles.	There are more gas molecules on the right side of the equation.
Identify how the system will temporarily shift before equilibrium is restored.	The position of equilibrium will shift to favour the formation of more particles. Therefore, increasing the volume of the container causes the reaction to shift to the right.

Check Your Solution
The shift in the position of equilibrium agrees with the prediction based on Le Châtelier's principle.

25. In the gaseous equilibrium system shown below, the volume of the container is increased, causing a decrease in pressure. In which direction does the reaction shift? Explain why.

$$CH_4(g) + 2H_2S(g) \rightleftharpoons CS_2(g) + 4H_2(g)$$

What Is Required?
You have to determine the direction a system at equilibrium will shift when the volume of the container is increased causing a decrease in the pressure.

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

Plan Your Strategy	Act on Your Strategy
Determine on which side of the equation there are more particles.	There are more gas molecules on the right side of the equation.
Identify how the system will temporarily shift before equilibrium is restored.	The position of equilibrium will shift to favour the formation of more particles. Therefore, increasing the volume of the container causes the reaction to shift to the right.

Check Your Solution
The shift in the position of equilibrium agrees with the prediction based on Le Châtelier's principle.

26. Toluene, $C_7H_8(\ell)$, is an important organic solvent. It is made industrially from methylcyclohexane, $C_7H_{14}(g)$:

$$C_7H_{14}(g) + \text{heat} \rightleftharpoons C_7H_8(g) + 3H_2(g)$$
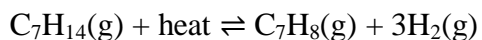
State three different changes to an equilibrium mixture of these reacting gases that would shift the reaction toward greater production of toluene.

What Is Required?
You need to determine three different changes to an equilibrium mixture of reacting gases that would shift the reaction toward greater production of toluene, $C_7H_8(\ell)$.

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

Plan Your Strategy	Act on Your Strategy
Identify the side of the equation on which toluene, $C_7H_8(g)$, is found.	Toluene, $C_7H_8(g)$, is on the right side of the reaction
Apply Le Châtelier's principle to determine how the forward reaction can be favoured.	The forward reaction is endothermic. An increase in temperature will cause the position of equilibrium to shift to use some of the heat that is introduced to the system with an increase in temperature. The reaction will shift to the right, in the direction of the endothermic reaction. There are more particles on the right side of the equation. An increase in the volume of the system (or a decrease in the total pressure) will cause the position of equilibrium to shift to favour the formation of more particles. The reaction will shift to the right. An increase in the amount of reactant methylcyclohexane, $C_7H_{14}(g)$, will favour product formation. (Alternatively, product can be removed for the same effect). The reaction will shift to the right.

Check Your Solution
The shift in the position of equilibrium agrees with the predictions based on Le Châtelier's principle.

27. In which direction does the following reaction shift as a result of each of the following changes?

$$2NO(g) + 2H_2(g) \rightleftharpoons N_2(g) + 2H_2O(g) + heat$$

a. increasing the pressure of gases in the reaction vessel by decreasing the volume
b. increasing the pressure of gases in the reaction vessel by adding inert argon gas while keeping the volume of the vessel constant
c. increasing the temperature

What Is Required?
You need to predict in which direction the reaction will shift as a result of:
a. increasing the pressure of gases in the reaction vessel by decreasing the volume
b. increasing the pressure of gases in the reaction vessel by adding inert argon gas while keeping the volume of the vessel constant
c. increasing the temperature

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

Plan Your Strategy	Act on Your Strategy
a. Identify on which side of the reaction there are fewer particles. **b.** Identify the effect of adding an inert gas at constant volume. **c.** Identify the heat term and in which direction in which the endothermic reaction occurs.	**a.** There are fewer particles on the right side of the reaction. **b.** At constant volume, there is no change in the concentration of any of the gases in this reaction. **c.** The reverse reaction is endothermic.
For each change, identify how the system will temporarily shift before equilibrium is restored.	**a.** The position of equilibrium will shift to favour the formation of fewer particles. Therefore, decreasing the volume of the container causes the reaction to shift to the right. **b.** There will no change in the position of equilibrium. **c.** The position of equilibrium will shift to use some of the heat that is introduced to the system with an increase in temperature. The reaction will shift to the left, in the direction of the endothermic reaction.

Check Your Solution
The shift in the positions of equilibrium agrees with the predictions based on Le Châtelier's principle.

28. In question 27, how does each of the changes affect K_{eq}? Explain your answers.

What Is Required?
You need to determine how each of the changes in the previous question affects K_{eq}.

What Is Given?
You know the balanced thermochemical equation from the previous question.

Plan Your Strategy	Act on Your Strategy
a. Identify the mole ratio of the components in the reaction.	The equation remains the same and the products and reactants are in the same mole ratio. Since the ratio of products to reactants does not change, the value of K_{eq} does not change.
b. Identify the mole ratio of the components in the reaction.	The equation remains the same and the products and reactants are in the same mole ratio. Since the ratio of products to reactants does not change, the value of K_{eq} does not change.
c. Identify which reaction is endothermic.	The reverse reaction is endothermic. An increase in temperature results in the formation of more reactant. A temperature change affects the forward and reverse reactions differently. The value of K_{eq} decreases when the temperature increases.

Check Your Solution
The answers agree with the predictions based on Le Châtelier's principle.

29. The following reaction is endothermic when read from left to right $N_2O_4(g) \rightleftharpoons 2NO_2(g)$. In which direction does the reaction shift as a result of each of the following changes?
a. adding $NO_2(g)$
b. adding a catalyst

What Is Required?
You need to determine in which direction the reaction shift as a result of each of:
a. adding $NO_2(g)$
b. adding a catalyst

What Is Given?
You know the balanced equation for the reaction and that the heat term is on the left.

Plan Your Strategy	Act on Your Strategy
a. Identify on which side of the reaction $NO_2(g)$ is found.	$NO_2(g)$ is on the right.
Identify how the system will temporarily shift before equilibrium is restored.	Increasing the concentration of $NO_2(g)$ will favour the formation of reactant. Therefore, adding more $NO_2(g)$ causes the reaction to shift to the left.
b. Identify how a catalyst affects the rate of reaction.	A catalyst increases the rate of both the forward and reverse reactions in the same way. There is no effect on the position of equilibrium.

Check Your Solution
The shift in the position of equilibrium agrees with the predictions based on Le Châtelier's principle.

30. In question 29, how do each of the changes affect K_{eq}? Explain your answers.

What Is Required?
You need to determine how each of the changes in the previous question affects K_{eq}.

What Is Given?
From the previous question, you know the balanced equation and that the heat term is on the left.

| Plan Your Strategy	Act on Your Strategy
Identify the mole ratio of the components in the reaction.	Neither of the changes affects K_{eq} because changing the concentration of reactants and products or adding a catalyst does not affect the ratio of products to reactants in the reaction.

Check Your Solution
The answers agree with the predictions based on Le Châtelier's principle.

Calculating an Equilibrium Constant for a Homogeneous Chemical System
Calculating the Molar Concentration of a Substance Using the Equilibrium Constant
(Student textbook page 444)

31. When 1.0 mol of ammonia gas is injected into a 0.50 L flask at a given temperature, the following reaction proceeds to equilibrium:

$$2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$$
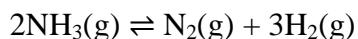
At equilibrium, the mixture contains 0.30 mol $H_2(g)$, 0.10 mol $N_2(g)$, and 0.80 mol $NH_3(g)$.
a. Write the equilibrium expression for the reaction.
b. What is the value of the equilibrium constant?

What Is Required?
You need to write the expression for the equilibrium constant.
You need to calculate the value of the equilibrium constant.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container is 0.50 L.
You know the amount in moles of the reactant and products at equilibrium:

$n_{NH_3} = 0.80$ mol

$n_{N_2} = 0.10$ mol

$n_{H_2} = 0.30$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to determine the molar equilibrium concentration, c, of each component.	$\begin{aligned}[NH_3] &= \dfrac{0.80 \text{ mol}}{0.50 \text{ L}} \\ &= 1.6 \text{ mol/L}\end{aligned}$ $\begin{aligned}[H_2] &= \dfrac{0.30 \text{ mol}}{0.50 \text{ L}} \\ &= 0.6 \text{ mol/L}\end{aligned}$ $\begin{aligned}[N_2] &= \dfrac{0.10 \text{ mol}}{0.50 \text{ L}} \\ &= 0.20 \text{ mol/L}\end{aligned}$
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[N_2][H_2]^3}{[NH_3]^2}$

Substitute the equilibrium concentrations into the expression for K_{eq} and solve.	$K_{eq} = \dfrac{[N_2][H_2]^3}{[NH_3]^2}$ $= \dfrac{(0.20)(0.60)^3}{(1.6)^2}$ $= 0.01687$ $= 0.017$ The value of the equilibrium constant is 0.017.

Check Your Solution

The expression for K_{eq} shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly and the answer has the correct number of significant digits.

32. Phosphorus trichloride, $PCl_3(g)$, reacts with chlorine, $Cl_2(g)$, to form phosphorus pentachloride, $PCl_5(g)$, as follows:

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$$
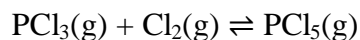
When 0.75 mol each of $PCl_3(g)$ and $Cl_2(g)$ were placed into an 8.0 L reaction vessel at 500 K, the equilibrium concentration of the mixture was determined to contain 0.035 mol/L of $PCl_3(g)$ and $Cl_2(g)$. The value of K_{eq} at 500 K is 49. Calculate the equilibrium concentration of $PCl_5(g)$.

What Is Required?
You need to calculate the equilibrium concentration of $PCl_5(g)$.

What Is Given?
You know the balanced equation for the reaction.
You know the volume of the reaction vessel: 8.0 L
You know the concentrations at equilibrium: $[PCl_3(g)] = [Cl_2(g)] = 0.035$ mol/L
You know that the value of K_{eq} at 500 K is 49.

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_{eq}.	$$K_{eq} = \frac{[PCl_5]}{[PCl_3][Cl_2]}$$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve for $[PCl_5]$.	$$K_{eq} = \frac{[PCl_5]}{[PCl_3][Cl_2]}$$ $$49 = \frac{[PCl_5]}{(0.035 \text{ mol/L})(0.035 \text{ mol/L})}$$ $[PCl_5] = 0.06002$ $\qquad = 0.060$ mol/L The equilibrium concentration of $PCl_5(g)$ is 0.060 mol/L.

Check Your Solution
The expression for K_{eq} shows the product in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

33. Iodine, $I_2(g)$, and bromine, $Br_2(g)$, react to form iodine monobromide, $IBr(g)$:

$$I_2(g) + Br_2(g) \rightleftharpoons 2IBr(g)$$
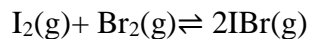
At 150°C, an equilibrium mixture in a 2.0 L flask contains 0.024 mol of iodine, 0.050 mol of bromine, and 0.38 mol of iodine monobromide. Determine K_{eq} for this reaction.

What Is Required?
You need to determine K_{eq} for this reaction.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume, V, of the container: 2.0 L
You know the amount in moles of the components at equilibrium:

$n_{I_2} = 0.024$ mol

$n_{B_2} = 0.050$ mol

$n_{IBr} = 0.38$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to determine the molar equilibrium concentration, c, of each component.	$[I_2] = \dfrac{0.024 \text{ mol}}{2.0 \text{ L}}$ $= 0.012$ mol/L $[Br_2] = \dfrac{0.050 \text{ mol}}{2.0 \text{ L}}$ $= 0.025$ mol/L $[IBr] = \dfrac{0.38 \text{ mol}}{2.0 \text{ L}}$ $= 0.19$ mol/L
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[IBr]^2}{[I_2][Br_2]}$

Substitute the equilibrium concentrations into the expression for K_{eq} and solve.	$K_{eq} = \dfrac{[IBr]^2}{[I_2][Br_2]}$ $= \dfrac{(0.19 \text{ mol/L})^2}{(0.012 \text{ mol/L})(0.025 \text{ mol/L})}$ $= 120.33$ $= 1.2 \times 10^2$ K_{eq} for the reaction is 1.2×10^2.

Check Your Solution
The expression for K_{eq} shows the product in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly and the answer has the correct number of significant digits.

34. For the synthesis of ammonia, the equilibrium reaction is as follows:

$$N_2 + 3H_2(g) \rightleftharpoons 2NH_3(g)$$
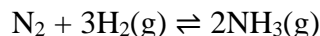
At 475 K, $K_{eq} = 625$. The equilibrium mixture is analyzed and found to contain a nitrogen gas concentration of 2.11 mol/L and a hydrogen gas concentration of 1.74 mol/L. What is the equilibrium concentration of ammonia gas in this mixture?

What Is Required?
You need to determine the equilibrium concentration of ammonia gas in the mixture.

What Is Given?
You know the balanced equation for the reaction.
You know the concentration of nitrogen and hydrogen at equilibrium:
   $[N_2(g)] = 2.22$ mol/L
   $[H_2(g)] = 1.74$ mol/L
You know $K_{eq} = 625$ at 475 K.

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve for $[NH_3]$	$K_{eq} = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$ $625 = \dfrac{[NH_3]^2}{(2.11 \text{ mol/L})(1.74 \text{ mol/L})^3}$ $[NH_3] = 83.349$ mol/L $= 83.3$ mol/L The equilibrium concentration of $NH_3(g)$ is 83.3 mol/L.

Check Your Solution
The expression for K_{eq} shows the product in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

35. Hydrogen sulfide, $H_2S(g)$, is a pungent, poisonous gas. At 1400 K, an equilibrium mixture contains 0.013 mol/L of hydrogen, $H_2(g)$, 0.18 mol/L of hydrogen sulfide, and an undetermined amount of sulfur, $S_2(g)$. The reaction is

$$2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$$
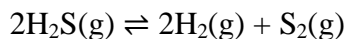
If the value of K_{eq} at 1400 K is 2.4×10^{-4}, what concentration of $S_2(g)$ is present at equilibrium?

What Is Required?
You need to determine the concentration of $S_2(g)$ present at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the concentrations of hydrogen sulfide and hydrogen at equilibrium:
  $[H_2S(g)] = 0.18$ mol/L
  $[H_2(g)] = 0.013$ mol/L
You know that $K_{eq} = 2.4 \times 10^{-4}$ at 1400 K.

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2]^2[S_2]}{[H_2S]^2}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve for $[S_2(g)]$.	$K_{eq} = \dfrac{[H_2]^2[S_2]}{[H_2S]^2}$ $2.4 \times 10^{-4} = \dfrac{(0.013 \text{ mol/L})^2[S_2]}{(0.18 \text{ mol/L})^2}$ $[S_2] = 0.04601 \text{ mol/L}$ $= 0.046 \text{ mol/L}$ The equilibrium concentration of $S_2(g)$ is 0.046 mol/L.

Check Your Solution
The expression for K_{eq} shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

36. Methane, $CH_4(g)$, ethyne, $C_2H_2(g)$, and hydrogen, $H_2(g)$, form the following equilibrium mixture:

$$2CH_4(g) \rightleftharpoons C_2H_2(g) + 3H_2(g)$$

While studying this reaction, a chemist analyzed a 4.0 L sealed flask containing an equilibrium mixture of the gases at 1700°C and found 0.64 mol of ethyne and 0.92 mol of hydrogen gas. If $K_{eq} = 0.15$ for the reaction at 1700°C, what is the expected concentration of methane gas at equilibrium?

What Is Required?
You need to determine the expected concentration of methane gas at equilibrium?

What Is Given?
You know the balanced equation for the reaction.
You know the volume of the container is 4.0 L
You know the amount in moles of ethyne and hydrogen at equilibrium:
$$n_{C_2H_2} = 0.64 \text{ mol}$$
$$n_{H_2} = 0.92 \text{ mol}$$
You know that $K_{eq} = 0.15$ at 1500°C.

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to determine the molar equilibrium concentration, c, of each component.	$[C_2H_2] = \dfrac{0.64 \text{ mol}}{4.0 \text{ L}}$ $= 0.16 \text{ mol/L}$ $[H_2] = \dfrac{0.92 \text{ mol}}{4.0 \text{ L}}$ $= 0.23 \text{ mol/L}$
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[C_2H_2][H_2]^3}{[CH_4]^2}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve for $[C_2H_2(g)]$.	$$K_{eq} = \frac{[C_2H_2][H_2]^3}{[CH_4]^2}$$ $$0.15 = \frac{(0.16 \text{ mol/L})(0.23 \text{ mol/L})^3}{[CH_4]^2}$$ $$[CH_4]^2 = \frac{(0.16 \text{ mol/L})(0.23 \text{ mol/L})^3}{(0.15)}$$ $$= 0.012978 \ (\text{mol/L})^2$$ $$[CH_4] = \sqrt{0.012978} \ \text{mol/L}$$ $$= 0.11392 \ \text{mol/L}$$ $$= 0.11 \ \text{mol/L}$$ The equilibrium concentration of $CH_4(g)$ is 0.11 mol/L.

Check Your Solution
The expression for K_{eq} shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

37. Nitrogen gas and oxygen gas are present in large quantities in the atmosphere. At a certain temperature, the value of K_{eq} is 4.2×10^{-8} for the following reaction:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$
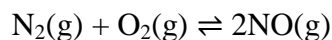
What can you infer about the equilibrium concentration of $NO(g)$ at this temperature?

What Is Required?
You must infer the equilibrium concentration of $NO(g)$ based upon a given value of K_{eq}.

What Is Given?
You know that for this reaction K_{eq} is 4.2×10^{-8}.
You know the balanced chemical equation for the reaction.

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant for the reaction.	$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$
Draw an inference based on the value of $K_{eq} = 4.2 \times 10^{-8}$.	The small value for the K_{eq} indicates that the position of equilibrium favours the reactants. Since K_{eq} is small, the concentration of $NO(g)$ will be small.

Check Your Solution
The answer agrees with the prediction based on Le Châtelier's principle.

38. At high temperatures, carbon dioxide gas, CO_2 (g), decomposes into carbon monoxide gas, CO(g), and oxygen gas, O_2 (g). The concentrations of the gases at equilibrium are $[CO_2] = 1.2$ mol/L, $[CO] = 0.35$ mol/L, and $[O_2] = 0.15$ mol/L. Determine K_{eq} at the equilibrium temperature.

What Is Required?
You need to determine the concentration of K_{eq} for the reaction.

What Is Given?
You know the balanced equation for the reaction.
You know the concentration of the gases at equilibrium:
    $[CO_2] = 1.2$ mol/L
    $[CO] = 0.35$ mol/L
    $[O_2] = 0.15$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[CO]^2[O_2]}{[CO_2]^2}$
Substitute the equilibrium concentrations into the expression for K_{eq} and solve.	$K_{eq} = \dfrac{[CO]^2[O_2]}{[CO_2]^2}$ $= \dfrac{(0.35)^2(0.15)}{(1.2)^2}$ $= 0.01276$ $= 0.013$ The value of the equilibrium constant is 0.013.

Check Your Solution
The expression for K_{eq} shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly and the answer has the correct number of significant digits.

39. The following reaction took place in a sealed flask at 250°C:

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

At equilibrium, the gases in the flask had the following concentrations:
$[PCl_5] = 1.2 \times 10^{-2}$ mol/L
$[PCl_3] = 1.5 \times 10^{-2}$ mol/L
$[Cl_2] = 1.5 \times 10^{-2}$ mol/L
Calculate the value of the equilibrium constant at 250°C.

What Is Required?
You need to calculate the value of the equilibrium constant.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the concentration of the gases at equilibrium:
$[PCl_5] = 1.2 \times 10^{-2}$ mol/L
$[PCl_3] = 1.5 \times 10^{-2}$ mol/L
$[Cl_2] = 1.5 \times 10^{-2}$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[PCl_3][Cl_2]}{[PCl_5]}$
Substitute the equilibrium concentrations into the expression for K_{eq} and solve.	$K_{eq} = \dfrac{[PCl_3][Cl_2]}{[PCl_5]}$ $= \dfrac{(1.5\times10^{-2} \ \text{mol/L})(1.5\times10^{-2} \ \text{mol/L})}{(1.2\times10^{-2} \ \text{mol/L})}$ $= 0.01875$ $= 0.019$ The equilibrium constant for the reaction is 1.9×10^{-2}.

Check Your Solution
The expression for K_{eq} shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly and the answer has the correct number of significant digits.

40. Hydrogen iodide gas decomposes into hydrogen gas and iodine gas according to the following equation:

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$
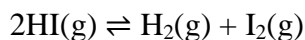
A 2.25 L container containing this equilibrium system is analyzed and found to contain 2.34 mol of hydrogen iodide and 1.76 mol of hydrogen gas. If the K_{eq} for the reaction is 2.86×10^{-2}, what amount (in moles) of iodine gas is present in the system?

What Is Required?
You need to calculate the amount in moles of iodine gas is present at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 2.25 L
You know the amount in moles of hydrogen iodide and hydrogen gas at equilibrium:

$n_{HI} = 2.34$ mol

$n_{H_2} = 1.76$ mol

You know the K_{eq} for the reaction is 2.86×10^{-2}.

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to determine the molar equilibrium concentration, c, of each component.	$[HI(g)] = \dfrac{2.34 \text{ mol}}{2.25 \text{ L}}$ $= 1.04$ mol/L $[H_2(g)] = \dfrac{1.76 \text{ mol}}{2.25 \text{ L}}$ $= 0.7822$ mol/L
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2][I_2]}{[HI]^2}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve for $[I_2(g)]$.	$K_{eq} = \dfrac{[H_2][I_2]}{[HI]^2}$ $2.86 \times 10^{-2} = \dfrac{(0.7822 \text{ mol/L})[I_2]}{(1.04 \text{ mol/L})^2}$ $[I_2] = 0.03955$ mol/L

Use the formula $n = cV$ to determine the amount in moles, n, of $I_2(g)$	$n_{I_2} = cV$ $= (0.03955 \text{ mol/}\cancel{L})(2.25 \ \cancel{L})$ $= 0.08898 \text{ mol}$ $= 0.0890 \text{ mol}$ The amount in moles of $I_2(g)$ is 8.90×10^{-2} mol.

Check Your Solution

The expression for K_{eq} shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The equilibrium concentrations have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

Calculating Equilibrium Constants Using Gas Partial Pressures
(Student textbook page 447)

41. In the decomposition of water vapour, equilibrium is established according to the following equation:

$$2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$$
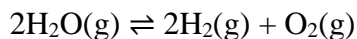
At a given temperature, the partial pressures are found to be as follows:
$P_{H_2O} = 0.63$ atm, $P_{H_2} = 1.15$ atm, and $P_{O_2} = 2.67$ atm
Determine the value of the equilibrium constant, K_p.

What Is Required?
You need to determine the value of the equilibrium constant, K_p.

What Is Given?
You know the balanced equation for the reaction.
The partial pressures are:

$P_{H_2O} = 0.63$ atm

$P_{H_2} = 1.15$ atm

$P_{O_2} = 2.67$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$K_p = \dfrac{P_{H_2}^2 \, P_{O_2}}{P_{H_2O}^2}$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, into the expression and solve for K_p	$K_p = \dfrac{P_{H_2}^2 \, P_{O_2}}{P_{H_2O}^2}$ $= \dfrac{(1.15)^2 (2.67)}{(0.63)^2}$ $= 8.8966$ $= 8.9$ The equilibrium constant is 8.9.

Check Your Solution
The expression for K_p shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

42. What is the numerical value of the equilibrium constant, K_p, for the system below?

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
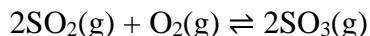
$P_{SO_2} = 0.54$ atm, $P_{O_2} = 0.58$ atm, and $P_{SO_3} = 2.67$ atm

What Is Required?
You need to determine the value of the equilibrium constant, K_p.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the partial pressure of each gas:

$P_{SO_2} = 0.54$ atm

$P_{O_2} = 0.58$ atm

$P_{SO_3} = 2.67$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$K_p = \dfrac{P_{SO_3}^2}{P_{SO_2}^2 \, P_{O_2}}$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, into the expression and solve for K_p.	$K_p = \dfrac{P_{SO_3}^2}{P_{SO_2}^2 \, P_{O_2}}$ $= \dfrac{(2.67)^2}{(0.54)^2(0.58)}$ $= 42.1509$ $= 42$ The equilibrium constant is 42.

Check Your Solution
The expression for K_p shows the product in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

43. A system is represented by the following reaction:

$$2HBr(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + Br_2(g)$$
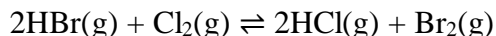
At 500 K, the partial pressures of this system at equilibrium were found to be as follows:
$P_{HBr} = 1.74$ atm, $P_{Cl_2} = 1.06$ atm, $P_{HCl} = 2.48$ atm, and $P_{Br_2} = 1.95$ atm
Calculate the numerical value of the equilibrium constant, K_p, at this temperature.

What Is Required?
You need to calculate the numerical value of the equilibrium constant, K_p.

What Is Given?
You know the balanced equation for the reaction.
You know the partial pressures of the gases:

$P_{HBr} = 1.74$ atm

$P_{Cl_2} = 1.06$ atm

$P_{HCl} = 2.48$ atm

$P_{Br_2} = 1.95$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$K_p = \dfrac{P_{HCl}^2 P_{Br_2}}{P_{HBr}^2 P_{Cl_2}}$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, into the expression and solve for K_p.	$K_p = \dfrac{P_{HCl}^2 P_{Br_2}}{P_{HBr}^2 P_{Cl_2}}$ $= \dfrac{(2.48)^2(1.95)}{(1.74)^2(1.06)}$ $= 3.7371$ $= 3.74$ The equilibrium constant is 3.74.

Check Your Solution
The expression for K_p shows the products in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

44. Hydrogen sulfide, $H_2S(g)$, is an extremely poisonous gas that has a characteristic foul smell similar to that of rotten eggs. It is used in many chemical applications, including the process of separation of heavy water from a sample of water. Heavy water is used to help control the release of energy in a CANDU nuclear reactor. The synthesis of hydrogen sulfide gas involves an equilibrium system in which sulfur gas, $S_2(g)$, reacts with hydrogen gas, $H_2(g)$, according to the following equation:

$$S_2(g) + 2H_2(g) \rightleftharpoons 2\,H_2S(g)$$
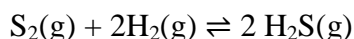
At a given temperature, the following partial pressures were found to exist in an equilibrium mixture of the gases: $P_{S_2} = 1.37$ atm, $P_{H_2} = 0.88$ atm, $P_{H_2S} = 1.68$ atm,

What is the numerical value of the equilibrium constant, K_p, at this temperature?

What Is Required?
You need to calculate the numerical value of the equilibrium constant, K_p,

What Is Given?
You know the balanced chemical equation for the reaction.
You know the partial pressures of the gases:

$P_{S_2} = 1.37$ atm

$P_{H_2} = 0.88$ atm

$P_{H_2S} = 1.68$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$K_p = \dfrac{P_{H_2S}^2}{P_{S_2}\,P_{H_2}^2}$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, into the expression and solve for K_p.	$K_p = \dfrac{P_{H_2S}^2}{P_{S_2}\,P_{H_2}^2}$ $= \dfrac{(1.68)^2}{(1.37)(0.88)^2}$ $= 2.6603$ $= 2.7$ The equilibrium constant is 2.7.

Check Your Solution
The expression for K_p shows the product in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

45. A system consisting of 18% carbon dioxide gas, 33% carbon monoxide gas, and oxygen gas has a total pressure of 5.45 atm. Determine the K_p value for this system:

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$$

What Is Required?
You need to determine the K_p for the system.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the total pressure: 5.45 atm
You know the mixture is 18% carbon dioxide gas, 33% carbon monoxide gas, and the remainder is oxygen gas.

Plan Your Strategy	Act on Your Strategy
Determine the percentage $O_2(g)$ in the mixture.	percentage $O_2(g) = 100\% - (18\% + 33\%)$ $\qquad\qquad\quad = 49\%$
The partial pressure exerted by each gas is directly proportional to its percentage in the mixture. Calculate the partial pressure of each gas.	$P_{CO_2} = 0.18 \times P_T$ $\qquad = 0.18 \times 5.45$ atm $\qquad = 0.981$ atm $P_{CO} = 0.33 \times P_T$ $\qquad = 0.33 \times 5.45$ atm $\qquad = 1.799$ atm $P_{O_2} = 0.49 \times P_T$ $\qquad = 0.49 \times 5.45$ atm $\qquad = 2.670$ atm
Write the expression for the equilibrium constant, K_p	$K_p = \dfrac{P_{CO}^2 P_{O_2}}{P_{CO_2}^2}$

The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, into the expression and solve for K_p.	$K_p = \dfrac{P_{CO}^2 P_{O_2}}{P_{CO_2}^2}$ $= \dfrac{(1.799)^2(2.670)}{(0.981)^2}$ $= 8.9791$ $= 8.99$ The equilibrium constant is 8.99.

Check Your Solution
The expression for K_p shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

46. A system of hydrogen chloride gas, HCl(g), hydrogen gas, H_2(g), and chlorine gas, Cl_2(g), has a total pressure of 5.88 atm. It is determined that the pressure composition is 30% H_2(g), 28% Cl_2(g), and the remaining pressure exerted by HCl(g). Determine the K_p value for the following reaction:
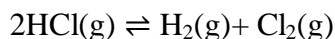
$$2HCl(g) \rightleftharpoons H_2(g) + Cl_2(g)$$

What Is Required?
You need to determine the K_p for the system.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the total pressure of the system: 5.88 atm
You know that the mixture is 30% hydrogen gas, 28% chlorine gas, and the remainder is hydrogen chloride gas.

Plan Your Strategy	Act on Your Strategy
Determine the percentage HCl(g) in the mixture.	percentage HCl(g) = 100% − (30% + 28%) = 42%
The partial pressure exerted by each gas is directly proportional to its percentage in the mixture. Calculate the partial pressure of each gas.	$P_{H_2} = 0.30 \times P_T$ $= 0.30 \times 5.88$ atm $= 1.764$ atm $P_{Cl_2} = 0.28 \times P_T$ $= 0.28 \times 5.88$ atm $= 1.646$ atm $P_{HCl} = 0.42 \times P_T$ $= 0.42 \times 5.88$ atm $= 2.470$ atm
Write the expression for the equilibrium constant, K_p.	$K_p = \dfrac{P_{H_2} P_{Cl_2}}{P_{HCl}^2}$

The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, into the expression and solve for K_p.	$K_p = \dfrac{P_{H_2} P_{Cl_2}}{P_{HCl}^2}$ $= \dfrac{(1.764)(1.646)}{(2.470)^2}$ $= 0.4759$ $= 0.48$ The equilibrium constant is 0.48.

Check Your Solution

The expression for K_p shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

47. At a given temperature, the system $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ is at equilibrium with $K_p = 4.15$. The system has the following partial pressures: $P_{N_2} = 1.23$ atm and $P_{NO} = 5.35$ atm. Determine the partial pressure of the oxygen gas present in this equilibrium mixture.

What Is Required?
You need to calculate the partial pressure of oxygen gas.

What Is Given?
You know the balanced equation for the reaction.
You know the value of the equilibrium constant, K_p: 4.15
You know the partial pressure of nitrogen and nitrous oxide:

$P_{N_2} = 1.23$ atm.

$P_{NO} = 5.35$ atm.

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$$K_p = \frac{P_{NO}^2}{P_{N_2} P_{O_2}}$$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, and the value for K_p into the expression and solve for P_{O_2}.	$$K_p = \frac{P_{NO}^2}{P_{N_2} P_{O_2}}$$ $$4.15 = \frac{(5.35)^2}{P_{O_2}(1.23)}$$ $$P_{O_2} = 5.61 \text{ atm}$$ The partial pressure of oxygen is 5.61 atm.

Check Your Solution
The expression for K_p shows the product in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium and the K_p have been substituted correctly. The answer has the correct number of significant digits and seems reasonable

48. At 435 K, $K_p = 21.3$ for the following reaction:

$$H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g)$$
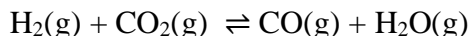
The system is analyzed and found to have the following partial pressures:
$P_{H_2} = 0.84$ atm, $P_{CO_2} = 1.31$ atm, and $P_{CO} = 5.35$ atm. Determine the partial pressure of the water vapour present in this equilibrium mixture.

What Is Required?
You need to determine the partial pressure of the water vapour present in this equilibrium mixture.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the equilibrium constant, K_p: 21.3
You know the partial pressures of the system:
$P_{H_2} = 0.84$ atm
$P_{CO_2} = 1.31$ atm
$P_{CO} = 5.35$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$K_p = \dfrac{P_{CO} P_{H_2O}}{P_{H_2} P_{CO_2}}$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, and the value for K_p into the expression and solve for P_{H_2O}.	$K_p = \dfrac{P_{CO} P_{H_2O}}{P_{H_2} P_{CO_2}}$ $21.3 = \dfrac{(5.35)(P_{H_2O})}{(0.84)(1.31)}$ $P_{H_2O} = 4.380$ atm $\phantom{P_{H_2O}} = 4.4$ atm The partial pressure of the water vapour is 4.4 atm.

Check Your Solution
The expression for K_p shows the products in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium and the K_p have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

49. At 1000 K, the following reaction occurs:

$$NO(g) + SO_3(g) \rightleftharpoons SO_2(g) + NO_2(g)$$
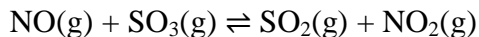
When this system reaches equilibrium at 1000 K, $K_p = 7.18$. A container of these gases at 1000 K is analyzed and found to have the following partial pressures: $P_{SO_3} = 0.32$ atm, $P_{SO_2} = 2.39$ atm, and $P_{NO_2} = 2.24$ atm. Determine the partial pressure of the nitrogen monoxide gas present in this system.

What Is Required?
You need to determine the partial pressure of the nitrogen monoxide gas present in this system.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the equilibrium constant, K_p: 7.18
You know the partial pressures of the system:
$\quad P_{SO_3} = 0.32$ atm
$\quad P_{SO_2} = 2.39$ atm
$\quad P_{NO_2} = 2.24$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$$K_p = \frac{P_{SO_2} P_{NO_2}}{P_{NO} P_{SO_3}}$$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, and the value for K_p into the expression and solve for P_{NO}.	$$K_p = \frac{P_{SO_2} P_{NO_2}}{P_{NO} P_{SO_3}}$$ $$7.18 = \frac{(2.39)(2.24)}{P_{NO}(0.32)}$$ $$P_{NO} = 2.330$$ $$= 2.3 \text{ atm}$$ The partial pressure of nitrogen monoxide is 2.3 atm.

Check Your Solution
The expression for K_p shows the products in the numerator, the reactants in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium and the K_p have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

50. At a given temperature, $K_p = 1.54$ for the following reaction:

$$2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$$

The system is analyzed and found to have the following partial pressures: $P_{NO} = 1.13$ atm and $P_{O_2} = 2.31$ atm. Determine the partial pressure of the nitrogen dioxide gas present in this equilibrium mixture at this temperature.

What Is Required?
You need to determine the partial pressure of the nitrogen dioxide gas present in this equilibrium mixture.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the equilibrium constant, K_p: 1.54
You know two of the partial pressures of the system:
$P_{O_2} = 2.31$ atm
$P_{NO} = 1.13$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the equilibrium constant, K_p.	$$K_p = \frac{P_{NO}^2 \, P_{O_2}}{P_{NO_2}^2}$$
The partial pressures in atm are divided by the reference state of 1 atm. Substitute the given partial pressures of the gases, without units, and the value for K_p into the expression and solve for P_{NO_2}.	$$K_p = \frac{P_{NO}^2 \, P_{O_2}}{P_{NO_2}^2}$$ $$1.54 = \frac{P_{NO}^2 \, P_{O_2}}{P_{NO_2}^2}$$ $$= \frac{(1.13)^2 (2.31)}{P_{NO_2}^2}$$ $$P_{NO_2} = 1.3839$$ $$= 1.38 \text{ atm.}$$ The partial pressure of nitrogen dioxide is 1.38 atm.

Check Your Solution
The expression for K_p shows the products in the numerator, the reactant in the denominator, and each is raised to the power of its coefficient in the balanced equation. The partial pressures at equilibrium and the K_p have been substituted correctly. The answer has the correct number of significant digits and seems reasonable.

Using Stoichiometry and K_{eq} to Calculate Equilibrium Concentrations (Student textbook page 451)

51. At 1100 K, hydrogen, $H_2(g)$, and iodine, $I_2(g)$, combine to form hydrogen iodide:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$
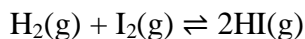
At equilibrium in a 1.0 L reaction vessel, the mixture of gases contained 0.30 mol of hydrogen, 1.3 mol of iodine, and 3.4 mol of hydrogen iodide. Determine the value of K_{eq}.

What Is Required?
You need to determine the value of K_{eq}.

What Is Given?
You know the balanced chemical equation for the equilibrium reaction.
You know the volume of the reaction vessel.
You know the initial amount in moles of each gas:

$n_{H_2} = 0.30$ mol

$n_{I_2} = 1.3$ mol

$n_{HI} = 3.4$ mol

Plan Your Strategy	Act on Your Strategy
Express the equilibrium concentration of the components in mol/L.	$[H_2] = 0.30$ mol/L $[I_2] = 1.3$ mol/L $[HI] = 3.4$ mol/L
Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$
Substitute the equilibrium concentrations into the expression and solve for K_{eq}.	$K_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$ $= \dfrac{(3.4)^2}{(0.30)(1.3)}$ $= 29.641$ $= 3.0 \times 10^1$ The equilibrium constant is 3.0×10^1.

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the numerator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. The answer is reasonable and has the correct number of significant digits.

52. At 25°C, the following reaction takes place:

$$I_2(g) + Cl_2(g) \rightleftharpoons 2ICl(g)$$
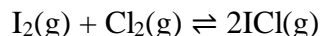
K_{eq} for the reaction is 82. If 0.83 mol of iodine gas, $I_2(g)$, and 0.83 mol of chlorine gas, $Cl_2(g)$, are placed in a 10 L container at 25°C, what are the concentrations of the various gases at equilibrium?

What Is Required?
You need to determine the equilibrium concentrations of each component.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 82
You know the volume of the container: 10 L
You know the initial amount in moles of the iodine gas and the chlorine gas:

$$n_{I_2} = 0.83 \text{ mol}$$

$$n_{Cl_2} = 0.83 \text{ mol}$$

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentration of the components in mol/L	$[I_2] = \dfrac{0.83 \text{ mol}}{10 \text{ L}}$ $= 0.083$ mol/L $[Cl_2] = \dfrac{0.83 \text{ mol}}{10 \text{ L}}$ $= 0.083$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the decrease in concentration in mol/L of $I_2(g)$ when the system reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$I_2(g)$	+	$Cl_2(g)$	$\rightleftharpoons$	$2ICl(g)$
	$[I_2]$ (mol/L)		$[Cl_2]$ (mol/L)		$[ICl]$ (mol/L)
I	0.083		0.083		0
C	$-x$		$-x$		$+2x$
E	$0.083 - x$		$0.083 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq}.	$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$
Substitute the equilibrium concentrations into the expression for K_{eq} and solve for x.	$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$ $$82 = \frac{(2x)^2}{(0.083-x)(0.083-x)}$$ $$\sqrt{82} = \sqrt{\frac{(2x)^2}{(0.083-x)(0.083-x)}}$$ $$9.05538 = \frac{2x}{0.083-x}$$ $$x = 0.06798 \text{ mol/L}$$ $$= 0.068 \text{ mol/L}$$
Determine the equilibrium concentrations of each component.	$[H_2] = [I_2]$ $= 0.083 \text{ mol/L} - 0.068 \text{ mol/L}$ $= 0.015 \text{ mol/L}$ $[HI] = 2(0.068 \text{ mol/L})$ $= 0.136 \text{ mol/L}$ $= 0.14 \text{ mol/L}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 87$. Considering the rounding for the final answers, this approximates the given K_{eq}. The answers are reasonable.

53. A chemist is studying the following reaction at a certain temperature:

$$SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$$
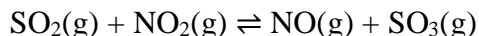
In a 1.0 L container, the chemist adds 0.17 mol of sulfur dioxide, $SO_2(g)$, to 0.11 mol of nitrogen dioxide, $NO_2(g)$. At equilibrium, the concentration of sulfur trioxide, $SO_3(g)$, is 0.089 mol/L. What is the value of K_{eq} for the reaction at this temperature?

What Is Required?
You need to determine the value of K_{eq} for the reaction.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 1.0 L
You know the initial amounts in moles of sulfur dioxide and nitrogen dioxide:

$n_{SO_2} = 0.17$ mol

$n_{NO_2} = 0.11$ mol

You know the concentration of sulfur trioxide at equilibrium: $[SO_3(g)] = 0.089$ mol/L

Plan Your Strategy	Act on Your Strategy
Express the initial concentrations of $SO_2(g)$ and $NO_2(g)$ in mol/L.	$[SO_2(g)] = 0.17$ mol/L $[NO_2(g)] = 0.11$ mol/L
Set up an ICE table showing the initial molar concentrations of each component. Use the stoichiometry of the reaction and the given equilibrium concentration of $SO_3(g)$ to determine the equilibrium concentrations of each component.	See the ICE table below.

	$SO_2(g)$	+	$NO_2(g)$	$\rightleftharpoons$	$NO(g)$	+	$SO_3(g)$
	$[SO_2]$ (mol/L)		$[NO_2]$ (mol/L)		$[NO]$ (mol/L)		$[SO_2]$ (mol/L)
I	0.170		0.110		0		0
C	0.170 – 0.089		0.110 – 0.089		0.089		0.089
E	0.081		0.021		0.089		0.089

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[NO][SO_3]}{[SO_2][NO_2]}$

Substitute the equilibrium concentrations into the expression and solve for K_{eq}.	$K_{eq} = \dfrac{[NO][SO_3]}{[SO_2][NO_2]}$ $= \dfrac{(0.089)(0.089)}{(0.081)(0.021)}$ $= 4.6566$ $= 4.7$ The equilibrium constant for the reaction is 4.7.

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. The answer is reasonable and has the correct number of significant digits.

54. Hydrogen bromide, HBr(g), decomposes at 700 K, as represented by the following reaction:

$$2HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$$
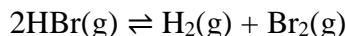
K_{eq} is 4.2×10^{-9} for this reaction. If 0.090 mol of hydrogen bromide is placed into a 2.0 L reaction vessel and heated to 700 K, what is the equilibrium concentration of each gas?

What Is Required?
You need to calculate the equilibrium concentration of each gas?

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 4.2×10^{-9}
You know the volume, V, of the reaction vessel: 2.0 L
You know the initial amount in moles of hydrogen bromide: $n_{HBr} = 0.090$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentration , c, of HBr in mol/L.	$[HBr] = \dfrac{0.090 \text{ mol}}{2.0 \text{ L}}$ $= 0.045$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the concentration, in mol/L, of $Br_2(g)$ when the system reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	2HBr(g) $\rightleftharpoons$	$H_2(g)$ +	$Br_2(g)$
	[HBr](mol/L)	$[H_2]$(mol/L)	$[Br_2]$(mol/L)
I	0.045	0	0
C	$-2x$	$+x$	$+x$
E	$0.045 - 2x$	x	x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2][Br_2]}{[HBr]^2}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$K_{eq} = \dfrac{[H_2][Br_2]}{[HBr]^2}$ $4.2 \times 10^{-9} = \dfrac{(x)(x)}{(0.045 - 2x)^2}$ $\sqrt{4.2 \times 10^{-9}} = \sqrt{\dfrac{(x)^2}{(0.045 - 2x)^2}}$ $6.4807 \times 10^{-5} = \dfrac{(x)}{0.045 - 2x}$ $x = 2.9 \times 10^{-6}\,\text{mol/L}$
Determine the equilibrium concentration of each component.	$[H_2] = [Br_2]$ $\quad = 2.9 \times 10^{-6}$ mol/L $[HBr] = 0.045 - 2(2.9 \times 10^{-6}$ mol/L$)$ $\quad = 0.045$ mol/L

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expressiongives $K_{eq} = 4.2 \times 10^{-9}$. This closely approximates the given K_{eq}. The answers are reasonable.

55. For the reaction of phosphorus trichloride, PCl_3, and nitrogen dioxide, NO_2, an equilibrium is established according to the following equation:

$$PCl_3(g) + NO_2(g) \rightleftharpoons POCl_3(g) + NO(g)$$
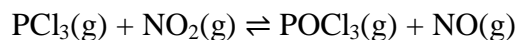
Initially, 1.24 mol of each reactant is placed into a 1.0 L container and equilibrium is established. If $K_{eq} = 3.77$ for the reaction, what are the equilibrium concentrations of all reactants and products?

What Is Required?
You need to calculate the equilibrium concentrations of all reactants and products.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the initial concentration of each reactant in the equilibrium mixture: 1.24 mol/L
You know the value of the equilibrium constant, K_{eq}: 3.77

Plan Your Strategy	Act on Your Strategy
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $PCl_3(g)$ had decreased when the system reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$PCl_3(g)$	+	$NO_2(g)$	$\rightleftharpoons$	$POCl_3(g)$	+	$NO(g)$
	$[PCl_3]$(mol/L)		$[NO_2]$(mol/L)		$[POCl_3]$(mol/L)		$[NO]$(mol/L)
I	1.24		1.24		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$1.24 - x$		$1.24 - x$		x		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[POCl_3][NO]}{[PCl_3][NO_2]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$$K_{eq} = \frac{[POCl_3][NO]}{[PCl_3][NO_2]}$$ $$3.77 = \frac{(x)(x)}{(1.24-x)(1.24-x)}$$ $$\sqrt{3.77} = \sqrt{\frac{(x)(x)}{(1.24-x)(1.24-x)}}$$ $$1.9416 = \frac{x}{1.24-x}$$ $$x = 0.818$$
Determine the equilibrium concentration of each component.	$[POCl_3]$ = $[NO]$ $\quad$ = 0.82 mol/L $[PCl_3]$ = $[NO_2]$ $\quad$ = 1.24 mol/L – 0.82 mol/L $\quad$ = 0.42 mol/L

Check Your Solution
The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives K_{eq} = 3.81. This closely approximates the given K_{eq}. The answers are reasonable.

56. When nitrogen gas reacts with oxygen gas, nitrogen monoxide gas forms:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$
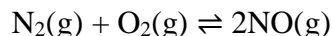
Initially, 0.30 mol of each reactant is placed into a 2.0 L container and equilibrium is established. If $K_{eq} = 52.1$ for the reaction, what are the equilibrium concentrations of all reactants and products?

What Is Required?
You need to determine the equilibrium concentrations of the reactants and products.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 52.1
You know the volume, V, of the container: 2.0 L
You know the initial amount in moles of each of $N_2(g)$ and $O_2(g)$: $n = 0.30$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentrations, c, of $N_2(g)$ and $O_2(g)$ in mol/L.	$[N_2] = [O_2]$ $= \dfrac{0.30 \text{ mol}}{2.0 \text{ L}}$ $= 0.15$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $N_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$N_2(g)$ +	$O_2(g)$ $\rightleftharpoons$	$2NO(g)$
	$[N_2]$ (mol/L)	$[O_2]$ (mol/L)	$[NO]$ (mol/L)
I	0.15	0.15	0
C	$-x$	$-x$	$+2x$
E	$0.15 - x$	$0.15 - x$	$2x$

Write the expression for the equilibrium constant, K_{eq}.	$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$ $$52.1 = \frac{(2x)^2}{(0.15-x)(015-x)}$$ $$\sqrt{52.1} = \sqrt{\frac{(2x)^2}{(0.15-x)(015-x)}}$$ $$7.2180 = \frac{2x}{(0.15-x)}$$ $$x = 0.11745 \text{ mol/L}$$
Determine the equilibrium concentration of each component.	$[NO] = 2x$ $\quad = 2(0.11745 \text{ mol/L})$ $\quad = 0.2349 \text{ mol/L}$ $\quad = 0.23 \text{ mol/L}$ $[N_2] = [O_2]$ $\quad = 0.15 - 0.1145$ $\quad = 3.3 \times 10^{-2} \text{ mol/L}$

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 48.6$. This approximates the given $K_{eq.}$ The answers are reasonable.

57. In the reaction of hydrogen gas reacting with iodine gas to form hydrogen iodide gas, $K_{eq} = 49.6$ at 730 K. The reaction is as follows:

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$
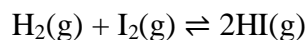
Initially, 3.51 mol of each reactant is placed into a 3.0 L container and equilibrium is established. Determine the amount in moles of each substance that will be present when the system reaches equilibrium at 730 K.

What Is Required?
You need to determine the amount in moles of each substance at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 49.6
You know the volume, V, of the container: 3.0 L
You know the initial amount in moles of each of $H_2(g)$ and $I_2(g)$: $n = 3.51$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentrations, c, in mol/L.	$[H_2] = [I_2]$ $= \dfrac{3.51 \text{ mol}}{3.0 \text{ L}}$ $= 1.17$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $H_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$H_2(g)$ +	$I_2(g)$ $\rightleftharpoons$	$2HI(g)$
	$[H_2]$ (mol/L)	$[I_2]$ (mol/L)	$[HI]$ (mol/L)
I	1.17	1.17	0
C	$-x$	$-x$	$+2x$
E	$1.17 - x$	$1.17 - x$	$2x$

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$ $$49.6 = \frac{(2x)^2}{(1.17-x)(1.17-x)}$$ $$\sqrt{49.6} = \sqrt{\frac{(2x)^2}{(1.17-x)(1.17-x)}}$$ $$7.0427 = \frac{2x}{1.17-x}$$ $$x = 0.91123 \text{ mol/L}$$ $$= 0.911 \text{ mol/L}$$
Determine the equilibrium concentration of each component.	$[H_2] = [I_2]$ $= 1.17\text{mol/L} - 0.911 \text{ mol/L}$ $= 0.259 \text{ mol/L}$ $= 0.26 \text{ mol/L}$ $[HI] = 2(0.911 \text{ mol/L })$ $= 1.822 \text{ mol/L}$ $= 1.8 \text{ mol/L}$
Use the formula $n = cV$ to determine the amount in moles, n, of each component.	$n_{H_2} = cV$ $= (0.26 \text{ mol/}\cancel{L})(3.0 \ \cancel{L})$ $= 0.78 \text{ mol}$ $n_{I_2} = cV$ $= (0.26 \text{ mol/}\cancel{L})(3.0 \ \cancel{L})$ $= 0.78 \text{ mol}$ $n_{HI} = cV$ $= (1.8 \text{ mol/}\cancel{L})(3.0 \ \cancel{L})$ $= 5.4 \text{ mol}$

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression, gives $K_{eq} = 47.9$. This approximates the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

58. Initially, 2.25 mol of carbon monoxide gas, $CO(g)$, and 2.25 mol of water vapour, $H_2O(g)$, are placed into a 1.50 L container and the following equilibrium is established, with a K_{eq} value of 4.2:
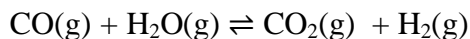
$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

What is the amount in moles of each reactant and product at equilibrium?

What Is Required?
You need to determine the amount in moles of each substance at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq},: 4.2
You know the volume, V, of the container = 1.50 L
You know initial amount in moles of $CO(g)$ and $H_2O(g)$: $n = 2.25$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentrations, c, in mol/L.	$[CO] = [H_2O]$ $= \dfrac{2.25 \text{ mol}}{1.50 \text{ L}}$ $= 1.50 \text{ mol/L}$
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $CO(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$CO(g)$	+	$H_2O(g)$	$\rightleftharpoons$	$CO_2(g)$	+	$H_2(g)$
	[CO] (mol/L)		[H₂O] (mol/L)		[CO₂] (mol/L)		[H₂] (mol/L)
I	1.50		1.50		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$1.50 - x$		$1.50 - x$		x		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$K_{eq} = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$ $4.2 = \dfrac{(x)(x)}{(1.50-x)(1.5-x)}$ $\sqrt{4.2} = \sqrt{\dfrac{(x)(x)}{(1.50-x)(1.5-x)}}$ $2.0494 = \dfrac{x}{1.50-x}$ $x = 1.0081 \text{ mol/L}$
Determine the equilibrium concentration of each component.	$[CO] = [H_2O]$ $\quad = 1.50 \text{ mol/L} - 1.0081 \text{ mol/L}$ $\quad = 0.4919 \text{ mol/L}$ $[CO_2] = [H_2]$ $\quad = 1.0081 \text{ mol/L}$
Use the formula $n = cV$ to determine the amount in moles, n, of each component.	$n_{CO} = cV$ $\quad = (0.4919 \text{ mol/}\cancel{L})(1.5 \ \cancel{L})$ $\quad = 0.73785 \text{ mol}$ $\quad = 0.74 \text{ mol}$ $n_{H_2O} = n_{CO}$ $\quad = 0.74 \text{ mol}$ $n_{CO_2} = cV$ $\quad = (1.0081 \text{ mol/}\cancel{L})(1.5 \ \cancel{L})$ $\quad = 1.5125 \text{ mol}$ $\quad = 1.5 \text{ mol}$ $n_{H_2} = n_{CO_2}$ $\quad = 1.5 \text{ mol}$

Check Your Solution
The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 4.2$. This agrees with the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

59. Bromine gas and chlorine gas establish equilibrium when they react to form bromine chloride gas. The K_{eq} for this reaction at 450 K is 28.8.

$$Br_2(g) + Cl_2(g) \rightleftharpoons 2BrCl(g)$$
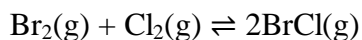
When 8.70 mol of each reactant is placed into a 5.00 L container, what is the amount in moles of each reactant and product at equilibrium?

What Is Required?
You need to determine the amount in moles of each substance at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 28.8
You know the volume, V, of the container: 5.00 L
You know the initial amount in moles of $Br_2(g)$ and $(Cl_2(g)$: $n = 8.70$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentrations, c, in mol/L.	$[Br_2] = [Cl_2]$ $= \dfrac{8.70 \text{ mol}}{5.00 \text{ L}}$ $= 1.740$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent amount by which the concentration of $Br_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$Br_2(g)$	+	$Cl_2(g)$	$\rightleftharpoons$	$2BrCl(g)$
	$[Br_2]$ (mol/L)		$[Cl_2]$ (mol/L)		$[BrCl]$ (mol/L)
I	1.740		1.740		0
C	$-x$		$-x$		$2x$
E					

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[BrCl]^2}{[Br_2][Cl_2]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$K_{eq} = \dfrac{[BrCl]^2}{[Br_2][Cl_2]}$ $28.8 = \dfrac{(2x)^2}{(1.740 - x)(1.740 - x)}$ $\sqrt{28.8} = \sqrt{\dfrac{(2x)^2}{(1.740 - x)(1.740 - x)}}$ $5.3665 = \dfrac{2x}{1.740 - x}$ $x = 1.2676$ mol/L
Determine the equilibrium concentration of each component.	$[Br_2] = [Cl_2]$ $= 1.740$ mol/L $- 1.2676$ mol/L $= 0.4724$ mol/L $= 0.472$ mol/L $[BrCl] = 2(1.2676$ mol/L$)$ $= 2.5352$ mol/L
Use the formula $n = cV$ to determine the amount in moles, n, of each component.	$n_{Br_2} = cV$ $= (0.472$ mol/L̸$)(5.00$ L̸$)$ $= 2.36$ mol $n_{Cl_2} = n_{Br_2}$ $= 2.36$ mol $n_{BrCl} = cV$ $= (2.5352$ mol/L̸$)(5.00$ L̸$)$ $= 12.676$ mol $= 12.7$ mol

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 28.9$. This approximates the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

60. Ozone, $O_3(g)$, reacts with nitrogen monoxide, $NO(g)$, and forms nitrogen dioxide $NO_2(g)$, and oxygen, $O_2(g)$. At 325 K, K_{eq} for the reaction is 6.70. The reaction is:

$$O_3(g) + NO(g) \rightleftharpoons NO_2(g) + O_2(g)$$
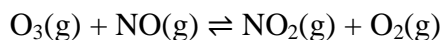
If 1.62 mol of each reactant is placed into a 2.00 L container, what is the amount in moles of each reactant and product at equilibrium?

What Is Required?
You need to determine the amount in moles of each substance at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 6.70
You know the volume, V, of the container: 2.00 L
You know the initial amount in moles of $O_3(g)$ and $NO(g)$: $n = 1.62$ mol

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to express the initial concentrations, c, in mol/L.	$[O_3] = [NO]$ $= \dfrac{1.62 \text{ mol}}{2.00 \text{ L}}$ $= 0.810$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the the amount by which the concentration of $O_3(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$O_3(g)$	+	$NO(g)$	$\rightleftharpoons$	$NO_2(g)$	+	$O_2(g)$
	$[O_3]$ (mol/L)		$[NO]$ (mol/L)		$[NO_2]$ (mol/L)		$[O_2]$ (mol/L)
I	0.810		0.810		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$0.810 - x$		$0.810 - x$		x		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[NO_2][O_2]}{[O_3][NO]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve for x.	$$K_{eq} = \frac{[NO_2][O_2]}{[O_3][NO]}$$ $$6.70 = \frac{(x)(x)}{(0.810-x)(0.810-x)}$$ $$\sqrt{6.70} = \sqrt{\frac{(x)(x)}{(0.810-x)(0.810-x)}}$$ $$2.5884 = \frac{x}{0.810-x}$$ $$x = 0.58427 \text{ mol/L}$$ $$= 0.584 \text{ mol/L}$$
Determine the equilibrium concentration of each component.	$[O_3] = [NO]$ $= 0.810 \text{ mol/L} - 0.584 \text{mol/L}$ $= 0.226 \text{ mol/L}$ $[NO_2] = [O_2]$ $= 0.584 \text{ mol/L}$
Use the formula $n = cV$ to determine the amount in moles, n, of each component.	$n_{O_3} = cV$ $= (0.226 \text{ mol/} \cancel{L})(2.00 \ \cancel{L})$ $= 0.452 \text{ mol}$ $n_{NO} = n_{O_3}$ $= 0.452 \text{ mol}$ $n_{NO_2} = cV$ $= (0.584 \text{ mol/} \cancel{L})(2.00 \ \cancel{L})$ $= 1.168 \text{ mol}$ $= 1.17 \text{ mol}$ $n_{O_2} = n_{NO_2}$ $= 1.17 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 6.67$. This approximates the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

Using the Quadratic Equation to Determine Equilibrium Information
(Student textbook page 454)

61. When solid carbon reacts with water vapour, hydrogen gas and carbon dioxide gas form according to the following equation:

$$C(s) + H_2O(g) \rightleftharpoons H_2(g) + CO(g)$$
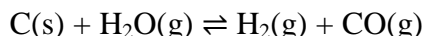
When 2.25 mol of water vapour is reacted with solid carbon in a 1.0 L container, and equilibrium is established, what concentration of each reactant and product in the gas phase will be present at equilibrium if K_{eq} equals 23.4?

What Is Required?
You need to calculate the concentration of each reactant and product in the gas phase at equilibrium.

What Is Given?
You know the balanced equation for the reaction and the state of each component in the reaction.
You know the volume of the container: 1.0 L
You know the initial amount in moles of $H_2O(g)$: $n = 2.25$ mol
You know the value of the equilibrium constant, K_{eq}: 23.4

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $H_2O(g)$ in mol/L.	$[H_2O] = 2.25$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration $H_2O(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	C(s)	+	$H_2O(g)$	$\rightleftharpoons$	$H_2(g)$	+	CO(g)
			$[H_2O]$ (mol/L)		$[H_2]$ (mol/L)		$[CO]$ (mol/L)
I			2.25		0		0
C			$-x$		$+x$		$+x$
E			$2.25 - x$		x		x

| Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2][CO]}{[H_2O]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[H_2][CO]}{[H_2O]}$ $23.4 = \dfrac{(x)(x)}{(2.25 - x)}$
Re-arrange the expression into a quadratic equation.	$x^2 + 23.4x - 52.65 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-23.4 \pm \sqrt{(23.4)^2 - (4)(-52.65)}}{2(1)}$ $= \dfrac{-23.4 \pm 27.535}{2}$ $x = 2.0674 \text{ or } x = -25.4674$ Because it is impossible to have a negative concentration, $x = 2.0674$.
Determine the equilibrium concentration of each component in the gas phase.	$[H_2O] = 2.25 \text{ mol/L} - 2.0674 \text{ mol/L}$ $= 0.1826 \text{ mol/L}$ $= 0.18 \text{ mol/L}$ $[H_2] = [CO]$ $= 2.0674$ $= 2.1 \text{ mol/L}$

Check Your Solution
The expression for K_{eq} shows the products in the numerator and the reactant in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 6.67$. This approximates the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

62. Dinitrogen tetroxide, $N_2O_4(g)$, reversibly decomposes into nitrogen dioxide, $NO_2(g)$, according to the following equation:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$
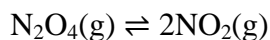
This reaction has resulted in one of the most important rocket propellant processes that researchers have developed. At 325 K, the K_{eq} value is 0.91. Initially, 0.34 mol $N_2O_4(g)$ is placed into a 1.00 L container and equilibrium is established. Determine the equilibrium concentration of each gas in the container.

What Is Required?
You need to determine the equilibrium concentration of each gas in the container.

What Is Given?
You know the balanced equation for the reaction and the state of each component in the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles of $N_2O_4(g)$: $n = 0.34$ mol
You know the value of the equilibrium constant, K_{eq}: 0.91

| Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $N_2O_4(g)$ in mol/L.	$[N_2O_4] = 0.34$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $N_2O_4(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	N_2O_4 (g) $\rightleftharpoons$	$2NO_2(g)$
	$[N_2O_4]$ (mol/L)	$[NO_2]$ (mol/L)
I	0.34	0
C	$-x$	$+2x$
E	$0.34 - x$	$2x$

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[NO_2]^2}{[N_2O_4]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$$ $$0.91 = \frac{(2x)^2}{(0.34 - x)}$$
Re-arrange the expression into a quadratic equation.	$4x^2 + 0.91x - 0.3094 = 0$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-0.91 \pm \sqrt{(0.91)^2 - (4)(4)(-0.3094)}}{2(4)}$$ $$= \frac{-0.91 \pm 2.4038}{2(4)}$$ $x = 0.18673$ or $x = -0.4142$ Because it is impossible to have a negative concentration, $x = 0.18673$.
Determine the equilibrium concentration of each component in the gas phase.	$[N_2O_4] = 0.34$ mol/L $- 0.18673$ mol/L $\quad = 0.15327$ mol/L $\quad = 0.15$ mol/L $[NO_2] = 2(0.18673)$ mol/L $\quad = 0.37346$ mol/L $\quad = 0.37$ mol/L

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactant in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 0.91$. This agrees with the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

63. In an equilibrium process, ethene, $C_2H_4(g)$, is reacted with hydrogen gas to produce ethane, $C_2H_6(g)$. At a given temperature, the K_{eq} value is 1.04 for the following reaction:

$$C_2H_4(g) + H_2(g) \rightleftharpoons C_2H_6(g)$$
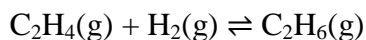
Initially, 0.34 mol $C_2H_4(g)$ and 0.53 mol $H_2(g)$ are placed into a 1.00 L container and equilibrium is established. Determine the equilibrium concentration of each gas in the container.

What Is Required?
You need to determine the equilibrium concentration of each gas in the container.

What Is Given?
You know the balanced equation for the reaction and the state of each component in the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles, n, of $C_2H_4(g)$: 0.34 mol
You know the initial amount in moles, n, of $H_2(g)$: 0.53 mol
You know the value of the equilibrium constant, K_{eq}: 1.04

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $C_2H_4(g)$ and $H_2(g)$ in mol/L.	$[N_2O_4] = 0.34$ mol/L $[H_2] = 0.53$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amont by which the concentration of $C_2H_4(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	$C_2H_4(g)$	+	$H_2(g)$	$\rightleftharpoons$	$C_2H_6(g)$
	$[C_2H_4]$ (mol/L)		$[H_2]$ (mol/L)		$[C_2H_6]$ (mol/L)
I	0.34		0.53		0
C	$-x$		$-x$		$+x$
E	$0.34 - x$		$0.53 - x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[C_2H_6]}{[C_2H_4][H_2]}$

Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$$K_{eq} = \frac{[C_2H_6]}{[C_2H_4][H_2]}$$ $$1.04 = \frac{x}{(0.34-x)(0.53-x)}$$
Re-arrange the expression into a quadratic equation.	$1.04(0.1802 - 0.87x + x^2) = x$ $1.04x^2 - 1.9048x + 0.1874 = 0$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-(-1.9048) \pm \sqrt{(-1.9048)^2 - (4)(1.04)(0.1874)}}{2(1.04)}$$ $$\frac{-(-1.9048) \pm 1.6878}{2.08}$$ $x = 0.1047$ or $x = 1.7272$ Because it is impossible to have a negative concentration, x cannot be greater than 0.34. Therefore, $x = 0.1047$.
Determine the equilibrium concentration of each component in the gas phase.	$[C_2H_4] = 0.34$ mol/L $- 0.1047$ mol/L $= 0.2353$ mol/L $= 0.24$ mol/L $[H_2] = 0.53$ mol/L $- 0.1047$ mol/L $= 0.4253$ mol/L $= 0.43$ mol/L $[C_2H_6] = 0.1047$ mol/L $= 0.10$ mol/L

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 0.97$. This approximates the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

64. When 1.88 mol of hydrogen gas and 2.86 mol of iodine gas are placed in a 2.00 L container, an equilibrium is reached with hydrogen iodide gas such that the K_{eq} value is 55.3 at 700 K. Determine the amount in moles of each reactant and product in the container when equilibrium is established.

What Is Required?

You need to determine the amount in moles of each reactant and product at equilibrium.

What Is Given?

You know the volume, V, of the container: 2.00 L
You know the initial amount in moles of $H_2(g)$: $n = 1.88$ mol
You know the initial amount in moles of $I_2(g)$: $n = 2.86$ mol
You know the value of the equilibrium constant, K_{eq}: 55.3

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the equilibrium reaction.	$H_2(g) \ + \ I_2(g) \ \rightleftharpoons \ 2HI(g)$
Use $c = \dfrac{n}{V}$ to express the initial concentration of $H_2(g)$ and $I_2(g)$ in mol/L.	$[H_2] = \dfrac{1.88 \text{ mol}}{2.00 \text{ L}}$ $\quad = 0.940$ mol/L $[I_2] = \dfrac{2.86 \text{ mol}}{2.00 \text{ L}}$ $\quad = 1.43$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $H_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	$2HI(g)$
	[H_2] (mol/L)		[I_2] (mol/L)		[HI] (mol/L)
I	0.940		1.43		0
C	$-x$		$-x$		$+2x$
E	$0.940 - x$		$1.43 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$ $55.3 = \dfrac{(2x)^2}{(0.940 - x)(1.43 - x)}$
Re-arrange the expression into a quadratic equation.	$55.3(1.3442 - 2.37x + x^2) = 4x^2$ $51.4x^2 - 131.061x + 74.3232 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-(-131.061) \pm \sqrt{(-131.061)^2 - (4)(51.3)(74.3232)}}{2(51.3)}$ $= \dfrac{131.061 \pm 43.885}{102.6}$ $x = 1.7051$ or $x = 0.8497$ Because it is impossible to have a negative concentration, x cannot be greater than 0.940. Therefore, $x = 0.8497$.
Determine the equilibrium concentration of each component in the gas phase.	$[H_2] = 0.940$ mol/L $- 0.8497$ mol/L $\qquad = 0.0903$ mol/L $[I_2] = 1.43$ mol/L $- 0.8497$ mol/L $\qquad = 0.5803$ mol/L $[HI] = 2(0.8497$ mol/L$)$ $\qquad = 1.6994$ mol/L
Use the formula $n = cV$ to calculate the amount in moles, n, of each component.	$n_{H_2} = cV$ $\qquad = (0.903$ mol/$\cancel{L})(2.00 \ \cancel{L})$ $\qquad = 0.181$ mol $n_{I_2} = cV$ $\qquad = (0.5803$ mol/$\cancel{L})(2.00 \ \cancel{L})$ $\qquad = 1.16$ mol $n_{HI} = cV$ $\qquad = (1.6994$ mol/$\cancel{L})(2.00 \ \cancel{L})$ $\qquad = 3.40$ mol

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression $K_{eq} = 55.1$. This is very close to the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

65. Carbonyl chloride gas, also called phosgene, $COCl_2(g)$, was used during World War I as a chemical weapon. Now, it is used in a process to manufacture plastics such as the ones used to make lenses in eyeglasses. It is formed by reacting carbon monoxide gas with chlorine gas in the following equilibrium system:

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$$
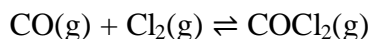
The K_{eq} for this reaction at 1000 K is 24.0. If 2.35 mol $CO(g)$ and 1.14 mol $Cl_2(g)$ are placed into a 1.00 L container, what concentration of each gas will exist at equilibrium at 1000 K?

What Is Required?
You need to determine the equilibrium concentration of each gas in the container.

What Is Given?
You know the balanced equation for the reaction.
You know the state of each component in the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles of $CO(g)$: $n = 2.35$ mol
You know the initial amount in moles of $Cl_2(g)$: $n = 1.14$ mol
You know the value of the equilibrium constant, K_{eq}: 24.0

| Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $CO(g)$ and $Cl_2(g)$ in mol/L.	[CO] = 2.35 mol/L [Cl$_2$] = 1.14 mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $CO(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	CO(g)	+	Cl₂(g)	⇌	COCl₂(g)
	[CO] (mol/L)		[Cl₂] (mol/L)		[COCl₂] (mol/L)
I	2.35		1.14		0
C	$-x$		$-x$		$+x$
E	$2.35 - x$		$1.14 - x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[COCl_2]}{[CO][Cl_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[COCl_2]}{[CO][Cl_2]}$ $24.0 = \dfrac{x}{(2.35 - x)(1.14 - x)}$
Re-arrange the expression into a quadratic equation.	$24.0(2.679 - 3.49x + x^2) = x$ $24.0x^2 - 84.76x + 64.296 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-(-84.76) \pm \sqrt{(-84.76)^2 - (4)(24.0)(64.296)}}{2(24.0)}$ $= \dfrac{84.76 \pm 31.809}{48.0}$ $x = 2.4285$ or $x = 1.103$ Because it is impossible to have a negative concentration, x cannot be greater than 1.14. Therefore, $x = 1.1031$.
Determine the equilibrium concentration of each component in the gas phase.	The equilibrium concentrations are: $[CO]$ = 2.35 mol/L – 1.1031 mol/L $\quad\quad$ = 1.2469 mol/L $\quad\quad$ = 1.25 mol/L $[Cl_2]$ = 1.14 mol/L – 1.1031 mol/L $\quad\quad$ = 0.0370 mol/L $[COCl_2]$ = 1.1031 mol/L $\quad\quad$ = 1.10 mol/L

Check Your Solution
The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 23.8$. This agrees closely with the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

66. Phosphoryl chloride, $POCl_3(g)$, is used in the manufacturing of flame retardants. It is manufactured in an equilibrium process in which phosphorus trichloride reacts with nitrogen dioxide to form $POCl_3(g)$ and $NO(g)$ according to the following equation:

$$PCl_3(g) + NO_2(g) \rightleftharpoons POCl_3(g) + NO(g)$$
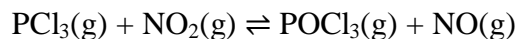
The K_{eq} for this reaction at 800 K is 1.82. If 1.86 mol $PCl_3(g)$ and 1.64 mol $NO_2(g)$ are placed into a 2.00 L container, what is the amount in moles of each gas at equilibrium at 800 K?

What Is Required?
You need to determine the amount in moles of each reactant and product at equilibrium.

What Is Given?
You know the volume of the container: 2.00 L
You know the initial amount in moles of $PCl_3(g)$: $n = 1.86$ mol
You know the initial amount in moles of NO_2 (g): $n = 1.64$ mol
You know the value of the equilibrium constant, K_{eq}: 1.82

| Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $PCl_3(g)$ and $NO_2(g)$ in mol/L.	$[PCl_3] = \dfrac{1.86 \text{ mol}}{2.00 \text{ L}}$ $= 0.930$ mol/L $[NO_2] = \dfrac{1.64 \text{ mol}}{2.00 \text{ L}}$ $= 0.820$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $PCl_3(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component in the gas phase.	See the ICE table below.

	PCl₃(g)	+	NO₂(g)	⇌	POCl₃(g)	+	NO(g)
	$[PCl_3]$ (mol/L)		$[NO_2]$ (mol/L)		$[POCl_3]$ (mol/L)		$[NO]$ (mol/L)
I	0.930		0.820		0		0
C	$-x$		$-x$		$+x$		$+x$
E	$0.930 - x$		$0.820 - x$		x		x

Write the expression for the equilibrium constant, K_{eq}.	$$K_{eq} = \frac{[POCl_3][NO]}{[PCl_3][NO_2]}$$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$$K_{eq} = \frac{[POCl_3][NO]}{[PCl_3][NO_2]}$$ $$1.82 = \frac{(x)(x)}{(0.930 - x)(0.820 - x)}$$
Re-arrange the expression into a quadratic equation.	$$1.82(0.7626 - 1.75x + x^2) = x^2$$ $$0.820x^2 - 3.185x + 1.3879 = 0$$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-(-3.185) \pm \sqrt{(-3.185)^2 - (4)(0.820)(1.3879)}}{2(0.820)}$$ $$= \frac{3.185 \pm 2.3647}{1.64}$$ $$x = 3.3836 \text{ or } x = 0.5005$$ Because it is impossible to have a negative concentration, x cannot be greater than 0.820. Therefore, $x = 0.5005$.
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[PCl_3] = 0.930 \text{ mol/L} - 0.5005 \text{ mol/L}$ $\qquad = 0.430 \text{ mol/L}$ $[NO_2] = 0.820 \text{ mol/L} - 0.5005 \text{ mol/L}$ $\qquad = 0.320 \text{ mol/L}$ $[POCl_3] = [NO]$ $\qquad = 0.5005 \text{ mol/L}$
Use the formula $n = cV$ to calculate the amount in moles, n, of each component.	$n_{PCl_3} = cV$ $\qquad = (0.430 \text{ mol/}\cancel{L})(2.00 \text{ }\cancel{L})$ $\qquad = 0.860 \text{ mol}$ $n_{NO_2} = cV$ $\qquad = (0.320 \text{ mol/}\cancel{L})(2.00 \text{ }\cancel{L})$ $\qquad = 0.640 \text{ mol}$ $n_{POCl_3} = cV$ $\qquad = (0.5005 \text{ mol/}\cancel{L})(2.00 \text{ }\cancel{L})$ $\qquad = 1.00 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression $K_{eq} = 1.82$. This agrees with the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

67. In the reaction $PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$, the K_{eq} value at 305 K is 14.15. Initially, 0.81 mol $PCl_3(g)$ and 1.37 mol $Cl_2(g)$ are placed into a 1.00 L container and allowed to reach equilibrium. What are the equilibrium concentrations of all these substances?

What Is Required?
You need to determine the equilibrium concentration of each gas in the container.

What Is Given?
You know the balanced equation for the reaction.
You know the state of each component in the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles of $PCl_3(g)$: $n = 0.81$ mol
You know the initial amount in moles of $Cl_2(g)$: $n = 1.37$ mol
You know the value of the equilibrium constant, K_{eq}: 14.15

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $PCl_3(g)$ and $Cl_2(g)$ in mol/L.	$[PCl_3] = 0.81$ mol/L $[Cl_2] = 1.37$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $PCl_3(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$PCl_3(g)$ +	$Cl_2(g)$ $\rightleftharpoons$	$PCl_5(g)$
	$[PCl_3]$ (mol/L)	$[Cl_2]$ (mol/L)	$[PCl_5]$ (mol/L)
I	0.81	1.37	0
C	$-x$	$-x$	$+x$
E	$0.81 - x$	$1.37 - x$	x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[PCl_5]}{[PCl_3][Cl_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[PCl_5]}{[PCl_3][Cl_2]}$ $14.15 = \dfrac{x}{(0.81 - x)(1.37 - x)}$
Re-arrange the expression into a quadratic equation.	$14.15(1.1087 - 2.18x + x^2) = x$ $14.15x^2 - 31.847x + 15.7022 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-(-31.847) \pm \sqrt{(-31.847)^2 - (4)(14.15)(15.7022)}}{2(14.15)}$ $= \dfrac{31.847 \pm 11.2021}{28.3}$ $x = 0.7295$ or $x = 1.521$ Because it is impossible to have a negative concentration, x cannot be greater than 0.81. Therefore, $x = 0.7295$
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[PCl_3] = 0.81$ mol/L $- 0.7295$ mol/L $\qquad = 0.080$ mol/L $[Cl_2] = 1.37$ mol/L $- 0.7295$ mol/L $\qquad = 0.64$ mol/L $[PCl_5] = 0.7295$ mol/L $\qquad = 0.73$ mol/L

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 14.3$. This agrees closely with the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

68. The synthesis of nitrogen monoxide, NO(g), is an equilibrium reaction, as shown below:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$
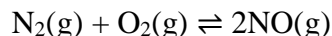
Initially, 1.33 mol of nitrogen gas and 2.80 mol of oxygen gas are placed into a 2.50 L container and allowed to reach equilibrium. What is the amount in moles of each gas at equilibrium at 200 K, if K_{eq} is 8.68 at this temperature?

What Is Required?
You need to determine the amount in moles of each reactant and product at equilibrium.

What Is Given?
You know the volume of the container: 2.50 L
You know the initial amount in moles of $N_2(g)$: $n = 1.33$ mol
You know the initial amount in moles of $O_2(g)$: $n = 2.80$ mol
You know the value of the equilibrium constant, K_{eq}: 8.68

| Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $N_2(g)$ and $O_2(g)$ in mol/L.	$[N_2] = \dfrac{1.33 \text{ mol}}{2.50 \text{ L}}$ $= 0.532$ mol/L $[O_2] = \dfrac{2.80 \text{ mol}}{2.50 \text{ L}}$ $= 1.12$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $N_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	$N_2(g)$ +	$O_2(g)$ $\rightleftharpoons$	$2NO(g)$
	$[N_2]$ (mol/L)	$[O_2]$ (mol/L)	$[NO]$ (mol/L)
I	0.532	1.12	0
C	$-x$	$-x$	$+2x$
E	$0.532 - x$	$1.12 - x$	$2x$

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[NO]^2}{[N_2][O_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[NO]^2}{[N_2][O_2]}$ $8.68 = \dfrac{(2x)^2}{(0.532 - x)(1.12 - x)}$
Re-arrange the expression into a quadratic equation.	$8.68(0.59584 - 1.652x + x^2 = 4x^2$ $4.68x^2 - 14.339x + 5.1719 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-(-14.339) \pm \sqrt{(-14.339)^2 - (4)(4.68)(5.1719)}}{2(4.68)}$ $= \dfrac{14.339 \pm 10.4302}{9.36}$ $x = 2.646 \text{ or } x = 0.4176$ Because it is impossible to have a negative concentration, x cannot be greater than 0.532. Therefore, $x = 0.4176$.
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[N_2] = 0.532 \text{ mol/L} - 0.4176 \text{ mol/L}$ $\qquad = 0.114 \text{ mol/L}$ $[O_2] = 1.12 \text{ mol/L} - 0.4176 \text{ mol/L}$ $\qquad = 0.70 \text{ mol/L}$ $[NO] = 2(0.4176)$ $\qquad = 0.8352 \text{ mol/L}$

Use the formula $n = cV$ to calculate the amount in moles of each component.	$n_{N_2} = cV$ $= (0.114 \text{ mol}/\cancel{L})(2.50 \ \cancel{L})$ $= 0.286 \text{ mol}$ $n_{O_2} = cV$ $= (0.70 \text{ mol}/\cancel{L})(2.50 \ \cancel{L})$ $= 1.76 \text{ mol}$ $n_{NO} = cV$ $= (0.8352 \text{ mol}/\cancel{L})(2.50 \ \cancel{L})$ $= 2.088 \text{ mol}$ $= 2.09 \text{ mol}$

Check Your Solution

The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression, $K_{eq} = 8.68$. This agrees with the given K_{eq}. The correct formula is used to convert to moles. The answers are reasonable and are expressed to the correct number of significant digits.

69. Carbon monoxide reaches equilibrium with water vapour as they react to form hydrogen gas and carbon dioxide gas. At 300 K, K_{eq} is equal to 0.52 for the reaction:

$$CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$$
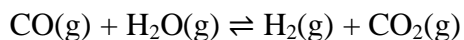
When 1.45 mol of carbon monoxide and 2.22 mol of water vapour are combined in a 1.00 L flask, equilibrium is established. What concentration of each reactant and product will be present at equilibrium?

What Is Required?
You need to determine the concentration of each reactant and product at equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the state of each component in the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles of $CO(g)$: $n = 1.45$ mol
You know the initial amount in moles of $H_2O(g)$: $n = 2.22$ mol
You know the value of the equilibrium constant, K_{eq}: 0.52

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $CO(g)$ and H_2O (g) in mol/L.	[CO] = 1.45 mol/L [H₂O] = 2.22 mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration in of $CO(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentrations of each component.	See the ICE table below.

	CO(g)	+	H₂O(g)	⇌	H₂(g)	+	CO₂(g)
	[CO] (mol/L)		[H₂O] (mol/L)		[H₂] (mol/L)		[CO₂] (mol/L)
I	1.45		2.22		0		0
C	−x		−x		+x		+x
E	1.45−x		2.22−x		x		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2][CO_2]}{[CO][H_2O]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[H_2][CO_2]}{[CO][H_2O]}$ $0.52 = \dfrac{(x)(x)}{(1.45-x)(2.22-x)}$
Re-arrange the expression into a quadratic equation.	$0.52(3.219 - 3.67x + x^2) = x^2$ $0.48x^2 + 1.9084x - 1.6734 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-(1.9084) \pm \sqrt{(1.9084)^2 - (4)(0.48)(-1.6734)}}{2(0.48)}$ $= \dfrac{-1.9084 \pm 2.6182}{0.96}$ $x = 0.7394$ or $x = -4.7152$ Because it is impossible to have a negative concentration, $x = 0.7394$.

Determine the equilibrium concentration of each component in the gas phase.	At equilibrium:
	
	 $[H_2]$ = 1.45 mol/L – 0.7394 mol/L
	       = 0.7106 mol/L
	       = 0.71 mol/L
	
	 $[CO_2]$ = 2.22 mol/L – 0.7394 mol/L
	       = 1.4806 mol/L
	       = 1.5 mol/L
	
	 $[CO] = [H_2O]$
	     = 0.7394 mol/L
	     = 0.74 mol/L

Check Your Solution
The expression for K_{eq} shows the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 0.52$. This agrees with the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

70. Sulfuryl chloride, $SO_2Cl_2(g)$ is used in a process to produce pesticides. It is also used as a convenient source of chlorine gas, because it can be cooled to form a pourable liquid that is easier to store and dispense. To generate this compound, the following reaction can be used, with a K_{eq} value of 0.45 at 650 K:

$$SO_2(g) + Cl_2(g) \rightleftharpoons SO_2Cl_2(g),$$
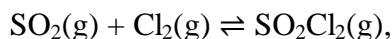
Initially, 3.15 mol $SO_2(g)$ and 2.14 mol $Cl_2(g)$ are placed into a 1.00 L container and allowed to reach equilibrium. What are the equilibrium concentrations of all these substances?

What Is Required?
You need to determine the concentration of each reactant and product at equilibrium.

What Is Given?
You know the balanced equation for the reaction.
You know the state of each component in the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles of SO_2 (g): $n = 3.15$ mol
You know the initial amount in moles of $Cl_2(g)$: $n = 2.14$ mol
You know the value of the equilibrium constant, K_{eq}: 0.45

| Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $SO_2(g)$ and $Cl_2(g)$ in mol/L.	$[SO_2] = 3.15$ mol/L $[Cl_2] = 2.14$ mol/L
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which theconcentration of $SO_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$SO_2(g)$	+	$Cl_2(g)$	$\rightleftharpoons$	$SO_2Cl_2(g)$ (mol/L)
	$[SO_2]$ (mol/L)		$[Cl_2]$ (mol/L)		$[SO_2Cl_2]$
I	3.15		2.14		0
C	$-x$		$-x$		$+x$
E	$3.15 - x$		$2.14 - x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[SO_2Cl_2]}{[SO_2][Cl_2]}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[SO_2Cl_2]}{[SO_2][Cl_2]}$ $0.45 = \dfrac{(x)}{(3.15-x)(2.14-x)}$
Re-arrange the expression into a quadratic equation.	$0.45(6.741 - 5.29x + x^2) = x$ $0.45x^2 - 3.3805x + 3.0334 = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-(-3.3805) \pm \sqrt{(-3.3805)^2 - (4)(0.45)(3.0334)}}{2(0.45)}$ $= \dfrac{3.3805 \pm 2.4429}{0.90}$ $x = 6.4773 \text{ or } x = 1.0418$ Because it is impossible to have a negative concentration, x cannot be greater than 2.14. Therefore, $x = 1.0418$.
Determine the equilibrium concentration of each component in the gas phase.	At equilibrium: $[SO_2] = 3.15 \text{ mol/L} - 1.0418 \text{ mol/L}$ $\qquad\quad = 2.1082 \text{ mol/L}$ $\qquad\quad = 2.1 \text{ mol/L}$ $[Cl_2] = 2.14 \text{ mol/L} - 1.0418 \text{ mol/L}$ $\qquad\quad = 1.0982 \text{ mol/L}$ $\qquad\quad = 1.1 \text{ mol/L}$ $[SO_2Cl_2] = 1.0418 \text{ mol/L}$ $\qquad\qquad = 1.0 \text{ mol/L}$

Check Your Solution
The expression for K_{eq} shows the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The concentrations of the gases are expressed in mol/L. Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 0.45$. This agrees with the given K_{eq}. The answers are reasonable and are expressed to the correct number of significant digits.

Using the Approximation Method in Equilibrium Calculations
(Student textbook page 457)

71. The following equation represents the equilibrium reaction for the decomposition of phosgene gas, $COCl_2(g)$:

$$COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g)$$

At 100°C, the value of K_{eq} for this reaction is 2.2×10^{-8}. The initial concentration of phosgene, $COCl_2(g)$, in a closed container at 100°C is 1.5 mol/L. What are the equilibrium concentrations of carbon monoxide gas, $CO(g)$, and chlorine gas, $Cl_2(g)$?

What Is Required?
You need to determine the equilibrium concentrations of carbon monoxide gas, $CO(g)$, and chlorine gas, $Cl_2(g)$

What Is Given?
You know the balanced chemical equation for the reaction.
You know the initial concentration of $COCl_2(g)$: 1.5 mol/L
You know the value of the equilibrium constant, K_{eq}: 2.2×10^{-8}

Plan Your Strategy	Act on Your Strategy
Set up an ICE table showing the initial molar concentrations. Let x represent the amount by which the concentration of $COCl_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$COCl_2(g)$ $\rightleftharpoons$	$CO(g)$ +	$Cl_2(g)$
	$[COCl_2]$ (mol/L)	$[CO]$ (mol/L)	$[Cl_2]$ (mol/L)
I	1.5	0	0
C	$-x$	$+x$	$+x$
E	$1.5 - x$	x	x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[CO][Cl_2]}{[COCl_2]}$

Test whether the approximation method can be used in solving for x. If the initial concentration of $COCl_2(g)$ is more than 1000 times greater than K_{eq}, the change in concentration of the reactant $COCl_2(g)$ is negligible. The initial concentration of the reactant can be used to represent the equilibrium concentration.	$\dfrac{[COCl_2]_{initial}}{K_{eq}} = \dfrac{1.5 \text{ mol/L}}{2.2 \times 10^{-8}}$ $\approx 7 \times 10^7$ The initial concentration of $COCl_2$ is more than 1000 times the K_{eq}. Therefore, $1.5 - x \approx 1.5$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$K_{eq} = \dfrac{[CO][Cl_2]}{[COCl_2]}$ $2.2 \times 10^{-8} = \dfrac{x^2}{1.5}$ $x^2 = 2.20$ $x = 1.816 \times 10^{-4} \text{ mol/L}$ $= 1.8 \times 10^{-4} \text{ mol/L}$
Determine the equilibrium concentrations of $CO(g)$ and $Cl_2(g)$.	At equilibrium: $[CO] = [Cl_2]$ $= 1.8 \times 10^{-4} \text{ mol/L}$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression gives $K_{eq} = 2.2 \times 10^{-8}$. This agrees with the given K_{eq}. The answer has the correct number of significant digits and is reasonable.

72. Hydrogen sulfide, $H_2S(g)$, is a poisonous gas with a characteristic offensive odour. At 1400°C, the gas decomposes, and K_{eq} is equal to 2.4×10^{-4}:

$$2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$$

If 4.0 mol of hydrogen sulfide is placed in a 3.0 L container, what is the equilibrium concentration of hydrogen gas, $H_2(g)$, at 1400°C?

What Is Required?
You need to determine the equilibrium concentration of hydrogen gas, $H_2(g)$

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 3.0 L
You know the initial amount in moles of H_2S (g): $n = 4.0$ mol
You know the value of the equilibrium constant, K_{eq}: 2.4×10^{-4}

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $H_2S(g)$ in mol/L.	$[H_2S] = \dfrac{4.0 \text{ mol}}{3.0 \text{ L}}$ $= 1.33$ mol/L
Set up an ICE table showing the initial molar concentration of $H_2S(g)$. Let $2x$ represent the amount by which the concentration of H_2S (g) had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	2H₂S(g)	⇌	2H₂(g)	+	S₂(g)
	$[H_2S]$ (mol/L)		$[H_2]$ (mol/L)		$[S_2]$ (mol/L)
I	1.33		0		0
C	$-2x$		$+2x$		$+x$
E	$1.33 - 2x$		$2x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2]^2[S_2]}{[H_2S]^2}$

Test whether the approximation method can be used in solving for x. If the initial concentration of reactant is more than 1000 times greater than K_{eq}, the change in concentration of the reactant is negligible. The initial concentration of the reactant can be used to represent the equilibrium concentration.	$$\frac{[H_2S]_{initial}}{K_{eq}} = \frac{1.33 \text{ mol/L}}{2.4 \times 10^{-4}}$$ $$\approx 5541$$ The initial concentration of H_2S is more than 1000 times the K_{eq}. Therefore, $1.33 - 2x \approx 1.33$.
Substitute the equilibrium concentratins and the value of K_{eq} into the expression and solve the equation for x.	$$K_{eq} = \frac{[H_2]^2[S_2]}{[H_2S]^2}$$ $$2.4 \times 10^{-4} = \frac{(2x)^2 x}{1.33^2}$$ $$x^3 = \frac{2.4 \times 10^{-4} \times 1.33^2}{4}$$ $$x = \sqrt[3]{\frac{2.4 \times 10^{-4} \times 1.33^2}{4}}$$ $$= 4.735 \times 10^{-2}$$
Determine the equilibrium concentration of $H_2(g)$.	At equilibrium: $$[H_2] = 2x$$ $$= 2(4.735 \times 10^{-2})$$ $$= 9.47 \times 10^{-2} \text{ mol/L}$$ $$= 9.5 \times 10^{-2} \text{ mol/L}$$

Check Your Solution
The answer has the correct number of significant digits and is reasonable.

73. At a particular temperature, K_{eq} for the decomposition of carbon dioxide gas, $CO_2(g)$, is 2.0×10^{-6}:

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$$

If 3.0 mol of carbon dioxide gas is put in a 5.0 L container, calculate the equilibrium concentration of each gas.

What Is Required?
You need to calculate the equilibrium concentration of each gas.

What Is Given?
You know the balanced equation for the reaction.
You know the volume of the container: 5.0 L
You know the initial amount in moles of $CO_2(g)$: $n = 3.0$ mol
You know the value of the equilibrium constant, K_{eq}: 2.0×10^{-6}

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $CO_2(g)$ in mol/L	$[CO_2] = \dfrac{3.0 \text{ mol}}{5.0 \text{ L}}$ $= 0.60$ mol/L
Set up an ICE table showing the initial molar concentrations. Let $2x$ represent the amount by which the concentration of $CO_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$2CO_2(g)$	$\rightleftharpoons$	$2CO(g)$	+	$O_2(g)$
	$[CO_2]$ (mol/L)		$[CO]$ (mol/L)		$[O_2]$ (mol/L)
I	0.60		0		0
C	$-2x$		$+2x$		$+x$
E	$0.60 - 2x$		$2x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[CO]^2[O_2]}{[CO_2]^2}$

Test whether the approximation method can be used in solving for x. If the initial concentration of $CO_2(g)$ is more than 1000 times greater than K_{eq}, the change in concentration of the reactant $CO_2(g)$ is negligible. The initial concentration of the reactant can be used to represent the equilibrium concentration.	$$\frac{[CO_2]_{initial}}{K_{eq}} = \frac{0.60 \text{ mol/L}}{2.0 \times 10^{-6}}$$ $$= 3 \times 10^5$$ The initial concentration of CO_2 is more than 1000 times the K_{eq}. Therefore $0.60 - 2x \approx 0.60$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$$K_{eq} = \frac{[CO]^2[O_2]}{[CO_2]^2}$$ $$2.0 \times 10^{-6} = \frac{(2x)^2 x}{0.60^2}$$ $$x^3 = \frac{2.0 \times 10^{-6} \times 0.60^2}{4}$$ $$x = \sqrt[3]{\frac{2.0 \times 10^{-6} \times 0.60^2}{4}}$$ $$= 5.646 \times 10^{-3}$$
Determine the equilibrium concentration of each component.	At equilibrium: $[CO_2] = 0.60 \text{ mol/L}$ $[CO] = 2x$ $\quad = 2(5.646 \times 10^{-3})$ $\quad = 1.129 \times 10^{-2} \text{ mol/L}$ $\quad = 1.1 \times 10^{-2} \text{ mol/L}$ $[O_2] = x$ $\quad = 5.646 \times 10^{-3} \text{ mol/L}$ $\quad = 5.6 \times 10^{-3} \text{ mol/L}$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression gives $K_{eq} = 1.1 \times 10^{-6}$. This approximates the given K_{eq}. The answers have the correct number of significant digits and seem reasonable.

74. At a certain temperature, the value of K_{eq} for the following reaction is 3.3×10^{-12}:

$$2NCl_3(g) \rightleftharpoons N_2(g) + 3Cl_2(g)$$
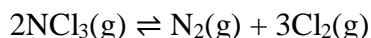
A certain amount of nitrogen trichloride, $NCl_3(g)$, is put in a 1.0 L reaction vessel at this temperature. At equilibrium, 4.6×10^{-6} mol of nitrogen gas, $N_2(g)$, is present. What amount of nitrogen trichloride was put in the reaction vessel?

What Is Required?
You need to determine the amount in mol of nitrogen trichloride that was put in the reaction vessel.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 1.0 L
You know the equilibrium amount in moles of $N_2(g)$: $n = 4.6 \times 10^{-6}$ mol
You know the value of the equilibrium constant, K_{eq}: 3.3×10^{-12}

Plan Your Strategy	Act on Your Strategy
Express the equilibrium concentration of $N_2(g)$ in mol/L	$[N_2] = 4.6 \times 10^{-6}$ mol/L
Estimate whether $[NCl_3]$ is $> 1000K_{eq}$.	The equilibrium constant is less than one so the reaction favours the reactant. Therefore, the equilibrium concentration and thus the initial concentration must be larger than the equilibrium concentration of the product, N_2. So: $[NCl_3]_I > 4.6 \times 10^{-6} \gg 3.3 \times 10^{-12}$. The approximation is valid. $[NCl_3]_E \approx [NCl_3]_I$
Since the equilibrium concentration of $N_2(g)$ is known, use the stoichiometry of the reaction to express the equilibrium concentration of $Cl_2(g)$ and $N_2(g)$. Set up an ICE table showing the equilibrium concentrations.	See the ICE table below.

	$2NCl_3(g)$ $\rightleftharpoons$	$N_2(g)$ +	$3Cl_2(g)$
	$[NCl_3]$ (mol/L)	$[N_2]$ (mol/L)	$[Cl_2]$ (mol/L)
I	$[NCl_3]_I$	0	0
C	$[NCl_3]_C$	4.6×10^{-6}	$3(4.6 \times 10^{-6})$
E	$[NCl_3]_E$	4.6×10^{-6}	1.38×10^{-5}

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[N_2][Cl_2]^3}{[NCl_3]_E^{\ 2}}$
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq}.	$K_{eq} = \dfrac{[N_2][Cl_2]^3}{[NCl_3]_E^{\ 2}}$ $3.3 \times 10^{-12} = \dfrac{(4.6 \times 10^{-6})(1.38 \times 10^{-5})^3}{[NCl_3]_E^{\ 2}}$
Solve for the equilibrium concentration of $NCl_3(g)$.	$[NCl_3]_E = \dfrac{(4.6 \times 10^{-6})(1.38 \times 10^{-5})^3}{(3.3 \times 10^{-12})}$ $= 6.0526 \times 10^{-5}$
Determine the initial concentration of $NCl_3(g)$.	The equilibrium concentration of $NCl_3(g) \approx$ initial concentration $NCl_3(g) = 6.1 \times 10^{-5}$ mol/L.

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression, $K_{eq} = 3.2 \times 10^{-12}$. This approximates the given K_{eq}. The answer has the correct number of significant digits and seems reasonable.

75. At a certain temperature, the value of K_{eq} for the following reaction is 4.2×10^{-8}:

$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
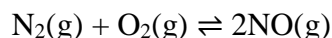
If 0.45 mol of nitrogen gas, $N_2(g)$, and 0.26 mol of oxygen gas, $O_2(g)$, are put in a 6.0 L reaction vessel, what is the equilibrium concentration of nitrogen monoxide, $NO(g)$, at this temperature?

What Is Required?
You need to determine the equilibrium concentration of nitrogen monoxide, $NO(g)$.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 6.0 L
You know the initial amount in moles of $N_2(g)$: $n = 0.45$ mol
You know the initial amount in moles of $O_2(g)$: $n = 0.26$ mol
You know the value of the equilibrium constant, K_{eq}: 4.2×10^{-8}

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $N_2(g)$ and $O_2(g)$ in mol/L	$[N_2] = \dfrac{0.45 \text{ mol}}{6.0 \text{ L}}$ $= 0.0750$ mol/L $[O_2] = \dfrac{0.26 \text{ mol}}{6.0 \text{ L}}$ $= 0.0433$ mol/L
Set up an ICE table showing the initial molar concentration of $N_2(g)$ and $O_2(g)$. Let x represent the amount by which the concentration of $N_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$N_2(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2NO(g)$
	$[N_2]$ (mol/L)		$[O_2]$ (mol/L)		$[NO]$ (mol/L)
I	0.0750		0.0433		0
C	$-x$		$-x$		$+2x$
E	$0.0750 - x$		$0.0433 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq}.	$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$
Test whether the approximation method can be used in solving for x. If the initial concentration of reactant is more than 1000 times greater than K_{eq}, the change in concentration of the reactant is negligible. The initial concentrations of the reactants can be used to represent the equilibrium concentration.	$$\frac{[O_2]_{initial}}{K_{eq}} = \frac{0.0433 \text{ mol/L}}{4.2 \times 10^{-8}}$$ $$\approx 1 \times 10^6$$ The initial concentration of O_2 is more than 1000 times the K_{eq}. Therefore, $0.0750 - x \approx 0.0750$ and $0.0433 - x \approx 0.0433$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression for K_{eq} and solve the equation for x.	$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$ $$4.2 \times 10^{-8} = \frac{(2x)^2}{(0.0750)(0.0433)}$$ $$x = 5.839 \times 10^{-6}$$
Determine the equilibrium concentration of $NO(g)$.	At equilibrium: $$[NO] = 2x$$ $$= 2(5.839 \times 10^{-6})$$ $$= 1.1678 \times 10^{-5} \text{ mol/L}$$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression, $K_{eq} = 4.4 \times 10^{-8}$. This approximates the given K_{eq}. The answer has the correct number of significant digits and seems reasonable.

76. For the reaction in question 75, K_{eq} changes to 1.0×10^{-5} at 1500 K. What equilibrium concentration of nitrogen oxide gas exists in the container at this new temperature?

What Is Required?
You have to determine the equilibrium concentration of nitrogen monoxide gas at a different temperature where the K_{eq} has a different value?

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 6.0 L
You know the initial amount in moles of $N_2(g)$: $n = 0.45$ mol
You know the initial amount in moles of $O_2(g)$: $n = 0.26$ mol
You know the value of the equilibrium constant, K_{eq}: 1.0×10^{-5}

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $N_2(g)$ and $O_2(g)$ in mol/L	$[N_2] = \dfrac{0.45 \text{ mol}}{6.0 \text{ L}}$ $= 0.0750$ mol/L $[O_2] = \dfrac{0.26 \text{ mol}}{6.0 \text{ L}}$ $= 0.0433$ mol/L
Set up an ICE table showing the initial molar concentration of $N_2(g)$ and $O_2(g)$. Let x represent the amount by which the concentration of $N_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$N_2(g)$	+	$O_2(g)$	$\rightleftharpoons$	2NO(g)
	$[N_2]$ (mol/L)		$[O_2]$ (mol/L)		[NO] (mol/L)
I	0.0750		0.0433		0
C	$-x$		$-x$		$+2x$
E	$0.0750 - x$		$0.0433 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[NO]^2}{[N_2][O_2]}$

Test whether the approximation method can be used in solving for x. If the initial concentration of reactant is more than 1000 times greater than K_{eq}, the change in concentration of the reactant is negligible. The initial concentrations of the reactants can be used to represent the equilibrium concentration.	$$\frac{[O_2]_{initial}}{K_{eq}} = \frac{0.0433 \text{ mol/L}}{1.0 \times 10^{-5}}$$ $$\approx 4 \times 10^3$$ The initial concentration of O_2 is more than 1000 times the K_{eq}. Therefore, $0.0750 - x \approx 0.0750$ and $0.0433 - x \approx 0.0433$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$$K_{eq} = \frac{[NO]^2}{[N_2][O_2]}$$ $$1.0 \times 10^{-5} = \frac{(2x)^2}{(0.0750)(0.0433)}$$ $$x^2 = 9.1188 \times 10^{-9}$$ $$x = 9.0104 \times 10^{-5}$$
Determine the equilibrium concentration of NO(g).	At equilibrium: $$[NO] = 2x$$ $$= 2(9.0104 \times 10^{-5})$$ $$= 1.80208 \times 10^{-5} \text{ mol/L}$$ $$= 1.8 \times 10^{-5} \text{ mol/L}$$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression gives $K_{eq} = 1.0 \times 10^{-5}$. This agrees with the given K_{eq}. The answer has the correct number of significant digits and seems reasonable.

77. At 300 K, K_{eq} is 1.43×10^{-26} for the following reaction:

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$$

Initially, 1.18 mol of sulfur trioxide, $SO_3(g)$, are placed into a 1.00 L container and equilibrium is established. What concentrations of $SO_2(g)$ and $O_2(g)$ exist in the container at equilibrium?

What Is Required?
You need to determine the equilibrium concentrations of $SO_2(g)$ and $O_2(g)$.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 1.00 L
You know the initial amount in moles of SO_3 (g): $n = 1.18$ mol
You know the value of the equilibrium constant, K_{eq}: 1.43×10^{-26}

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $SO_3(g)$ in mol/L	$[SO_3] = 1.18$ mol/L
Set up an ICE table showing the initial molar concentration of $SO_3(g)$. Let $2x$ represent the amount by which the concentration of $SO_3(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$2SO_3(g)$	$\rightleftharpoons$	$2SO_2(g)$	$+$	$O_2(g)$
	$[SO_3]$ (mol/L)		$[SO_2]$ (mol/L)		$[O_2]$ (mol/L)
I	1.18		0		0
C	$-2x$		$+2x$		$+x$
E	$1.18 - 2x$		$2x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[SO_2]^2[O_2]}{[SO_3]^2}$

Test whether the approximation method can be used in solving for x. If the initial concentration of reactant is more than 1000 times greater than K_{eq}, the change in concentration of the reactant is negligible. The initial concentrations of the reactants can be used to represent the equilibrium concentration.	$\dfrac{[SO_3]_{initial}}{K_{eq}} = \dfrac{1.18 \text{ mol/L}}{1.43 \times 10^{-26}}$ $= 8.2517 \times 10^{25}$ The initial concentration of SO_3 is more than 1000 times the K_{eq}. Therefore, $1.18 - x \approx 1.18$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$K_{eq} = \dfrac{[SO_2]^2[O_2]}{[SO_3]^2}$ $1.43 \times 10^{-26} = \dfrac{(2x)^2 x}{(1.18)^2}$ $x^3 = \dfrac{1.43 \times 10^{-26} \times 1.18^2}{4}$ $x = \sqrt[3]{\dfrac{1.43 \times 10^{-26} \times 1.18^2}{4}}$ $= 1.7074 \times 10^{-9}$
Determine the equilibrium concentration of $SO_2(g)$ and $O_2(g)$.	At equilibrium: $[SO_2] = 2x$ $= 2(1.7074 \times 10^{-9})$ $= 3.4148 \times 10^{-9} \text{ mol/L}$ $= 3.41 \times 10^{-9}$ $[O_2] = 1.7074 \times 10^{-9} \text{ mol/L}$ $= 1.71 \times 10^{-9}$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression gives $K_{eq} = 1.68 \times 10^{-26}$. This approximates the given K_{eq}. The answer has the correct number of significant digits and seems reasonable.

78. Carbon dioxide decomposes into carbon monoxide and oxygen gases with a K_{eq} value of 6.4×10^{-7} at 2300 K:

$$2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g)$$

Initially, 0.95 mol of carbon dioxide, $CO_2(g)$, is placed into a 1.00 L container and equilibrium is established. What concentrations of carbon monoxide gas and oxygen gas exist in the container at equilibrium at this temperature?

What Is Required?
You need to calculate the equilibrium concentration of carbon monoxide, $CO(g)$, and oxygen, $O_2(g)$.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the initial amount in moles of $CO_2(g)$: $n = 0.95$ mol
You know the volume of the container: 1.00 L
You know the value of the equilibrium constant: $K_{eq} = 6.4 \times 10^{-7}$

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $CO_2(g)$ in mol/L	$[CO_2] = 0.95$ mol/L
Set up an ICE table showing the initial molar concentrations. Let $2x$ represent the amount by which the concentration of $CO_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$2CO_2(g)$	$\rightleftharpoons$	$2CO(g)$	$+$	$O_2(g)$
	$[CO_2]$ (mol/L)		$[CO]$ (mol/L)		$[O_2]$ (mol/L)
I	0.95		0		0
C	$-2x$		$+2x$		$+x$
E	$0.95 - 2x$		$2x$		x

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[CO]^2[O_2]}{[CO_2]^2}$
Test whether the approximation method can be used in solving for x. If the initial concentration of $CO_2(g)$ is more than 1000 times greater than K_{eq}, the change in concentration of the reactant $CO_2(g)$ is negligible. The initial concentration of the reactant can be used to represent the equilibrium concentration.	$\dfrac{[CO_2]_{initial}}{K_{eq}} = \dfrac{0.95 \text{ mol/L}}{6.4 \times 10^{-7}}$ $= 1.5 \times 10^6$ The initial concentration of CO_2 is more than 1000 times the K_{eq}. Therefore, $0.95 - 2x \approx 0.95$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$K_{eq} = \dfrac{[CO]^2[O_2]}{[CO_2]^2}$ $6.4 \times 10^{-7} = \dfrac{(2x)^2 x}{0.95^2}$ $x^3 = \dfrac{6.4 \times 10^{-7} \times 0.95^2}{4}$ $x = \sqrt[3]{\dfrac{6.4 \times 10^{-7} \times 0.95^2}{4}}$ $= 5.246 \times 10^{-3}$
Determine the equilibrium concentration of $CO(g)$ and $O_2(g)$.	At equilibrium: $[CO] = 2x$ $\quad = 2(5.246 \times 10^{-3})$ $\quad = 1.04926 \times 10^{-3} \text{ mol/L}$ $\quad = 1.05 \times 10^{-2} \text{ mol/L}$ $[O_2] = x$ $\quad = 5.246 \times 10^{-3} \text{ mol/L}$ $\quad = 5.25 \times 10^{-3} \text{ mol/L}$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression gives $K_{eq} = 6.5 \times 10^{-7}$. This approximates the given K_{eq}. The answers have the correct number of significant digits and seem reasonable.

79. $K_{eq} = 1.0 \times 10^{-7}$ for the reaction below:

$$2H_2(g) + O_2(g) \rightleftharpoons 2H_2O(g)$$
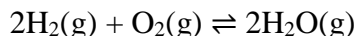
If 0.18 mol of hydrogen gas, $H_2(g)$, and 0.23 mol of oxygen gas, $O_2(g)$, are put in a 1.0 L reaction vessel, what is the equilibrium concentration of the water vapour at this temperature?

What Is Required?
You need to determine the equilibrium concentration of water vapour, $H_2O(g)$.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 1.0 L
You know the initial amount in moles of $H_2(g)$: $n = 0.18$ mol
You know the initial amount in moles of $O_2(g)$: $n = 0.23$ mol
You know the value of the equilibrium constant, K_{eq}: 1.0×10^{-7}

| Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $H_2(g)$ and $O_2(g)$ in mol/L	$[H_2] = 0.18$ mol/L $[O_2] = 0.23$ mol/L
Set up an ICE table showing the initial molar concentration of $H_2(g)$ and $O_2(g)$. Let x represent the amount by which the concentration of $H_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$2H_2(g)$	$+$	$O_2(g)$	$\rightleftharpoons$	$2H_2O(g)$
	$[H_2]$ (mol/L)		$[O_2]$ (mol/L)		$[H_2O]$ (mol/L)
I	0.18		0.23		0
C	$-2x$		$-x$		$+2x$
E	$0.18 - 2x$		$0.23 - x$		$2x$

Write the expression for the equilibrium constant, K_{eq}.	$K_{eq} = \dfrac{[H_2O]^2}{[H_2]^2[O_2]}$

Test whether the approximation method can be used in solving for x. If the initial concentration of reactant is more than 1000 times greater than K_{eq}, the change in concentration of the reactant is negligible. The initial concentrations of the reactants can be used to represent the equilibrium concentration.	$\dfrac{[H_2]_{initial}}{K_{eq}} = \dfrac{0.18 \text{ mol/L}}{1.0 \times 10^{-7}}$ $\approx 2 \times 10^6$ The initial concentration of H_2 is more than 1000 times the value of K_{eq}. Therefore, $0.18 - x \approx 0.18$ and $0.23 - x \approx 0.23$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$K_{eq} = \dfrac{[H_2O]^2}{[H_2]^2[O_2]}$ $1.0 \times 10^{-7} = \dfrac{(2x)^2}{(0.18)^2(0.23)}$ $x^2 = \dfrac{1.0 \times 10^{-7} \times 0.18^2 \times 0.23}{4}$ $x = \sqrt{\dfrac{1.0 \times 10^{-7} \times 0.18^2 \times 0.23}{4}}$ $= 1.3649 \times 10^{-5}$
Determine the equilibrium concentration of $H_2O(g)$.	At equilibrium: $[H_2O] = 2x$ $= 2(1.3649 \times 10^{-5})$ $= 2.729 \times 10^{-5}$ mol/L $= 2.7 \times 10^{-5}$ mol/L

Check Your Solution
Substituting the calculated equilibrium concentrations into the equilibrium expression gives $K_{eq} = 1.0 \times 10^{-7}$. This agrees with the given K_{eq}. The answer has the correct number of significant digits and seems reasonable.

80. When nitrogen gas and chloride gas react to form nitrogen trichloride gas, an equilibrium is established in which $K_{eq} = 4.15 \times 10^{-5}$ at a given temperature. The reaction is as follows:

$$N_2(g) + 3Cl_2(g) \rightleftharpoons 2NCl_3(g)$$
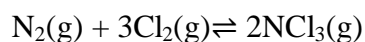
If 2.74 mol of nitrogen gas, $N_2(g)$, and 0.84 mol of chlorine gas, $Cl_2(g)$, are put in a 2.0 L reaction vessel, what is the equilibrium concentration of the nitrogen trichloride gas at this temperature?

What Is Required?
You need to determine the equilibrium concentration of nitrogen trichloride, $NCl_3(g)$.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the volume of the container: 2.0 L
You know the initial amount in moles of N_2 (g): $n = 2.74$ mol
You know the initial amount in moles of $Cl_2(g)$: $n = 0.84$ mol
You know the value of the equilibrium constant, K_{eq}: 4.15×10^{-5}

Plan Your Strategy	Act on Your Strategy
Express the initial concentration of $N_2(g)$ and $O_2(g)$ in mol/L	$[N_2] = \dfrac{2.74 \text{ mol}}{2.0 \text{ L}}$ $= 1.37$ mol/L $[Cl_2] = \dfrac{0.84 \text{ mol}}{2.0 \text{ L}}$ $= 0.42$ mol/L
Set up an ICE table showing the initial molar concentration of $N_2(g)$ and $Cl_2(g)$. Let x represent the amount by which the concentration of $N_2(g)$ had decreased when the system had reached equilibrium. Use the stoichiometry of the reaction to determine the equilibrium concentration of each component.	See the ICE table below.

	$N_2(g)$	+	$3Cl_2(g)$	$\rightleftharpoons$	$2NCl_3(g)$
	$[N_2]$ (mol/L)		$[Cl_2]$ (mol/L)		$[NCl_3]$ (mol/L)
I	1.37		0.42		0
C	$-x$		$-3x$		$+2x$
E	$1.37 - x$		$0.42 - 3x$		$2x$

Write the expression for the equilibrium constant, K_{eq}.	$$K_{eq} = \frac{[NCl_3]^2}{[N_2][Cl_2]^3}$$
Test whether the approximation method can be used in solving for x. If the initial concentration of reactant is more than 1000 times greater than K_{eq}, the change in concentration of the reactant is negligible. The initial concentrations of the reactants can be used to represent the equilibrium concentration.	$$\frac{[N_2]_{initial}}{K_{eq}} = \frac{1.37 \text{ mol/L}}{4.15 \times 10^{-5}}$$ $$= 3.301 \times 10^4$$ The initial concentration of N_2 is more than 1000 times the value of K_{eq}. Therefore, $1.37 - x \approx 1.37$ and $0.42 - 3x \approx 0.42$.
Substitute the equilibrium concentrations and the value of K_{eq} into the expression and solve the equation for x.	$$K_{eq} = \frac{[NCl_3]^2}{[N_2][Cl_2]^3}$$ $$4.15 \times 10^{-5} = \frac{(2x)^2}{(1.37)(0.42)^3}$$ $$x^2 = \frac{(4.15 \times 10^{-5})(1.37)(0.42)^3}{4}$$ $$x = \sqrt{\frac{(4.15 \times 10^{-5})(1.37)(0.42)^3}{4}}$$ $$x = 1.026 \times 10^{-3}$$
Determine the equilibrium concentration of $NCl_3(g)$.	At equilibrium: $$[NCl_3] = 2x$$ $$= 2(1.026 \times 10^{-3})$$ $$= 2.0523 \times 10^{-3} \text{ mol/L}$$ $$= 2.1 \times 10^{-3} \text{ mol/L}$$

Check Your Solution
Substituting the calculated equilibrium concentration into the equilibrium expression gives $K_{eq} = 4.3 \times 10^{-5}$. This approximates the given K_{eq}. The answer has the correct number of significant digits and seems reasonable.

Calculating and Using a Reaction Quotient
(Student textbook page 461)

81. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 300 K, $K_{eq} = 35.5$ for the reaction below:

$$3C(s) + 3H_2(g) \rightleftharpoons CH_4(g) + C_2H_2(g)$$
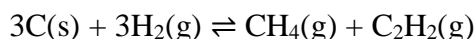
The following concentrations are found to be present at a particular point in time:
$[H_2(g)] = 0.34$ mol/L, $[CH_4(g)] = 2.13$ mol/L, and $[C_2H_2(g)] = 1.77$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know the value of the equilibrium constant, K_{eq}: 35.5
You know the concentration of the components:
   $[H_2(g)] = 0.34$ mol/L
   $[CH_4(g)] = 2.13$ mol/L
   $[C_2H_2(g)] = 1.77$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \dfrac{[CH_4][C_2H_2]}{[H_2]^3}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \dfrac{[CH_4][C_2H_2]}{[H_2]^3}$ $= \dfrac{(2.13)(1.77)}{(0.34)^3}$ $= 95.9$ $= 96$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} > K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too large. The reaction will shift to the left to lower the concentration of products and increase the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

82. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. A 3.00 L reaction vessel contains 9.00 mol of hydrogen iodide gas, 6.00 mol of hydrogen gas, and 4.50 mol of iodine gas. At a given temperature, the K_{eq} value for the reaction below is 50.0.
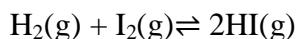
$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

What Is Required?

You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?

You know the balanced chemical equation for the reaction.
You know the volume of reaction vessel: 3.00 L
You know the value of the equilibrium constant, K_{eq}: 50.0.
You know the amounts in moles of the components:

$H_2(g) = 6.00$ mol
$HI(g) = 9.00$ mol
$I_2(g) = 4.50$ mol

Plan Your Strategy	Act on Your Strategy
Express the concentration of each component in mol/L.	$[H_2] = \dfrac{6.0 \text{ mol}}{3.0 \text{ L}}$ $= 2.0$ mol/L $[HI] = \dfrac{9.0 \text{ mol}}{3.0 \text{ L}}$ $= 3.0$ mol/L $[I_2] = \dfrac{4.5 \text{ mol}}{3.0 \text{ L}}$ $= 1.5$ mol/L
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \dfrac{[HI]^2}{[H_2][I_2]}$ $= \dfrac{(3.0)^2}{(2.0)(1.5)}$ $= 3.0$

Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} < K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

83. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 525 K, $K_{eq} = 0.041$ for the following reaction:

$$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$$

At one point in time, measurements were taken and the following concentrations were measured: $[PCl_5(g)] = 0.77$ mol/L, $[PCl_3(g)] = 1.21$ mol/L, and $[Cl_2(g)] = 0.49$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know that $K_{eq} = 0.041$.
You know the concentration of the components:
  $PCl_5(g)] = 0.77$ mol/L
  $PCl_3(g)] = 1.21$ mol/L
  $[Cl_2(g)] = 0.49$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \dfrac{[PCl_3][Cl_2]}{[PCl_5]}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \dfrac{[PCl_3][Cl_2]}{[PCl_5]}$ $= \dfrac{(1.21)(0.49)}{(0.77)}$ $= 0.77$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} > K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too large. The reaction will shift to the left to lower the concentration of products and increase the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

84. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. Consider the following reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$
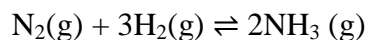
$K_{eq} = 6.0 \times 10^{-2}$ at 875 K. At a point just after nitrogen gas and hydrogen gas are injected into a container, the concentrations of each reactant and product are $[N_2(g)] = 1.50 \times 10^{-5}$ mol/L , $[H_2(g)] = 0.354$ mol/L, and $[NH_3(g)] = 2.00 \times 10^{-4}$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know that $K_{eq} = 6.0 \times 10^{-2}$.
You know the concentration of components:
  $[N_2(g)] = 1.50 \times 10^{-5}$ mol/L
  $[H_2(g)] = 0.354$ mol/L
  $[NH_3(g)] = 2.00 \times 10^{-4}$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$ $= \dfrac{(2.00 \times 10^{-4})^2}{(1.50 \times 10^{-5})(0.354)^3}$ $= 0.060$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} = K_{eq}$ the system is at equilibrium. There is no shift in the position of equilibrium.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

85. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. In the reaction of phosphorus trichloride, $PCl_3(g)$, and ammonia, $NH_3(g)$, hydrogen chloride gas and $P(NH_2)_3(g)$ forms by this reaction:

$$PCl_3(g) + 3NH_3(g) \rightleftharpoons P(NH_2)_3(g) + 3HCl(g)$$
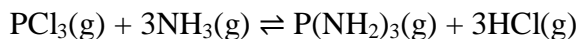
At 600 K, this reaction has an equilibrium constant, K_{eq}, that is 142.3. A system of these gases was analyzed and the following concentrations were determined: $[PCl_3(g)] = 2.15$ mol/L , $[NH_3(g)] = 1.21$ mol/L, $[P(NH_2)_3(g)] = 5.74$ mol/L and $[HCl(g)] = 2.04$ mol/L.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation.
You know that $K_{eq} = 142.3$.
You know the concentration of the components:
  $[PCl_3(g)] = 2.15$ mol/L
  $[NH_3(g)] = 1.21$ mol/L
  $[P(NH_2)_3(g)] = 5.74$ mol/L
  $[HCl(g)] = 2.04$ mol/L

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \dfrac{[P(NH_2)_3][HCl]^3}{[PCl_3][NH_3]^3}$
Substitute the given concentrations into the reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \dfrac{[P(NH_2)_3][HCl]^3}{[PCl_3][NH_3]^3}$ $= \dfrac{(5.74)(2.04)^3}{(2.15)(1.21)^3}$ $= 12.8$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} < K_{eq}$, the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

86. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 865 K, $K_p = 2.11$ for the reaction below:
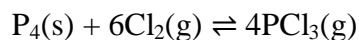
$$P_4(s) + 6Cl_2(g) \rightleftharpoons 4PCl_3(g)$$

At one point in time, measurements were taken and the following partial pressures were measured: $P_{Cl_2} = 1.21$ atm, $P_{PCl_3} = 3.04$ atm

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation.
You know the value of K_p: 2.11
You know the partial pressures of the components:

$P_{Cl_2} = 1.21$ atm

$P_{PCl_3} = 3.04$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_p.	$Q_p = \dfrac{P_{PCl_3}^4}{P_{Cl_2}^6}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_p.	$Q_p = \dfrac{P_{PCl_3}^4}{P_{Cl_2}^6}$ $= \dfrac{(3.04)^4}{(1.21)^6}$ $= 27.2$
Compare the calculated value of Q_p and the given K_p to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_p > K_p$ the system is not at equilibrium. The ratio of products to reactants is too large. The reaction will shift to the left to lower the concentration of products and increase the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the product in the numerator and the reactant in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

87. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. At 660 K, $K_p = 25.2$ for the reaction below:
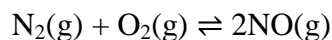
$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

At one point in time, measurements were taken and the following partial pressures were measured:
$P_{N_2} = 2.37$ atm, $P_{O_2} = 2.27$ atm, and $P_{NO} = 3.74$ atm.

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation for the reaction.
You know that $K_p = 25.2$.
You know the partial pressures of the components:
$P_{NO} = 3.74$ atm
$P_{N_2} = 2.37$ atm
$P_{O_2} = 2.27$ atm

Plan Your Strategy	Act on Your Strategy
Write the expression for the reaction quotient, Q_p.	$Q_p = \dfrac{P_{NO}^2}{P_{N_2} P_{O_2}}$
Substitute the given concentrations into the reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, the concentrations are written without units. Calculate the value of Q_p.	$Q_p = \dfrac{P_{NO}^2}{P_{N_2} P_{O_2}}$ $= \dfrac{(3.74)^2}{(2.37)(2.27)}$ $= 2.60$
Compare the calculated value of Q_p and the given K_p to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_p < K_p$ the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the product in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

88. Determine if the reaction is at equilibrium. If it is not at equilibrium, then determine the direction in which the reaction is favoured. If 1.00 mol of each reactant and product in the reaction below are placed into a 1.00 L container and if the system is kept at 700 K, in what direction will the reaction shift to reach equilibrium, if K_{eq} is 5.10?
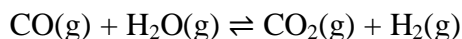
$$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$$

What Is Required?
You need to determine if a system is at equilibrium and, if not, to predict in which direction the reaction will shift to reach equilibrium.

What Is Given?
You know the balanced chemical equation.
You know the volume of the reaction vessel: 1.00 L
You know the value of the equilibrium constant: $K_{eq} = 5.10$
You know the amounts in moles of the components:
  $[CO(g)] = 1.00$ mol
  $[H_2O(g)] = 1.00$ mol
  $[CO(g)] = 1.00$ mol
  $[H_2(g)] = 1.00$ mol

Plan Your Strategy	Act on Your Strategy
Express the concentration of each component in mol/L.	$[CO(g)] = 1.00$ mol/L $[H_2O(g)] = 1.00$ mol/L $[CO(g)] = 1.00$ mol/L $[H_2(g)] = 1.00$ mol/L
Write the expression for the reaction quotient, Q_{eq}.	$Q_{eq} = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$
Substitute the given concentrations into this reaction quotient. Each concentration is divided by the reference state of 1 mol/L. Therefore, they are written without units. Calculate the value of Q_{eq}.	$Q_{eq} = \dfrac{[CO_2][H_2]}{[CO][H_2O]}$ $= \dfrac{(1.00)(1.00)}{(1.00)(1.00)}$ $= 1.0$
Compare the calculated value of Q_{eq} and the given K_{eq} to determine if the system is at equilibrium. If not, predict which way the equilibrium will shift to reach equilibrium.	Since $Q_{eq} < K_{eq}$ the system is not at equilibrium. The ratio of products to reactants is too small. The reaction will shift to the right to increase the concentration of products and decrease the concentration of reactants.

Check Your Solution
The expression for the reaction quotient has the products in the numerator and the reactants in the denominator, each raised to the power of its coefficient in the balanced equation. The substitution of concentration into this expression has been made correctly.

Solutions to Practice Problems in
Chapter 8 Acid-Base Equilibrium Systems

Identifying Conjugate Acid-Base Pairs
(Student textbook page 496)

1. Name and write the formula for the conjugate base of the molecule or ion below.
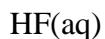
HF(aq)

What Is Required?
You need to name and write the formula for the conjugate base.

What Is Given?
You know the formula for the acid is HF(aq).

Plan Your Strategy	Act on Your Strategy
The conjugate base will have one less proton than its conjugate acid.	Remove one proton, H^+, from the molecule or ion. Adjust the charge on the molecule or ion to match the removal of the proton. F^-(aq)
Write the name and formula of the conjugate base.	The conjugate base is the fluoride ion, F^-(aq).

Check Your Solution
The acid and its conjugate base differ by one proton, H^+.

2. Name and write the formula for the conjugate base of the molecule or ion below.

HCO_3^-(aq)

What Is Required?
You need to name and write the formula for the conjugate base.

What Is Given?
You know the formula for the acid is HCO_3^-(aq).

Plan Your Strategy	Act on Your Strategy
The conjugate base will have one less proton than its conjugate acid.	Remove one proton, H^+, from the molecule or ion. Adjust the charge on the molecule or ion to match the removal of the proton. CO_3^{2-}(aq)
Write the name and formula of the conjugate base.	The conjugate base is the carbonate ion, CO_3^{2-}(aq).

Check Your Solution
The acid and its conjugate base differ by one proton, H^+.

3. Name and write the formula for the conjugate base of the molecule or ion below.

H_2SO_4(aq)

What Is Required?
You need to name and write the formula for the conjugate base.

What Is Given?
You know the formula for the acid is H_2SO_4(aq).

Plan Your Strategy	Act on Your Strategy
The conjugate base will have one less proton than its conjugate acid.	Remove one proton, H^+, from the molecule or ion. Adjust the charge on the molecule or ion to match the removal of the proton. HSO_4^-(aq)
Write the name and formula of the conjugate base.	The conjugate base is the hydrogen sulfate ion, HSO_4^-(aq).

Check Your Solution
The acid and its conjugate base differ by one proton, H^+.

4. Name and write the formula for the conjugate base of the molecule or ion below.
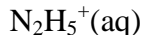
$N_2H_5^+(aq)$

What Is Required?
You need to name and write the formula for the conjugate base.

What Is Given?
You know the formula for the acid is $N_2H_5^+(aq)$.

Plan Your Strategy	Act on Your Strategy
The conjugate base will have one less proton than its conjugate acid.	Remove one proton, H^+, from the molecule or ion. Adjust the charge on the molecule or ion to match the removal of the proton. $N_2H_4(aq)$
Write the name and formula of the conjugate base.	The conjugate base is hydrazine, $N_2H_4(aq)$.

Check Your Solution
The acid and its conjugate base differ by one proton, H^+.

5. Name and write the formula for the conjugate acid of the molecule or ion.
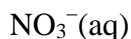
$NO_3^-(aq)$

What Is Required?
You need to name and write the formula for the conjugate acid.

What Is Given?
You know the formula for the base is $NO_3^-(aq)$.

Plan Your Strategy	Act on Your Strategy
The conjugate acid will have one more proton than its conjugate base.	Add one proton, H^+, to the molecule or ion. Adjust the charge on the molecule or ion to match the addition of the proton. $HNO_3(aq)$
Write the name and formula of the conjugate acid.	The conjugate acid is nitric acid, $HNO_3(aq)$.

Check Your Solution
The base and its conjugate acid differ by one proton, H^+.

6. Name and write the formula for the conjugate acid of the molecule or ion.
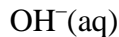
$OH^-(aq)$

What Is Required?
You need to name and write the formula for the conjugate acid.
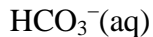
What Is Given?
You know the formula for the base is $OH^-(aq)$.

Plan Your Strategy	Act on Your Strategy
The conjugate acid will have one more proton than its conjugate base.	Add one proton, H^+, to the molecule or ion. Adjust the charge on the molecule or ion to match the addition of the proton. $H_2O(\ell)$
Write the name and formula of the conjugate acid.	The conjugate acid is water, $H_2O(\ell)$.

Check Your Solution
The base and its conjugate acid differ by one proton, H^+.

7. Name and write the formula for the conjugate acid of the molecule or ion.
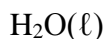
$H_2O(\ell)$

What Is Required?
You need to name and write the formula for the conjugate acid.

What Is Given?
You know the formula for the base is $H_2O(\ell)$.

Plan Your Strategy	Act on Your Strategy
The conjugate acid will have one more proton than its conjugate base.	Add one proton, H^+, to the molecule or ion. Adjust the charge on the molecule or ion to match the addition of the proton. $H_3O^+(aq)$
Write the name and formula of the conjugate acid.	The conjugate acid is the hydronium ion, $H_3O^+(aq)$.

Check Your Solution
The base and its conjugate acid differ by one proton, H^+.

8. Name and write the formula for the conjugate acid of the molecule or ion.
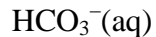
$HCO_3^-(aq)$

What Is Required?
You need to name and write the formula for the conjugate acid.

What Is Given?
You know the formula for the base is $HCO_3^-(aq)$.

Plan Your Strategy	Act on Your Strategy
The conjugate acid will have one more proton than its conjugate base.	Add one proton, H^+, to the molecule or ion. Adjust the charge on the molecule or ion to match the addition of the proton. $H_2CO_3(aq)$
Write the name and formula of the conjugate acid.	The conjugate acid is carbonic acid, $H_2CO_3(aq)$.

Check Your Solution
The base and its conjugate acid differ by one proton, H^+.

9. When perchloric acid dissolves in water, the following reaction occurs:

$$HClO_4(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$$
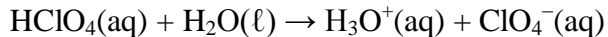
Identify the conjugate acid-base pairs in the reaction above.

What Is Required?
You need to identify the conjugate acid-base pairs for the reaction.

What Is Given?
You know the balanced acid-base equation.

Plan Your Strategy	Act on Your Strategy
On the left side of the equation, identify the acid as the molecule or ion that donates a proton, H^+. Identify the base as the molecule or ion that accepts the proton, H^+. Identify the conjugate base on the right side of the equation as the particle with one less proton than the acid on the left side of the equation. Identify the particle on the right side that has one proton more than the base on the left side as the conjugate acid.	Identify the conjugate acid-base pairs as $Acid_1/Base_1$ and $Acid_2/Base_2$. $HClO_4(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$ $\quad Acid_1 \qquad\quad Base_2 \qquad\quad Acid_2 \qquad\quad Base_1$
Name the conjugate acid-base pairs.	The conjugate acid-base pairs are $HClO_4(aq)/ClO_4^-(aq)$ and $H_2O(\ell)/H_3O^+(aq)$.

Check Your Solution
The formulas for the conjugate acid-base pairs differ by one proton, H^+.

10. Identify the conjugate acid-base pairs in the following reactions:

    a. $HS^-(aq) + H_2O(\ell) \rightleftharpoons H_2S(aq) + OH^-(aq)$

    b. $O_2^-(aq) + H_2O(\ell) \rightleftharpoons 2OH^-(aq)$

    c. $H_2S(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + HS^-(aq)$

    d. $H_2SO_4(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$
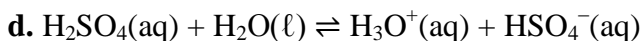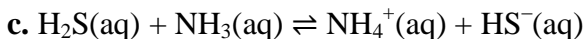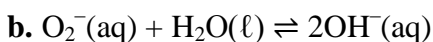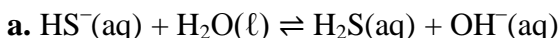
What Is Required?
You need to identify the conjugate acid-base pairs for each reaction.

What Is Given?
You know the balanced acid-base equation for each reaction.

Plan Your Strategy	Act on Your Strategy

On the left side of the equation, identify the acid as the molecule or ion that donates a proton, H^+. Identify the base as the molecule or ion that accepts the proton, H^+.
 	
Identify the conjugate base on the right side of the equation as the particle with one less proton than the acid on the left side of the equation. Identify the particle on the right side that has one proton more than the base on the left side as the conjugate acid.
	Identify the conjugate acid-base pairs as $Acid_1/Base_1$ and $Acid_2/Base_2$
	
	a. $HS^-(aq) + H_2O(\ell) \rightleftharpoons H_2S(aq) + OH^-(aq)$
	   $Base_2$ $Acid_1$ $Acid_2$ $Base_1$
	
	b. $O_2^-(aq) + H_2O(\ell) \rightleftharpoons OH^-(aq) + OH^-(aq)$
	   $Base_2$ $Acid_1$ $Base_1$ $Acid_2$
	
	c. $H_2S(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + HS^-(aq)$
	   $Acid_1$ $Base_2$ $Acid_2$ $Base_1$
	
	d. $H_2SO_4(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$
	   $Acid_1$ $Base_2$ $Acid_2$ $Base_1$

Name the conjugate acid-base pairs.	**a.** $H_2O(\ell)/OH^-(aq)$; $H_2S(aq)/HS^-(aq)$ **b.** $H_2O(\ell)/OH^-(aq)$; $OH^-(aq)/O^{2-}(aq)$ **c.** $H_2S(aq)/HS^-(aq)$; $NH_4^+(aq)/NH_3(aq)$ **d.** $H_2SO_4(aq)/HSO_4^-(aq)$; $H_3O^+(aq)/H_2O(\ell)$

Check Your Solution
The formulas for the conjugate acid-base pairs differ by one proton, H^+.

Calculating the [H₃O⁺] or [OH⁻] in Aqueous Solution
(Student textbook page 500)

11. The concentration of hydroxide ions, $OH^-(aq)$, in a solution at 25°C is 0.150 mol/L. Determine the concentration of hydronium ions, $H_3O^+(aq)$, in the solution.

What Is Required?
You need to determine the hydronium ion concentration in an aqueous solution.

What Is Given?
You know that $[OH^-]$ is 0.150 mol/L.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Rearrange the expression for the water constant to solve for $[H_3O^+]$.	$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ $\left[H_3O^+\right] = \dfrac{K_w}{[OH^-]}$ $= \dfrac{1.0 \times 10^{-14} \ \dfrac{\text{mol}^2}{\text{L}^2}}{0.150 \ \dfrac{\text{mol}}{\text{L}}}$ $= 6.67 \times 10^{-14} \, \text{mol/L}$

Check Your Solution
Multiplying the calculated $[H_3O^+]$ by the given $[OH^-]$ gives the water constant, 1.0×10^{-14}.

12. A solution of lithium hydroxide, LiOH(aq), is made by placing 2.00 mol of the base into 1.50 L of solution. What is the concentration of hydronium ions, H_3O^+(aq), in this solution at 25°C?

What Is Required?
You need to find the concentration of hydronium ions, $[H_3O^+]$, in the solution.

What Is Given?
You know there is 2.00 mol of LiOH in 1.50 L of solution.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Calculate the molar concentration of LiOH. Since LiOH is a strong base, $[OH^-]$ = [LiOH].	$$[LiOH] = [OH^-]$$ $$= \frac{n}{V}$$ $$= \frac{2.00 \text{ mol}}{1.50 \text{ L}}$$ $$= 1.33 \text{ mol/L}$$
Rearrange the expression for the water constant to solve for $[H_3O^+]$.	$$[H_3O^+] = \frac{K_w}{[OH^-]}$$ $$= \frac{1.0 \times 10^{-14} \frac{\text{mol}^2}{\text{L}^2}}{1.33 \frac{\text{mol}}{\text{L}}}$$ $$= 7.52 \times 10^{-15} \text{ mol/L}$$

Check Your Solution
Multiplying the calculated $[H_3O^+]$ and $[OH^-]$ values gives the water constant 1.0×10^{-14}.

13. Two solutions at the same temperature are to be compared for their hydronium ion concentrations. Solution A has a hydroxide ion concentration of 0.335 mol/L, and solution B has a hydroxide ion concentration of 0.285 mol/L. In which solution will the hydronium ion concentration be larger? Explain your answer.

What Is Required?
You need to compare the hydronium ion concentrations in two solutions.

What Is Given?
You know the hydroxide concentration $[OH^-]$ for the two solutions.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
From the expression for the water constant, $[H_3O^+]$ and $[OH^-]$ are inversely proportional.	Compare the $[OH^-]$ for the two solutions: Solution A: $[OH^-] = 0.335$ mol/L Solution B: $[OH^-] = 0.285$ mol/L
The solution with the smaller $[OH^-]$ will have the higher $[H_3O^+]$.	Since solution B has the lower $[OH^-]$, it must also have the higher $[H_3O^+]$.

Check Your Solution
Using the expression for the water constant, $[H_3O^+]$ for solution A is 2.99×10^{-14} mol/L. For solution B, $[H_3O^+]$ is 3.51×10^{-14} mol/L. Solution B has a higher $[H_3O^+]$.

14. Explain why the two solutions in question 13 need only to be at the same temperature for the comparison to be made, and do not necessarily have to be at 25°C.

Answer
$[H_3O^+]$ and $[OH^-]$ will still be inversely proportional even though the K_w will be different at other temperatures.

15. In a solution at 25°C, 4.75 g of sodium hydroxide, NaOH(s), is added to enough water to create a volume of 2.40 L. What is the hydronium ion, H_3O^+(aq), concentration in this solution?

What Is Required?
You need to find the concentration of hydronium ions, $[H_3O^+]$, in the solution.

What Is Given?
You know there is 4.75 g of NaOH in 2.40 L of solution.
The expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Calculate the amount in moles, n, of NaOH.	$n_{NaOH} = \dfrac{m}{M}$ $= \dfrac{4.75\ g}{40.00\ g/mol}$ $= 0.11875\ mol$
Use the formula $c = \dfrac{n}{V}$ to calculate the molar concentration, c, of NaOH. Since NaOH is a strong base, $[OH^-] = [NaOH]$.	$c_{NaOH} = \dfrac{n}{V}$ $= \dfrac{0.118\ 75\ mol}{2.40\ L}$ $= 0.04948\ mol/L$ $[OH^-] = 0.04948\ mol/L$
Rearrange the expression for the water constant to solve for $[H_3O^+]$.	$\left[H_3O^+ \right] = \dfrac{K_w}{[OH^-]}$ $= \dfrac{1.0 \times 10^{-14}\ \dfrac{mol^2}{L^2}}{0.04948\ \dfrac{mol}{L}}$ $= 2.02 \times 10^{-13}\ mol/L$

Check Your Solution
Multiplying the calculated $[H_3O^+]$ and $[OH^-]$ values gives the water constant, 1.0×10^{-14}.

16. A drain cleaner is a mixture of soluble hydroxides in solution. Would the drain cleaner be acidic, basic, or neutral? Explain your answer using one or two well-reasoned sentences.

Answer
Since the hydroxides are soluble, the solution would have a high concentration of hydroxide ions in solution. As a result, the hydronium ion concentration would be very low, which means that the solution is basic.

17. Verify that the hydronium ion, $H_3O^+(aq)$, concentration in a solution at 25°C containing 0.455 mol/L of hydroxide ions, $OH^-(aq)$, is 2.20×10^{-14} mol/L.

What Is Required?
You need to verify the $[H_3O^+]$ in a given solution.

What Is Given?
You know the $[OH^-]$ is 0.455 mol/L.
You know the value of $[H_3O^+]$ you need to verify is 2.20×10^{-14} mol/L.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Calculate a trial value of the K_w using the given values of $[H_3O^+]$ and $[OH^-]$.	trial $K_w = [H_3O^+][OH^-]$ $\quad = (2.20 \times 10^{-14} \text{mol/L}) (0.455 \text{ mol/L })$ $\quad = 1.0 \times 10^{-14}$
The $[H_3O^+]$ is correct if the trial value obtained is 1.0×10^{-14}.	Since the trial value calculated for K_w is the same as the accepted value, the given $[H_3O^+]$ is correct.

Check Your Solution
If you divide the water constant by the $[OH^-]$, 0.455 mol/L, you get a $[H_3O^+]$ of 2.20×10^{-14} mol/L. The number of significant digits is correct.

18. A solution has a hydronium ion, $H_3O^+(aq)$, concentration of 0.152 mol/L at 25°C. What is the hydroxide ion, $OH^-(aq)$, concentration of this solution?

What Is Required?
You need to find the find the $[OH^-]$ for a given solution.

What Is Given?
You know the $[H_3O^+]$ is 0.152 mol/L.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Rearrange the expression for the water constant to solve for $[OH^-]$.	$$\left[OH^- \right] = \frac{K_w}{[H_3O^+]}$$ $$= \frac{1.0 \times 10^{-14} \ \frac{mol}{L}}{0.152 \ \frac{mol}{L}}$$ $$= 6.58 \times 10^{-14} \, mol/L$$

Check Your Solution
Multiplying the calculated $[OH^-]$ by the given $[H_3O^+]$ gives the water constant, 1.0×10^{-14}.

19. What is the hydronium ion, $H_3O^+(aq)$, concentration of a solution at 25°C that has a hydroxide ion, $OH^-(aq)$, concentration of 0.0025 mol/L?

What Is Required?
You need to find the find the $[H_3O^+(aq)]$ for a given solution.

What Is Given?
You know the $[OH^-(aq)]$ is 0.0025 mol/L.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Rearrange the expression for the water constant to solve for $[H_3O^+]$.	$[H_3O^+] = \dfrac{K_w}{[OH^-]}$ $= \dfrac{1.0 \times 10^{-14}\,\dfrac{mol^{\cancel{2}}}{L^{\cancel{2}}}}{0.0025\dfrac{\cancel{mol}}{\cancel{L}}}$ $= 4.0 \times 10^{-12}\,mol/L$

Check Your Solution
Multiplying the calculated $[OH^-]$ by the given $[H_3O^+]$ gives the water constant, 1.0×10^{-14}.

20. A 1.55 mol/L hydrochloric acid solution, HCl(aq), is at 25°C. Determine the hydroxide ion, $OH^-(aq)$, concentration in this solution.

What Is Required?
You need to find the find the $[OH^-]$ for a given acid solution.

What Is Given?
You know the $[HCl] = 1.55$ mol/L.
You know the expression for the water constant is $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$.

Plan Your Strategy	Act on Your Strategy
Since HCl(aq) is a strong acid, $[H_3O^+] = [HCl]$. Rearrange the expression for the water constant to solve for $[OH^-]$.	$\left[OH^-\right] = \dfrac{K_w}{[H_3O^+]}$ $= \dfrac{1.0 \times 10^{-14} \ \dfrac{\text{mol}\cancel{/}}{\text{L}\cancel{/}}}{1.55 \ \dfrac{\text{mol}\cancel{/}}{\text{L}\cancel{/}}}$ $= 6.45 \times 10^{-15} \text{ mol/L}$

Check Your Solution
Multiplying the calculated $[OH^-]$ by the given $[H_3O^+]$ gives the water constant, 1.0×10^{-14}.

Calculating pH, pOH, and [OH⁻]
(Student textbook page 502)

21. An unknown solution has a pH of 5.84. If the solution is at 25°C, what is the pOH of the solution?

What Is Required?
You need to determine the pOH of a solution.

What Is Given?
You know the pH of the solution is 5.84.

Plan Your Strategy	Act on Your Strategy
Use the expression for pK_w to determine pH.	$pK_w = pH + pOH = 14.00$ $pH + pOH = 14.00$ $pH = 14.00 - 5.84$ $= 8.16$

Check Your Solution
Working in reverse, pK_w minus the pH gives the original pOH.

22. In a given solution, the pOH is 2.77. If the solution is at 25°C, what is the pH of the solution?

What Is Required?
You need to determine the pH of a solution.

What Is Given?
You know the pOH of the solution is 2.77.

Plan Your Strategy	Act on Your Strategy
Use the expression for pK_w to determine pH.	$pK_w = pH + pOH = 14.00$ $pH + pOH = 14.00$ $pH = 14.00 - 2.77$ $= 11.23$

Check Your Solution
Working in reverse, pK_w minus the pH gives the original pOH.

23. Determine the pH and pOH of a solution at 25°C with a hydronium ion, H_3O^+(aq), concentration of 3.20×10^{-10} mol/L.

What Is Required?
You need to determine the pH and pOH of a solution.

What Is Given?
You know the $[H_3O^+]$ for the solution is 3.20×10^{-10} mol/L (3 sig. digits).

Plan Your Strategy	Act on Your Strategy
Use the expression pH = –log $[H_3O^+]$ to determine the pH.	pH = –log $[H_3O^+]$ $\quad$ = –log (3.20×10^{-10}) $\quad$ = 9.495 (3 sig. digits after the decimal)
Use the expression for pK_w to calculate the pOH.	pK_w = pH + pOH = 14.00 pH + pOH = 14.00 $\quad\quad$ pOH = 14.00 – 9.495 $\quad\quad\quad\quad$ = 4.505

Check Your Solution
The low value of the $[H_3O^+$(aq)$]$ is consistent with the high pH value. Similarly, the pOH value would be expected to be low. The answers seem reasonable.

24. A solution is made by dissolving 0.45 mol of sodium hydroxide, NaOH(s), in enough water to make 3.75 L of solution. Determine the pH and pOH of this solution at 25°C.

What Is Required?
You must calculate the pH and pOH of a given solution.

What Is Given?
You know that 0.45 mol of NaOH(s) is in 3.75 L of solution.

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to calculate the molar concentration, c, of NaOH.	$c_{\text{NaOH}} = \dfrac{n}{V}$ $= \dfrac{0.45 \text{ mol}}{3.75 \text{ L}}$ $= 0.12 \text{ mol/L}$
Since NaOH is a strong base, [NaOH] = [OH⁻].	$[\text{OH}^-] = 0.12 \text{ mol/L}$
Calculate pOH using the equation pOH = –log [OH⁻].	$\text{pOH} = -\log [\text{OH}^-]$ $= -\log 0.12$ $= 0.92$
Use the expression for pK_w to determine the pH.	$pK_w = \text{pH} + \text{pOH} = 14.00$ $\text{pH} + \text{pOH} = 14.00$ $\text{pH} = 14.00 - 0.92$ $= 13.08$

Check Your Solution
The low pOH value is consistent with the high concentration of NaOH. Similarly, the high pH value is consistent with fairly high base concentration. The answers seem reasonable.

25. Determine the pH and pOH of a solution at 25°C that has 0.42 mol of hydroxide ions, $OH^-(aq)$, in 2.00 L of solution.

What Is Required?
You must calculate the pH and pOH of a given solution.

What Is Given?
You know that 0.45 mol of hydroxide ion is in 2.00 L of solution.

Plan Your Strategy	Act on Your Strategy
Use the formula $c = \dfrac{n}{V}$ to calculate the molar concentration, c, of the hydroxide ion.	$c_{OH^-} = \dfrac{n}{V}$ $= \dfrac{0.42 \text{ mol}}{2.00 \text{ L}}$ $= 0.21 \text{ mol/L}$
Calculate pOH using the equation $pOH = -\log [OH^-]$.	$pOH = -\log [OH^-]$ $= -\log 0.21$ $= 0.68$
Use the expression for pK_w to determine the pH.	$pK_w = pH + pOH = 14.00$ $pH + pOH = 14.00$ $pH = 14.00 - 0.68$ $= 13.32$

Check Your Solution
The low pOH value is consistent with the high concentration of $OH^-(aq)$. Similarly, the high pH value is consistent with fairly high base concentration. The answers seem reasonable.

26. Calculate the pH and pOH of a solution at 25°C that has a hydroxide ion, $OH^-(aq)$, concentration of 1.74×10^{-9} mol/L.

What Is Required?
You must calculate the pH and pOH of a given solution.

What Is Given?
You know the hydroxide ion concentration, $[OH^-]$, is 1.74×10^{-9} mol/L.

Plan Your Strategy	Act on Your Strategy
Calculate pOH using the equation $pOH = -\log [OH^-]$.	$\begin{aligned} pOH &= -\log [OH^-] \\ &= -\log (1.74 \times 10^{-9}) \\ &= 8.759 \end{aligned}$
Use the expression for pK_w to determine the pH.	$pK_w = pH + pOH = 14.00$ $\begin{aligned} pH + pOH &= 14.00 \\ pH &= 14.000 - 8.759 \\ &= 5.241 \end{aligned}$

Check Your Solution
The pH and pOH values are consistent with the concentration of their respective ions. The answers seem reasonable.

27. What are the pH and pOH at 25°C of 0.097 mol/L nitric acid, $HNO_3(aq)$, a strong acid?

What Is Required?
You need to determine the pH and pOH of a solution.

What Is Given?
You know the nitric acid concentration is 0.097 mol/L.

Plan Your Strategy	Act on Your Strategy
Since $HNO_3(aq)$ is a strong acid, $[H_3O^+] = [HNO_3]$.	$[H_3O^+] = [HNO_3]$ $\quad = 0.097$ mol/L
Use the expression pH = $-\log [H_3O^+]$ to determine the pH.	pH = $-\log [H_3O^+]$ $\quad = -\log 0.097$ $\quad = 1.01$
Use the expression for pK_w to calculate the pOH.	$pK_w = $ pH $+$ pOH $= 14.00$ pH $+$ pOH $= 14.00$ $\quad\quad$ pOH $= 14.00 - 1.01$ $\quad\quad\quad\quad = 12.99$

Check Your Solution
The relatively high value of the $H_3O^+(aq)$ is consistent with the low pH value. Similarly, the pOH value would be expected to be high. The answers seem reasonable.

28. A solution at 25°C is made by adding 0.083 g of calcium hydroxide, $Ca(OH)_2(s)$, (a soluble solid) into enough water to make 125 mL of solution. What are the pH and pOH of the resulting solution?

What Is Required?
You must calculate the pH and pOH of a given solution.

What Is Given?
You know that 0.083 g $Ca(OH)_2(s)$ is dissolved in 125 mL of solution.

Plan Your Strategy	Act on Your Strategy
Use the formula $n = \dfrac{m}{M}$ to calculate the amount in moles, n, of $Ca(OH)_2(s)$.	$n_{Ca(OH)_2} = \dfrac{m}{M}$ $= \dfrac{0.083 \text{ g}}{74.1 \text{ g /mol}}$ $= 0.00112 \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration, c, of $Ca(OH)_2(aq)$.	$c_{Ca(OH)_2} = \dfrac{n}{V}$ $= \dfrac{0.00112 \text{ mol}}{0.125 \text{ L}}$ $= 0.00896 \text{ mol/L}$
Use the mole ratio $\dfrac{OH^-}{Ca(OH)_2}$ to determine the $[OH^-]$.	$\dfrac{\text{mol } OH^-}{1 \text{ mol } Ca(OH)_2} = \dfrac{2}{1}$ $[OH^-] = 2 \times [Ca(OH)_2]$ $\qquad = 2 \times 0.008\ 96 \text{ mol/L}$ $\qquad = 0.01792 \text{ mol/L}$
Calculate pOH using the equation $pOH = -\log [OH^-]$	$pOH = -\log [OH^-]$ $\qquad = -\log 0.01792$ $\qquad = 1.75$
Use the expression for pK_w to determine the pH.	$pK_w = pH + pOH = 14.00$ $pH + pOH = 14.00$ $\qquad pH = 14.00 - 1.75$ $\qquad\quad = 12.25$

Check Your Solution
The low pOH value is consistent with the high concentration of $OH^-(aq)$. Similarly, the high pH value is consistent with fairly high base concentration. The answers seem reasonable.

29. What iIs the hydronium ion, H_3O^+(aq), concentration of a solution at 25°C with a pOH of 7.95?

What Is Required?
You need to calculate the concentration of H_3O^+(aq) in a solution of known pOH.

What Is Given?
You know the pOH of the solution is 7.95.

Plan Your Strategy	Act on Your Strategy
Use the pK_a to determine the pH of the solution.	$pK_w = pH + pOH = 14.00$ $pH + pOH = 14.00$ $\qquad pH = 14.00 - 7.95$ $\qquad\quad = 6.05$
Calculate $[H_3O^+]$ using the expression $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $\qquad = 10^{-6.05}$ $\qquad = 8.91 \times 10^{-7}$ mol/L

Check Your Solution
The concentration seems reasonable for a solution having a pOH (and pH) close to 7.

30. Explain why pH + pOH is equal to 14.00 when any aqueous solution is at 25°C.

Answer
Water undergoes autoionization:
$2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
Regardless of the temperature, the hydronium ion concentration and the hydroxide ion concentration in pure water are always equal.
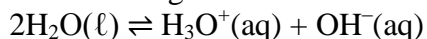
At 25°C, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L

$$K_w = [H_3O^+][OH^-]$$
$$1.0 \times 10^{-14} = (1.0 \times 10^{-7}\text{mol/L})(1.0 \times 10^{-7}\text{mol/L})$$
$$-\log K_w = (-\log [H_3O^+]) + (-\log [OH^-])$$
$$pK_w = pH + pOH$$
$$14.00 = 7.00 + 7.00$$

At another temperature, the value of K_w will be different and pH + pOH $\neq$ 14.

Predicting the Direction of an Acid-Base Reaction
(Student textbook page 508)

31. Predict the direction for the following reaction. State whether reactants or products are favoured, and give reasons to support your decision.

$$NH_4^+(aq) + H_2PO_4^-(aq) \rightarrow NH_3(aq) + H_3PO_4(aq)$$

What Is Required?
You must predict the direction of the reaction and give a reason for your answer.

What Is Given?
You have a balanced acid-base reaction.

Plan Your Strategy	Act on Your Strategy
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$NH_4^+(aq) + H_2PO_4^-(aq) \rightleftharpoons NH_3(aq) + H_3PO_4(aq)$ $\quad$ Acid$_1$ $\qquad\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_2$ (H_3PO_4) is stonger than Acid$_1$ (NH_4^+). Reactants are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

32. Predict the direction for the following reaction. State whether reactants or products are favoured, and give reasons to support your decision.

$$H_2O(\ell) + HS^-(aq) \rightarrow OH^-(aq) + H_2S(aq)$$

What Is Required?
You must predict the direction of the reaction and give a reason for your answer.

What Is Given?
You have a balanced acid-base reaction.

Plan Your Strategy	Act on Your Strategy
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$H_2O(\ell) + HS^-(aq) \rightleftharpoons OH^-(aq) + H_2S(aq)$ $\quad$ Acid $_1$ $\qquad\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_2$ (H_2S) is stonger than Acid$_1$ (H_2O). Reactants are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

33. Predict the direction for the following reaction. State whether reactants or products are favoured, and give reasons to support your decision.

$$HF(aq) + SO_4^{2-}(aq) \rightarrow F^-(aq) + HSO_4^-(aq)$$

What Is Required?
You must predict the direction of the reaction and give a reason for your answer.

What Is Given?
You have a balanced acid-base reaction.

Plan Your Strategy	Act on Your Strategy
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$HF(aq) + SO_4^{2-}(aq) \rightleftharpoons F^-(aq) + HSO_4^-(aq)$ $\quad$ Acid$_1$ $\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_2$ (HSO_4^-) is stonger than Acid$_1$ (HF). Reactants are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

34. In which direction will the following reaction proceed? Explain your decision.

$$HPO_4^{2-}(aq) + NH_4^+(aq) \rightleftharpoons H_2PO_4^-(aq) + NH_3(aq)$$

What Is Required?
You must predict the direction of the reaction and explain your reasoning.

What Is Given?
You have a balanced acid-base reaction.

Plan Your Strategy	Act on Your Strategy
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$HPO_4^{2-}(aq) + NH_4^+(aq) \rightleftharpoons H_2PO_4^-(aq) + NH_3(aq)$ $\quad$ Acid$_1$ $\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_2$ ($H_2PO_4^-$) is stronger than Acid$_1$ (HPO_4^{2-}). Reactants are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

35. In which direction will the following reaction proceed? Explain your decision.

$$H_2SO_4(aq) + H_2O(\ell) \rightleftharpoons HSO_4{}^-(aq) + H_3O^+(aq)$$

What Is Required?
You must predict the direction of the reaction and explain your reasoning.

What Is Given?
You have a balanced acid-base reaction.

Plan Your Strategy	Act on Your Strategy
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$H_2SO_4(aq) + H_2O(\ell) \rightleftharpoons HSO_4{}^-(aq) + H_3O^+(aq)$ $\quad$ Acid$_1$ $\qquad\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_1$ (H_2SO_4) is a stronger acid than Acid$_2$ (H_3O^+). Products are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

36. In which direction will the following reaction proceed? Explain your decision.

$$H_2S(aq) + NH_3(aq) \rightleftharpoons HS^-(aq) + NH_4^+(aq)$$

What Is Required?
You must predict the direction of the reaction and explain your reasoning.
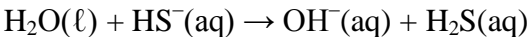
What Is Given?
You have a balanced acid-base reaction.

Plan Your Strategy	Act on Your Strategy
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$H_2S(aq) + NH_3(aq) \rightleftharpoons HS^-(aq) + NH_4^+(aq)$ $\quad$ Acid$_1$ $\qquad\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_1$ (H_2S) is a stronger acid than Acid$_2$ (NH_4^+). Products are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

37. Write the equilibrium equation for the following reaction. State whether the reactants or the products are favoured. (Use **Table 8.1** or the table in Appendix B.)

Aqueous carbonic acid, $H_2CO_3(aq)$, is combined with ammonia, $NH_3(g)$.

What Is Required?
You must write the equilibrium equation and predict the direction of the reaction.

What Is Given?
You know the reactants are $H_2CO_3(aq)$ and $NH_3(g)$.

Plan Your Strategy	Act on Your Strategy
Write the equilibrium equation.	$H_2CO_3(aq) + NH_3(g) \rightleftharpoons HCO_4^-(aq) + NH_4^+(aq)$
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$H_2CO_3(aq) + NH_3(g) \rightleftharpoons HCO_4^-(aq) + NH_4^+(aq)$ $\quad$ Acid$_1$ $\qquad\qquad\qquad\qquad\qquad$ Acid$_2$ Acid$_1$ (H_2CO_3) is a stronger acid than Acid$_2$ (NH_4^+). Products are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

38. Write the equilibrium equation for the following reaction. State whether the reactants or the products are favoured. (Use the table in Appendix B (student textbook page 747).)

Sodium hydrogen sulfite, $NaHSO_3$(s), is dissolved in water. (Hint: The sodium hydrogen sulfite dissociates completely into sodium ions and hydrogen sulfite ions.) What happens to the hydrogen sulfite ion in water?

What Is Required?
You must write the equilibrium equation and predict the direction of the reaction.

What Is Given?
You know the reactants are $NaHSO_3$(s) and $H_2O(\ell)$.
You know that $NaHSO_3$(s) dissociates completely in water, producing Na^+(aq) and HSO_3^-(aq) ions.

Plan Your Strategy	Act on Your Strategy
Write the equilibrium equation.	HSO_3^-(aq) + $H_2O(\ell)$ ⇌ SO_3^{2-} (aq) + H_3O^+(aq)
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	HSO_3^-(aq) + $H_2O(\ell)$ ⇌ SO_3^{2-} (aq) + H_3O^+(aq) Acid$_1$ Acid$_2$ Acid$_2$ (H_3O^+) is stronger than Acid$_1$ (HSO_3^-). Reactants are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

39. Write the equilibrium equation for the following reaction. State whether the reactants or the products are favoured. (Use **Table 8.1** or the table in Appendix B.)

Hydrofluoric acid, HF(aq), is mixed with potassium nitrate, KNO_3(s). (Consider the hint in question 38.)

What Is Required?
You must write the equilibrium equation and predict the direction of the reaction.

What Is Given?
You know the reactants are HF(aq) and KNO_3(s).
You know that KNO_3(s) dissociates completely in water, producing K^+(aq) and NO_3^-(aq) ions.

Plan Your Strategy	Act on Your Strategy
Write the equilibrium equation.	NO_3^-(aq) + HF(aq) $\rightleftharpoons$ HNO_3(aq) + F^-(aq)
Identify the acid on each side of the equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	NO_3^-(aq) + HF(aq) $\rightleftharpoons$ HNO_3(aq) + F^-(aq) $\qquad\qquad$ Acid$_1$ $\qquad$ Acid$_2$ Acid$_2$ (HNO_3) is stronger than Acid$_1$ (HF). Reactants are favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

40. Write the equilibrium equation for the following reaction. State whether the reactants or the products are favoured. (Use **Table 8.1** or the table in Appendix B.)

Write the reactions for the ionizations of the three hydrogen ions in $H_3PO_4(\ell)$.

What Is Required?
You must write the equilibrium equation and predict the direction for each reaction.

What Is Given?
You know the reactants for the first ionization are $H_3PO_4(\ell)$ and $H_2O(\ell)$.
You know the reactants for the second ionization are $H_2PO_4^-(aq)$ and $H_2O(\ell)$.
You know the reactants for the third ionization are $HPO_4^{2-}(aq)$ and $H_2O(\ell)$.

Plan Your Strategy	Act on Your Strategy
Write the equilibrium equation for each of the three ionizations.	$H_3PO_4(aq) + H_2O(\ell) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$ $H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$ $HPO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$
Identify the acid on each side of each equation. Use Appendix B to determine which acid is stronger. The reaction will proceed from the side with the stronger acid toward the side with the weaker acid.	$H_3PO_4(aq) + H_2O(\ell) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$ $\text{Acid}_1 \qquad\qquad\qquad\qquad\qquad\qquad \text{Acid}_2$ Acid_2 is stronger than Acid_1, so the formation of reactants is favoured. $H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$ $\text{Acid}_1 \qquad\qquad\qquad\qquad\qquad\qquad \text{Acid}_2$ Acid_2 is stronger than Acid_1, so the formation of reactants is favoured. $HPO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$ $\text{Acid}_1 \qquad\qquad\qquad\qquad\qquad\qquad \text{Acid}_2$ Acid_2 is stronger than Acid_1, so the formation of reactants is favoured.

Check Your Solution
The stronger acid gives up a proton more readily. The reaction proceeds from the stronger to the weaker acid.

Determining the Acid-Dissociation Constant, K_a, and Percent Dissociation
(Student textbook page 512)

41. In low doses, barbiturates act as sedatives. Barbiturates are made from barbituric acid, a weak monoprotic acid that was first prepared by the German chemist Adolph von Baeyer in 1864. The formula for barbituric acid is $HC_4H_3N_2O_3(s)$. A chemist prepares a 0.10 mol/L solution of barbituric acid. The chemist measured the pH of the solution and recorded the value as 2.50. What is the acid-dissociation constant for barbituric acid? What percentage of its molecules dissociated into ions?

What Is Required?
You need to calculate the acid-dissociation constant, K_a, and the percentage of the acid molecules that have dissociated.

What Is Given?
You know the original concentration of the barbituric acid is 0.10 mol/L.
You know the pH of the solution is 2.50.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $\quad = 10^{-2.50}$ $\quad = 3.16 \times 10^{-3}$ mol/L (carry 3 sig. digits)
Write a balanced chemical equation representing the acid-dissociation using a symbol for barbituric acid such as HBa. Set up an ICE table showing the initial [HBa] = 0.10 mol/L and the initial $[H_3O^+]$ ~0 mol/L. Use the stoichiometry of the equation and the $[H_3O^+]$ calculated in the previous step to complete the ICE table.	See the ICE table below.

	HBa(aq) +	$H_2O(\ell)$	$\rightleftharpoons$	$H_3O^+(aq)$ +	$Ba^-(aq)$
	[HBa] (mol/L)			$[H_3O^+]$ (mol/L)	$[Ba^-]$ (mol/L)
Initial	0.10			~0	0
Change	-3.16×10^{-3}			$+3.16 \times 10^{-3}$	$+3.16 \times 10^{-3}$
Equilibrium	$0.10 - 3.16 \times 10^{-3}$			3.16×10^{-3}	3.16×10^{-3}

Write the expression for K_a and solve for K_a.	$K_a = \dfrac{[\text{H}_3\text{O}^+][\text{Ba}^-]}{[\text{HBa}]}$ $= \dfrac{(3.16\times10^{-3})(3.16\times10^{-3})}{0.09684}$ $= 1.0 \times 10^{-4}\,(2\text{ sig. digits})$
Calculate the percent dissociation.	$\text{percent dissociation} = \dfrac{[\text{H}_3\text{O}^+]}{[\text{HBa}]}\times100\%$ $= \dfrac{3.16\times10^{-3}\ \cancel{\text{mol/L}}}{0.10\ \cancel{\text{mol/L}}}\times100\%$ $= 3.2\%$

Check Your Solution

The answers are expressed to the correct number of significant digits, and both the K_a value and the percent dissociation seem to be reasonable for a weak acid.

42. The word "butter" comes from the Greek *butyros*. Butanoic acid, $HC_4H_7O_2(aq)$ (common name butyric acid), gives rancid butter its distinctive odour. If the pH of a 1.00×10^{-2} mol/L solution of butanoic acid is 3.41, calculate the acid-dissociation constant for butyric acid. What percentage of butyric acid molecules in this solution ionized?

What Is Required?
You need to calculate the acid-dissociation constant, K_a, and the percentage of the acid molecules that have dissociated.

What Is Given?
You know the original concentration of the butanoic acid is 1.00×10^{-2} mol/L.
You know the pH of the solution is 3.41.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$\begin{aligned}[H_3O^+] &= 10^{-pH} \\ &= 10^{-3.41} \\ &= 3.89 \times 10^{-4} \text{ mol/L (carry 3 sig. digits)}\end{aligned}$
Write a balanced chemical equation representing the acid-dissociation using a symbol for butanoic acid such as HBu. Set up an ICE table showing the initial $[HBu] = 1.00 \times 10^{-2}$ mol/L and the initial $[H_3O^+] \sim 0$ mol/L. Use the stoichiometry of the equation and the $[H_3O^+]$ calculated in the previous step to complete the ICE table.	See the ICE table below.

	HBu(aq)	+	H₂O(ℓ)	⇌	H_3O^+(aq)	+	Bu⁻(aq)
	[HBu] (mol/L)				$[H_3O^+]$ (mol/L)		[Bu⁻] (mol/L)
I	1.00×10^{-2}				~0		0
C	$- 3.89 \times 10^{-4}$				$+3.89 \times 10^{-4}$		$+3.89 \times 10^{-4}$
E	$1.00 \times 10^{-2} - 3.89 \times 10^{-3}$				3.89×10^{-4}		3.89×10^{-4}

Write the expression for K_a and solve for K_a.	$\begin{aligned} K_a &= \frac{[H_3O^+][Bu^-]}{[HBu]} \\ &= \frac{(3.89\times10^{-4})(3.89\times10^{-4})}{0.009611} \\ &= 1.6 \times 10^{-5} \end{aligned}$

Calculate the percent dissociation.	$\text{percent dissociation} = \dfrac{[H_3O^+]}{[HBu]} \times 100\%$
	$= \dfrac{3.89 \times 10^{-4} \ \cancel{mol/L}}{1.00 \times 10^{-2} \ \cancel{mol/L}} \times 100\%$
	$= 3.9\%$

Check Your Solution

The answer is expressed to the correct number of significant digits (since the pH value given in the question has two significant digits to the right of the decimal) and seems to be reasonable for the K_a value for a weak acid.

43. Wild almonds taste bitter (and are dangerous to eat!) because they contain hydrocyanic acid, $HCN(aq)$. When a chemist prepared a 0.75 mol/L solution of $HCN(aq)$, the pH was found to be 4.67. What is the acid-dissociation constant, K_a?

What Is Required?
You need to calculate the acid-dissociation constant, K_a, for hydrocyanic acid.

What Is Given?
You know the original concentration of the hydrocyanic acid is 0.75 mol/L.
You know the pH of the solution is 4.67.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-4.67}$ $= 2.14 \times 10^{-5}$ mol/L (carry 3 sig. digits)
Write a balanced chemical equation representing the acid-dissociation. Set up an ICE table showing the initial $[HCN] = 0.75$ mol/L and the initial $[H_3O^+] \sim 0$ mol/L. Use the stoichiometry of the equation and the $[H_3O^+]$ calculated in the previous step to complete the ICE table.	See the ICE table below.

	HCN(aq)	+	H₂O(ℓ)	⇌	H₃O⁺(aq)	+	CN⁻(aq)
	[HCN] (mol/L)				[H₃O⁺] (mol/L)		[CN⁻] (mol/L)
I	0.75				~0		0
C	-2.14×10^{-5}				$+2.14 \times 10^{-5}$		$+2.14 \times 10^{-5}$
E	$0.75 - 2.14 \times 10^{-5}$				2.14×10^{-5}		2.14×10^{-5}

Write the expression for K_a and solve for K_a.	$K_a = \dfrac{[H_3O^+][CN^-]}{[HCN]}$ $= \dfrac{(2.14 \times 10^{-5})(2.14 \times 10^{-5})}{0.7498}$ $= 6.1 \times 10^{-10}$

Check Your Solution
The answer is expressed to the correct number of significant digits and seems to be reasonable for the K_a value for a weak acid.

44. Some sunscreen lotions contain salts of para-aminobenzoic acid (PABA). The structure of PABA is shown below. A saturated solution of PABA was prepared by dissolving 4.7g in a 1.0 L solution. The pH of the solution was found to be 3.19. Calculate the acid-dissociation constant, K_a, for PABA.

What Is Required?
You need to calculate the acid-dissociation constant, K_a, for PABA.

What Is Given?
You know the mass of PABA dissolved is 4.7 g /L of solution and the pH of the solution is 3.19.

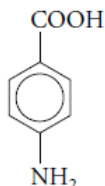| Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-3.19}$ $= 6.46 \times 10^{-4}$ mol/L (carry 3 sig. digits)
Use the molar mass, M, of PABA and the formula $n = \dfrac{m}{M}$ to calculate the amount in moles, n, of PABA.	$n_{PABA} = \dfrac{m}{M}$ $= \dfrac{4.7\ \cancel{g}}{139.15\ \cancel{g}/mol}$ $= 0.03427$ mol
Use the formula $c = \dfrac{n}{V}$ to determine the initial molar concentration, c, of PABA.	$c_{PABA} = \dfrac{n}{V}$ $= \dfrac{0.03427\ mol}{1.0\ L}$ $= 0.03427$ mol/L
Write a balanced equation representing the acid-dissociation. Set up an ICE table showing the initial [PABA] and the initial $[H_3O^+]$ ~0 mol/L. Use the stoichiometry of the equation and the $[H_3O^+]$ calculated in the first step, above, to complete the ICE table.	See the ICE table on the following page.

	PABA(aq)	+	H₂O(ℓ)	⇌	H₃O⁺(aq)	+	PABA⁻(aq)

	[PABA] (mol/L)		[H$_3$O$^+$] (mol/L)	[PABA$^-$] (mol/L)
I	0.03427		~0	0
C	-6.46×10^{-4}		$+6.46 \times 10^{-4}$	$+6.46 \times 10^{-4}$
E	$0.03427 - 6.46 \times 10^{-4}$		6.46×10^{-4}	6.46×10^{-4}

Write the expression for K_a and solve for K_a.	$\begin{aligned} K_a &= \dfrac{[\text{H}_3\text{O}^+][\text{PABA}^-]}{[\text{PABA}]} \\ &= \dfrac{(6.46\times10^{-4})(6.46\times10^{-4})}{0.03362} \\ &= 1.2\times10^{-5} \end{aligned}$

Check Your Solution

The answer is expressed to the correct number of significant digits and seems to be reasonable for the K_a value for a weak acid.

45. Aspirin™ (acetylsalicylic acid) is a monoprotic acid with molar mass of 180 g/mol. An aqueous solution containing 3.3 g/L has a pH of 2.62. What percentage of acetylsalicylic acid molecules ionize in the solution?

What Is Required?
You need to calculate the percentage of the acid molecules that have ionized.

What Is Given?
You know the mass concentration of acetylsalicylic acid (ASA) is 3.3 g/L.
You know the pH of the solution is 2.62.
You know the molar mass is 180 g/mol.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-2.62}$ $= 2.40 \times 10^{-3}$ mol/L (carry 3 sig. digits)
Use the mass concentration and molar mass to find the molar concentration.	$M = \dfrac{3.3 \text{ g/L}}{180 \text{ g/mol}}$ $= 0.0183 \text{ mol/L}$
Calculate the percent dissociation.	$\text{percent dissociation} = \dfrac{[H_3O^+]}{[ASA]} \times 100\%$ $= \dfrac{2.40 \times 10^{-3} \text{ mol/L}}{0.0183 \text{ mol/L}}$ $= 0.1311$ $= 0.13 \times 100\%$ $= 13\%$

Check Your Solution
The answer is expressed to the correct number of significant digits and is a fairly low percent ionization for a weak acid which is reasonable.

46. Quinoxaline is a heterocyclic compound with the molecular formula $C_8H_6N_2$. It is used in the manufacturing of dyes, pharmaceuticals, and antibiotics. When a chemist dissolved 7.19 g of quinoxaline in a 1.0 L solution, the pH of the solution was 1.33. Calculate the acid-dissociation constant, K_a, for quinoxaline.

What Is Required?
You need to calculate the acid-dissociation constant, K_a, for quinoxaline.

What Is Given?
You know the mass of quinoxaline dissolved is 7.19 g.
The volume of the solution is 1L.
You know the pH of the solution is 1.33.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-1.33}$ $= 4.68 \times 10^{-2}$ mol/L (carry 3 sig. digits)
Use the molar mass, M, of quinoxaline and the formula $n = \dfrac{m}{M}$ to calculate the amount in moles, n, of quinoxaline. Use the periodic table to calculate the molar mass. The molar mass of quinoxaline (HQ) equals 130.16 g/mol.	$n_{HQ} = \dfrac{m}{M}$ $= \dfrac{7.19\,\cancel{g}}{130.16\,\cancel{g}\,/\text{mol}}$ $= 0.05524$ mol
Use the formula $c = \dfrac{n}{V}$ to determine the initial molar concentration, c, of quinoxaline.	$c_{HQ} = \dfrac{n}{V}$ $= \dfrac{0.05524 \text{ mol}}{1.0 \text{ L}}$ $= 0.05524$ mol/L
Write a balanced equation representing the acid-dissociation using the symbol HQ to represent this acid. Set up an ICE table showing the initial [HQ] and the initial $[H_3O^+] \sim 0$ mol/L. Use the stoichiometry of the equation and $[H_3O^+]$ to complete the ICE table.	See the ICE table below.

	HQ(aq)	+	H$_2$O(ℓ)	⇌	H$_3$O$^+$(aq)	+	Q$^-$(aq)
	[HQ] (mol/L)				$[H_3O^+]$ (mol/L)		$[Q^-]$ (mol/L)
I	0.05524				~0		0
C	-4.68×10^{-2}				$+4.68 \times 10^{-2}$		$+4.68 \times 10^{-2}$
E	$0.05524 - 4.68 \times 10^{-2}$				4.68×10^{-2}		4.68×10^{-2}

Write the expression for K_a and solve for $K_{a.}$	$K_a = \dfrac{[H_3O^+][Q^-]}{[HQ]}$
	$= \dfrac{(4.68 \times 10^{-2})(4.68 \times 10^{-2})}{0.00844}$
	$= 0.26$

Check Your Solution

The answer is expressed to the correct number of significant digits and he K_a seems to be reasonable for a weak acid.

47. Caprytic acid, also called octanoic acid, is an 8-carbon fatty acid that is used as a sanitizer on dairy equipment that comes in contact with milk. The acid has a molar mass of 144.21 g/mol. An aqueous solution of the sanitizer containing 12.25 g/L has a pH of 2.98. What Is the acid-dissociation constant, K_a, for caprytic acid?

What Is Required?
You need to calculate the acid-dissociation constant, K_a, for caprytic acid.

What Is Given?
You know the mass concentration of caprytic acid is 12.25 g/L.
The molar mass of caprytic acid is 144.21g/mol.
The pH of the solution is 2.98.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH} = 10^{-2.98}$ $= 1.05 \times 10^{-3}$ mol/L (carry 3 sig. digits)
Use the mass concentration and molar mass to find the molar concentration of caprytic acid.	$M = \dfrac{12.25 \text{ g/L}}{144.21 \text{ g/mol}}$ $= 0.0849$ mol/L
Write a balanced equation representing the acid-dissociation using the symbol HCa to represent this acid. Set up an ICE table showing the initial [HCa] and the initial $[H_3O^+] \sim 0$ mol/L. Use the stoichiometry of the equation and $[H_3O^+]$ to complete the ICE table.	See the ICE table below.

	HCa(aq)	+	H₂O(ℓ)	⇌	H₃O⁺(aq)	+	Ca⁻(aq)
	[HCa] (mol/L)				$[H_3O^+]$ (mol/L)		$[Ca^-]$ (mol/L)
I	0.0849				~0		0
C	-1.05×10^{-3}				$+1.05 \times 10^{-3}$		$+1.05 \times 10^{-3}$
E	$0.0849 - 1.05 \times 10^{-3}$				1.05×10^{-3}		1.05×10^{-3}

Write the expression for K_a and solve for K_a.	$K_a = \dfrac{[\text{H}_3\text{O}^+][\text{Ca}^-]}{[\text{HCa}]}$ $= \dfrac{(1.05 \times 10^{-3})(1.05 \times 10^{-3})}{0.0838}$ $= 1.32 \times 10^{-5}$

Check Your Solution
The answer is expressed to the correct number of significant digits and he K_a seems to be reasonable for a weak acid.

48. Citric acid, $C_6H_8O_7$, is used to add a sour taste to foods and soft drinks. It is prepared in a solution with a concentration of 0.52 mol/L, and has a pH of 1.72. Calculate the acid-dissociation constant for citric acid. What percentage of citric acid molecules ionizes in the solution?

What Is Required?
You need to calculate the acid-dissociation constant, K_a, for citric acid and the percentage of the acid molecules that have dissociated.

What Is Given?
You know the original concentration of the citric acid is 0.52 mol/L.
You know the pH of the solution is 1.72.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-1.72}$ $= 0.0190$ mol/L (carry 3 sig. digits)
Write a balanced equation representing the acid-dissociation using a symbol for citric acid such as HCi. Set up an ICE table showing the initial [HCi] = 0.52 mol/L and the initial $[H_3O^+]$ ~0 mol/L. Use the stoichiometry of the equation and $[H_3O^+]$ to complete the ICE table.	See the ICE table below.

	HCi(aq)	+	H₂O(ℓ)	⇌	H₃O⁺(aq)	+	Ci⁻(aq)
	[HCi] (mol/L)				$[H_3O^+]$ (mol/L)		$[Ci^-]$ (mol/L)
I	0.52				~0		0
C	−0.0190				+0.0190		+0.0190
E	0.52 − 0.0190				0.0190		0.0190

Write the expression for K_a and solve for K_a.	$K_a = \dfrac{[H_3O^+][Ci^-]}{[HCi]}$ $= \dfrac{(0.0190)(0.0190)}{0.501}$ $= 7.2 \times 10^{-4}$

Calculate the percent dissociation.	$\text{percent dissociation} = \dfrac{[H_3O^+]}{[HCi]} \times 100\%$
	$= \dfrac{0.0190 \; \cancel{mol/L}}{0.52 \; \cancel{mol/L}} \times 100\%$
	$= 3.7\%$

Check Your Solution

The answer is expressed to the correct number of significant digits and the K_a and percent dissociation seem reasonable for a weak acid.

49. Creatine is a naturally occurring organic acid that helps supply energy to all cells in the body, with particular emphasis on muscle cells. It has a molar mass of 131 g/mol. An aqueous solution containing 2.68 g/L has a pH of 2.23. What percentage of creatine molecules ionize in the solution?

What Is Required?
You need to calculate the percentage of the creatine molecules that have ionized.

What Is Given?
You know the mass of creatine dissolved is 2.68 g/L of solution.
You know the pH of the solution is 2.23.
You know the molar mass is 131 g/mol.

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-2.23}$ $= 5.89 \times 10^{-3}$ mol/L (carry 3 sig. digits)
Use the molar mass, M, of creatine and the formula $n = \dfrac{m}{M}$ to calculate the amount in moles, n, of creatine.	$n_{creatine} = \dfrac{m}{M}$ $= \dfrac{2.68 \cancel{g}}{131 \cancel{g}/\text{mol}}$ $= 0.02046$ mol
Use the formula $c = \dfrac{n}{V}$ to determine the initial molar concentration, c, of creatine.	$c_{creatine} = \dfrac{n}{V}$ $= \dfrac{0.02046 \text{ mol}}{1.0 \text{ L}}$ $= 0.02046$ mol/L
Calculate the percent dissociation.	percent dissociation $= \dfrac{[H_3O^+]}{[\text{creatine}]} \times 100\%$ $= \dfrac{5.89 \times 10^{-3} \cancel{\text{mol/L}}}{0.02046 \cancel{\text{mol/L}}} \times 100\%$ $= 29\%$

Check Your Solution
The answer is expressed to the correct number of significant digits and the percent dissociation is reasonable for the given solubility of creatine.

50. Uracil is one of the four nucleotide monomers that make up the nucleic acid RNA. It was originally isolated in 1900, by the hydrolysis of yeast nuclein found in bovine thymus and spleen. Uracil has a molar mass of 112 g/mol. A biologist prepared a solution of the acid by dissolving 1.24 g of the compound in 1.0 L of solution. The pH of the solution is 5.71. Calculate the acid-dissociation constant, K_a, for uracil.

What Is Required?
You need to calculate the acid-dissociation constant, K_a, for uracil.

What Is Given?
You know the molar mass of uracil is 112 g/mol.
You know the initial mass of uracil to prepare the solution is 1.24 g.
You know the initial volume of the solution is 1.0 L.
You know the pH of the solution is 5.71

Plan Your Strategy	Act on Your Strategy
Determine the hydronium ion concentration using the equation $[H_3O^+] = 10^{-pH}$.	$[H_3O^+] = 10^{-pH}$ $= 10^{-5.71}$ $= 1.95 \times 10^{-6}$ mol/L (carry 3 sig. digits)
Use the molar mass, M, of uracil and the formula $n = \dfrac{m}{M}$ to calculate the amount in moles, n, of uracil.	$n_{HU} = \dfrac{m}{M}$ $= \dfrac{1.24 \cancel{g}}{112 \cancel{g}/mol}$ $= 0.01107$ mol
Use the formula $c = \dfrac{n}{V}$ to determine the initial molar concentration, c, of uracil.	$c_{HU} = \dfrac{n}{V}$ $= \dfrac{0.01107 \text{ mol}}{1.0 \text{ L}}$ $= 0.01107$ mol/L
Write a balanced equation representing the acid-dissociation using the symbol HU to represent uracil. Set up an ICE table showing the initial [HU] and the initial $[H_3O^+]$ ~0 mol/L. Use the stoichiometry of the equation and $[H_3O^+]$ to complete the ICE table.	See the ICE table below.

	HU(aq)	+	H₂O(ℓ)	⇌	H₃O⁺(aq)	+	U⁻(aq)
	[HU] (mol/L)				$[H_3O^+]$ (mol/L)		$[U^-]$ (mol/L)
I	0.01107				~0		0
C	-1.95×10^{-6}				$+1.95 \times 10^{-6}$		$+1.95 \times 10^{-6}$
E	$0.01107 - 1.95 \times 10^{-6}$				1.95×10^{-6}		1.95×10^{-6}

| Write the expression for K_a and solve for K_a.	$K_a = \dfrac{[H_3O^+][U^-]}{[HU]}$ $= \dfrac{(1.95 \times 10^{-6})(1.95 \times 10^{-6})}{0.01107}$ $= 3.4 \times 10^{-10}$

Check Your Solution
The answer is expressed to the correct number of significant digits and the K_a is small, which is reasonable for an acid having this pH.

Determining pH Using K_a
(Student textbook page 514)

51. A 0.075 mol/L solution of chlorous acid, $HClO_2(aq)$, is made in a lab. It has a K_a value of 1.15×10^{-2}, with an equilibrium reaction given as follows:
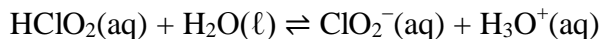
$$HClO_2(aq) + H_2O(\ell) \rightleftharpoons ClO_2^-(aq) + H_3O^+(aq)$$

What Is the pH of the solution?

What Is Required?
You must determine the pH of the solution.

What Is Given?
For $HClO_2(aq)$, you know the initial concentration is 0.075 mol/L.
You know $HClO_2(aq)$ has a K_a of 1.15×10^{-2}.
You have a balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of $HClO_2(aq)$ that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial $HClO_2(aq)$ concentration, set up an ICE table for the equilibrium.	See the ICE table below.

	$HClO_2(aq)$ +	$H_2O(\ell)$	$\rightleftharpoons$	$ClO_2^-(aq)$ +	$H_3O^+(aq)$
	$[HClO_2]$ (mol/L)			$[ClO_2^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.075			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.075 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][ClO_2^-]}{[HClO_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][ClO_2^-]}{[HClO_2]}$ $1.15 \times 10^{-2} = \dfrac{(x)(x)}{0.075 - x}$

Re-arrange the expression into a quadratic equation.	$1.15 \times 10^{-2} = \dfrac{(x)(x)}{0.075 - x}$ $1.15 \times 10^{-2}(0.075 - x) = x^2$ $8.625 \times 10^{-4} - 1.15 \times 10^{-2}x = x^2$ $x^2 + 1.15 \times 10^{-2}x - 8.625 \times 10^{-4} = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-1.15\times10^{-2} \pm \sqrt{(1.15\times10^{-2})^2 - (4)(1)(-8.625\times10^{-4})}}{2(1)}$ $= \dfrac{-1.15\times10^{-2} \pm (5.985\times10^{-2})}{2}$ $x = 2.418 \times 10^{-2}$ or $x = -3.568 \times 10^{-2}$ Because it is impossible to have a negative concentration, x cannot equal -3.568×10^{-2}. Therefore, $[H_3O^+] = 2.418 \times 10^{-2}$ mol/L.
Calculate the pH using the equation pH = $-\log [H_3O^+]$.	pH = $-\log [H_3O^+]$ $= -\log (2.418 \times 10^{-2})$ $= 1.62$

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The answer seems reasonable for a weak acid having this K_a.

52. Cyanic acid, HCNO(aq), has a K_a value of 3.47×10^{-4}. It ionizes according to the equation below:

$$HCNO(aq) + H_2O(\ell) \rightleftharpoons CNO^-(aq) + H_3O^+(aq)$$
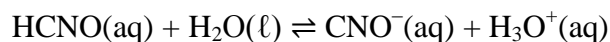
Determine the pH of a 0.015 mol/L acid sample.

What Is Required?
You must determine the pH of the solution.

What Is Given?
For HCNO(aq), you know the initial concentration is 0.015 mol/L.
You know HCNO(aq) has a K_a of 3.47×10^{-4}.
You have a balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of HCNO(aq) that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial HCNO(aq) concentration, set up an ICE table for the equilibrium.	See the ICE table below.

	HCNO(aq)	+	H₂O(ℓ)	⇌	CNO⁻(aq)	+	H₃O⁺(aq)
	[HCNO] (mol/L)				[CNO⁻] (mol/L)		[H₃O⁺] (mol/L)
I	0.015				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.015 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][CNO^-]}{[HCNO]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][CNO^-]}{[HCNO]}$ $3.47 \times 10^{-4} = \dfrac{(x)(x)}{0.015 - x}$
Re-arrange the expression into a quadratic equation.	$3.47 \times 10^{-4} = \dfrac{(x)(x)}{0.015 - x}$ $3.47 \times 10^{-4}(0.015 - x) = x^2$ $5.205 \times 10^{-6} - 3.47 \times 10^{-4}\,x = x^2$ $x^2 + 3.47 \times 10^{-4}x - 5.205 \times 10^{-6} = 0$

Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$\begin{aligned} x &= \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a} \\[6pt] &= \dfrac{-3.47 \times 10^{-4} \pm \sqrt{(3.47 \times 10^{-4})^2 - (4)(1)(-5.205 \times 10^{-6})}}{2(1)} \\[6pt] &= \dfrac{-3.47 \times 10^{-4} \pm (4.576 \times 10^{-3})}{2} \end{aligned}$ $x = 2.114 \times 10^{-3}$ or $x = -2.462 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -2.462×10^{-3}. Therefore, $[H_3O^+] = 2.114 \times 10^{-3}$ mol/L.
Calculate the pH using the equation pH $= -\log [H_3O^+]$.	$\begin{aligned} \text{pH} &= -\log [H_3O^+] \\ &= -\log (2.114 \times 10^{-3}) \\ &= 2.67 \end{aligned}$

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The answer seems reasonable for a weak acid having this K_a.

53. What is the change in pH of the cyanic acid solution from question 52, if the solution concentration triples from 0.015 to 0.045 mol/L?

What Is Required?
You must determine the pH of the solution when the initial concentration triples from 0.015 mol/L to 0.045 mol/L.

What Is Given?
For HCNO(aq), you know the initial concentration is 0.045 mol/L.
You know that HCNO(aq) has a K_a of 3.47×10^{-4}.
You have a balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of HCNO(aq) that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial HCNO(aq) concentration, set up an ICE table for the equilibrium.	See the ICE table below.

	HCNO(aq) +	$H_2O(\ell)$	$\rightleftharpoons$	CNO^-(aq) +	H_3O^+(aq)
	[HCNO] (mol/L)			[CNO$^-$] (mol/L)	[H$_3$O$^+$] (mol/L)
I	0.045			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.045 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][CNO^-]}{[HCNO]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][CNO^-]}{[HCNO]}$ $3.47 \times 10^{-4} = \dfrac{(x)(x)}{(0.045 - x)}$
Re-arrange the expression into a quadratic equation.	$3.47 \times 10^{-4} = \dfrac{(x)(x)}{(0.045 - x)}$ $3.47 \times 10^{-4}(0.045 - x) = x^2$ $1.5615 \times 10^{-5} - 3.47 \times 10^{-4}x = x^2$ $x^2 + 3.47 \times 10^{-4}x - 1.5615 \times 10^{-5} = 0$

| Use the quadratic formula to solve for x.

$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ | $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$= \dfrac{-3.47 \times 10^{-4} \pm \sqrt{(3.47 \times 10^{-4})^2 - (4)(1)(-1.5615 \times 10^{-5})}}{2(1)}$

$= \dfrac{-3.47 \times 10^{-4} \pm 7.91 \times 10^{-3}}{2}$
$x = 3.782 \times 10^{-3}$ or $x = -4.128 \times 10^{-3}$

Because it is impossible to have a negative concentration, x cannot equal $= -4.128 \times 10^{-3}$.
Therefore, $[H_3O^+] = 3.782 \times 10^{-3}$ mol/L. |
| Calculate the pH using the equation $pH = -\log [H_3O^+]$. | $pH = -\log [H_3O^+]$
$\quad = -\log (3.781 \times 10^{-3})$
$\quad = 2.42$

The pH changes by 0.25 units when the initial concentration of acid triples. |

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The change in pH is small for a large change in concentration which is reasonable since the pH scale is logarithmic.

54. Determine the pH of a 0.084 mol/L periodic acid solution, $HIO_4(aq)$. The K_a value for the reaction below is 2.29×10^{-2}.

$$HIO_4(aq) + H_2O(\ell) \rightleftharpoons IO_4^-(aq) + H_3O^+(aq)$$
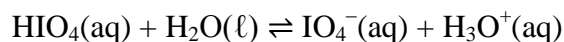
What Is Required?
You must determine the pH of the solution of periodic acid.

What Is Given?
For $HIO_4(aq)$, you know the initial concentration is 0.084 mol/L.
You know that $HIO_4(aq)$ has a K_a of 2.29×10^{-2}.
You have a balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of $HIO_4(aq)$ that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial HCNO(aq) concentration, set up an ICE table for the equilibrium.	See the ICE table below.

	$HIO_4(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$IO_4^-(aq)$	+	$H_3O^+(aq)$
	$[HIO_4]$ (mol/L)				$[IO_4^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.084				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.084 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][IO_4^-]}{[HIO_4]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][IO_4^-]}{[HIO_4]}$ $2.29 \times 10^{-2} = \dfrac{(x)(x)}{0.084 - x}$
Re-arrange the expression into a quadratic equation.	$2.29 \times 10^{-2} = \dfrac{(x)(x)}{0.084 - x}$ $2.29 \times 10^{-2}(0.084 - x) = x^2$ $1.9236 \times 10^{-3} - 2.29 \times 10^{-2}x = x^2$ $x^2 + 2.29 \times 10^{-2}x - 1.9236 \times 10^{-3} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-2.29 \times 10^{-2} \pm \sqrt{(2.29 \times 10^{-2})^2 - (4)(1)(-1.9236 \times 10^{-3})}}{2(1)}$$ $$= \frac{-2.29 \times 10^{-2} \pm 9.066 \times 10^{-2}}{2}$$ $x = 3.388 \times 10^{-2}$ or $x = -5.678 \times 10^{-2}$ Because it is impossible to have a negative concentration, x cannot equal -5.678×10^{-2}. Therefore, $[H_3O^+] = 3.388 \times 10^{-2}$ mol/L.
Calculate the pH using the equation pH $= -\log [H_3O^+]$.	pH $= -\log [H_3O^+]$ $= -\log (3.388 \times 10^{-2})$ $= 1.47$

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The answer seems reasonable for a weak acid having this K_a.

55. Butylamine is a monoprotic organic acid used in the manufacturing of pesticides. It has a K_a value of 2.51×10^{-11}. A chemist prepares a solution of the acid with a concentration of 0.064 mol/L. Determine the pH of the solution.

What Is Required?
You must determine the pH of the solution of butylamine.

What Is Given?
For butylamine, you know the initial concentration is 0.064 mol/L.
You know that butylamine has a K_a of 2.51×10^{-11}.

Plan Your Strategy	Act on Your Strategy
Write a balanced equation for the acid-dissociation using the symbol HBu for butylamine. Let x represent the change in concentration in mol/L of butylamine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial HBu(aq) concentration, set up an ICE table for the equilibrium.	See the ICE table below.

	HBu(aq)	+	H₂O(ℓ)	⇌	Bu⁻(aq)	+	H₃O⁺(aq)
	[HBu] (mol/L)				[Bu⁻] (mol/L)		[H₃O⁺] (mol/L)
I	0.064				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.064 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[\text{H}_3\text{O}^+][\text{Bu}^-]}{[\text{HBu}]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[\text{H}_3\text{O}^+][\text{Bu}^-]}{[\text{HBu}]}$ $2.51 \times 10^{-11} = \dfrac{(x)(x)}{0.064 - x}$
Re-arrange the expression into a quadratic equation.	$2.51 \times 10^{-11} = \dfrac{(x)(x)}{0.064 - x}$ $2.51 \times 10^{-11}(0.064 - x) = x^2$ $1.6064 \times 10^{-12} - 2.51 \times 10^{-11}x = x^2$ $x^2 + 2.51 \times 10^{-11}x - 1.6064 \times 10^{-12} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-2.51 \times 10^{-11} \pm \sqrt{(2.51 \times 10^{-11})^2 - (4)(1)(-1.6064 \times 10^{-12})}}{2(1)}$$ $$= \frac{-2.5110 \times 10^{-11} \pm 2.535 \times 10^{-6}}{2}$$ $$x = 1.267 \times 10^{-6} \text{ or } x = -1.268 \times 10^{-6}$$ Because it is impossible to have a negative concentration, x cannot equal -1.268×10^{-6}. Therefore, $[H_3O^+] = 1.267 \times 10^{-6}$ mol/L.
Calculate the pH using the equation pH $= -\log [H_3O^+]$.	pH $= -\log [H_3O^+]$ $= -\log (1.267 \times 10^{-6})$ $= 5.90$

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The answer seems reasonable for a weak acid having this K_a.

56. Triclopyr is a weak monoprotic organic acid used to control broadleaf weeds while leaving grasses unaffected. It has a K_a value of 2.09×10^{-3}. A solution of this acid is prepared with an acid concentration of 0.072 mol/L. What is the pH of this solution?

What Is Required?
You must determine the pH of the solution of triclopyr.

What Is Given?
For triclopyr, you know the initial concentration is 0.072 mol/L.
You know the triclopyr has a K_a of 2.09×10^{-3}.

Plan Your Strategy	Act on Your Strategy
Write a balanced equation for the acid-dissociation using the symbol HTr for triclopyr. Let x represent the change in concentration in mol/L of triclopyr that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial HTr(aq) concentration, set up an ICE table for the equilibrium.	See the ICE table below.

|	HTr(aq)	+	H$_2$O(ℓ)	$\rightleftharpoons$	Tr$^-$(aq)	+	H$_3$O$^+$(aq)
	[HTr] (mol/L)				[Tr$^-$] (mol/L)		[H$_3$O$^+$] (mol/L)
I	0.072				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.072 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[\text{H}_3\text{O}^+][\text{Tr}^-]}{[\text{HTr}]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[\text{H}_3\text{O}^+][\text{Tr}^-]}{[\text{HTr}]}$ $2.09 \times 10^{-3} = \dfrac{(x)(x)}{0.072 - x}$
Re-arrange the expression into a quadratic equation.	$2.09 \times 10^{-3} = \dfrac{(x)(x)}{0.072 - x}$ $2.09 \times 10^{-3}(0.072 - x) = x^2$ $1.5048 \times 10^{-4} - 2.09 \times 10^{-3}x = x^2$ $x^2 + 2.09 \times 10^{-3}x - 1.5048 \times 10^{-4} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$\frac{-2.09 \times 10^{-3} \pm \sqrt{(2.09 \times 10^{-3})^2 - (4)(1)(-1.5048 \times 10^{-4})}}{2(1)}$$ $$= \frac{-2.09 \times 10^{-3} \pm 2.462 \times 10^{-2}}{2}$$ $x = 1.126 \times 10^{-2}$ or $x = -1.336 \times 10^{-2}$ Because it is impossible to have a negative concentration, x cannot equal -1.336×10^{-2}. Therefore, $[H_3O^+] = 1.126 \times 10^{-2}$ mol/L.
Calculate the pH using the equation pH $= -\log [H_3O^+]$.	pH $= -\log [H_3O^+]$ $= -\log (1.126 \times 10^{-2})$ $= 1.95$

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The answer seems reasonable for a weak acid having this K_a.

57. Nitrobenzene is a monoprotic organic acid that has many uses, one of which is in the manufacturing of shoe and floor polishes. $K_a = 1.05 \times 10^{-4}$. The ionization of the acid is given by the following reaction:

$$C_6H_5NO_2(aq) + H_2O(\ell) \rightleftharpoons C_6H_4NO_2^-(aq) + H_3O+(aq)$$
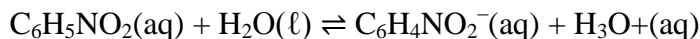
Determine the pH of a 0.10 mol/L solution of the acid.

What Is Required?
You must determine the pH of the solution of nitrobenzene.

What Is Given?
You know the initial concentration of nitrobenzene is 0.10 mol/L.
You know that nitrobenzene has a K_a of 1.05×10^{-4}.
You have the balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of nitrobenzene that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of nitrobenzene set up an ICE table for the equilibrium.	See the ICE table below.

	$C_6H_5NO_2(aq)$ +	$H_2O(\ell)$ $\rightleftharpoons$	$C_6H_4NO_2^-(aq)$ +	$H_3O+(aq)$
	$[C_6H_5NO_2]$ (mol/L)		$[C_6H_4NO_2^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.10		0.000	0.000
C	$-x$		$+x$	$+x$
E	$0.10 - x$		x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_6H_4NO_2^-]}{[C_6H_5NO_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_6H_4NO_2^-]}{[C_6H_5NO_2]}$ $1.05 \times 10^{-4} = \dfrac{(x)(x)}{0.10 - x}$

Re-arrange the expression into a quadratic equation.	$$1.05 \times 10^{-4} = \frac{(x)(x)}{0.10 - x}$$ $$1.05 \times 10^{-4}(0.10 - x) = x^2$$ $$1.05 \times 10^{-5} - 1.05 \times 10^{-4}x = x^2$$ $$x^2 + 1.05 \times 10^{-4}x - 1.05 \times 10^{-5} = 0$$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-1.05 \times 10^{-4} \pm \sqrt{(1.05 \times 10^{-4})^2 - (4)(1)(-1.05 \times 10^{-5})}}{2(1)}$$ $$= \frac{-1.05 \times 10^{-4} \pm 6.482 \times 10^{-3}}{2}$$ $x = 3.188 \times 10^{-3}$ or $x = -3.294 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot -3.294×10^{-3}. Therefore, $[H_3O^+] = 3.188 \times 10^{-3}$ mol/L.
Calculate the pH using the equation $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $= -\log (3.188 \times 10^{-3})$ $= 2.50$

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the initial concentration. The answer seems reasonable for a weak acid having this K_a.

58. Which solution is more acidic: a 0.045 mol/L solution of a weak monoprotic acid with $K_a = 7.74 \times 10^{-7}$ or a 0.087 mol/L solution of a weak monoprotic acid with $K_a = 9.81 \times 10^{-8}$?

What Is Required?
You must determine which one of two given acids is more acidic.

What Is Given?
For one acid, the initial concentration is 0.045 mol/L and the K_a is 7.74×10^{-7}.
For the second acid, the initial concentration is 0.087 mol/L and the K_a is 9.81×10^{-8}.

Plan Your Strategy	Act on Your Strategy
Designate the first acid as HA_1. Write a balanced equation for the acid-dissociation of the first acid. Let x represent the change in concentration in mol/L of HA_1 that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of HA_1, set up an ICE table for the equilibrium.	See the ICE table below.

	HA_1(aq)	+	H_2O(ℓ)	⇌	A_1^-(aq)	+	H_3O+(aq)
	$[HA_1]$ (mol/L)				$[A_1^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.045				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.045 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][A_1^-]}{[HA_1]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][A_1^-]}{[HA_1]}$ $7.74 \times 10^{-7} = \dfrac{(x)(x)}{0.045 - x}$
Re-arrange the expression into a quadratic equation.	$7.74 \times 10^{-7} = \dfrac{(x)(x)}{0.045 - x}$ $7.74 \times 10^{-7}(0.045 - x) = x^2$ $3.483 \times 10^{-8} - 7.74 \times 10^{-7}x = x^2$ $x^2 + 7.74 \times 10^{-7}x - 3.483 \times 10^{-8} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-7.74 \times 10^{-7} \pm \sqrt{(7.74 \times 10^{-7})^2 - (4)(1)(-3.4830 \times 10^{-8})}}{2(1)}$$ $$= \frac{-7.74 \times 10^{-7} \pm 3.733 \times 10^{-4}}{2}$$ $$x = 1.863 \times 10^{-4} \text{ or } x = -1.870 \times 10^{-4}$$ Because it is impossible to have a negative concentration, x cannot equal -1.870×10^{-4}. Therefore, $[H_3O^+] = 1.863 \times 10^{-4}$ mol/L.
Calculate the pH using the equation pH = $-\log [H_3O^+]$.	pH = $-\log [H_3O^+]$ $= -\log (1.863 \times 10^{-4})$ $= 3.73$
Designate the second acid as HA_2. Write a balanced equation for the acid-dissociation of the second acid. Let x represent the change in concentration in mol/L of HA_2 that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of HA_2, set up an ICE table for the equilibrium.	See the ICE table below.

	$HA_2(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$A_2^-(aq)$	+	$H_3O^+(aq)$
	$[HA_2]$ (mol/L)				$[A_2^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.087				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.087 - x$				x		x

Write the K_a expression for the reaction.	$$K_a = \frac{[H_3O^+][A_2^-]}{[HA_2]}$$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$$K_a = \frac{[H_3O^+][A_2^-]}{[HA_2]}$$ $$9.81 \times 10^{-8} = \frac{(x)(x)}{0.087 - x}$$

Re-arrange the expression into a quadratic equation.	$$9.81 \times 10^{-8} = \frac{(x)(x)}{0.087 - x}$$ $$9.81 \times 10^{-8}(0.087 - x) = x^2$$ $$8.535 \times 10^{-9} - 9.81 \times 10^{-8}x = x^2$$ $$x^2 + 9.81 \times 10^{-8}x - 8.535 \times 10^{-9} = 0$$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-9.81 \times 10^{-8} \pm \sqrt{(9.81 \times 10^{-8})^2 - (4)(1)(-8.535 \times 10^{-9})}}{2(1)}$$ $$= \frac{-9.81 \times 10^{-8} \pm 1.8477 \times 10^{-4}}{2}$$ $$x = 9.234 \times 10^{-5} \text{ or } x = -9.243 \times 10^{-5}$$ Because it is impossible to have a negative concentration, x cannot equal -9.243×10^{-5}. Therefore, $[H_3O^+] = 9.234 \times 10^{-5}$ mol/L.
Calculate the pH using the equation pH = $-\log$ $[H_3O^+]$.	pH = $-\log$ $[H_3O^+]$ $= -\log (9.234 \times 10^{-5})$ $= 4.03$
Identify the acid having the higher $[H_3O^+]$.	Acid HA_1 has a lower pH and a higher $[H_3O^+]$.

Check Your Solution
The answer for the concentration of $[H_3O^+]$ seems reasonable for these K_a values.

59. Which solution is more acidic: a 0.15 mol/L solution of a weak monoprotic acid with $K_a = 0.025$ or a 0.34 mol/L solution of a weak monoprotic acid with $K_a = 0.018$?

What Is Required?
You must determine which of two given acids is more acidic.

What Is Given?
For one acid, the initial concentration is 0.15 mol/L and the K_a is 0.025.
For the second acid, the initial concentration is 0.34 mol/L and the K_a is 0.018.

Plan Your Strategy	Act on Your Strategy
Designate the first acid as HA_1. Write a balanced equation for the acid-dissociation of HA_1. Let x represent the change in concentration in mol/L of HA_1 that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of HA_1, set up an ICE table for the equilibrium.	See the ICE table below.

	$HA_1(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$A_1^-(aq)$	+	$H_3O+(aq)$
	$[HA_1]$ (mol/L)				$[A_1^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.15				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.15 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][A_1^-]}{[HA_1]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][A_1^-]}{[HA_1]}$ $0.025 = \dfrac{(x)(x)}{0.15 - x}$
Re-arrange the expression into a quadratic equation.	$0.025 = \dfrac{(x)(x)}{0.15 - x}$ $(0.025)(0.15 - x) = x^2$ $3.75 \times 10^{-3} - 0.025x = x^2$ $x^2 + 0.025x - 3.75 \times 10^{-3} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-0.025 \pm \sqrt{(0.025)^2 - (4)(1)(-3.75 \times 10^{-3})}}{2(1)}$$ $$= \frac{-0.025 \pm 0.125}{2}$$ $$x = 5.00 \times 10^{-2} \text{ or } x = -0.0750$$ Because it is impossible to have a negative concentration, x cannot equal -0.0750. Therefore, $[H_3O^+] = 5.00 \times 10^{-2}$ mol/L.
Calculate the pH using the equation $pH = -\log [H_3O^+]$.	$$pH = -\log [H_3O^+]$$ $$= -\log (5.00 \times 10^{-2})$$ $$= 1.30$$
Designate the second acid as HA_2. Write a balanced equation for the acid-dissociation of HA_2. Let x represent the change in concentration in mol/L of HA_2 that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of HA_2, set up an ICE table for the equilibrium.	See the ICE table below.

	$HA_2(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$A_2^-(aq)$	+	$H_3O+(aq)$
	$[HA_2]$ (mol/L)				$[A_2^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.34				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.34 - x$				x		x

| Write the K_a expression for the reaction.	$$K_a = \frac{[H_3O^+][A_2^-]}{[HA_2]}$$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$$K_a = \frac{[H_3O^+][A_2^-]}{[HA_2]}$$ $$0.018 = \frac{(x)(x)}{0.34 - x}$$
Re-arrange the expression into a quadratic equation.	$$0.018 = \frac{(x)(x)}{0.34 - x}$$ $$(0.018)(0.34 - x) = x^2$$ $$6.12 \times 10^{-3} - 0.018x = x^2$$ $$x^2 + 0.018x - 6.12 \times 10^{-3} = 0$$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-0.018 \pm \sqrt{(0.018)^2 - (4)(1)(-6.12 \times 10^{-3})}}{2(1)}$$ $$= \frac{-0.018 \pm 0.1575}{2}$$ $x = 0.1395$ or $x = -0.0878$ Because it is impossible to have a negative concentration, x cannot equal -0.0878. Therefore, $[H_3O^+] = 0.1395$ mol/L.
Calculate the pH using the equation $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $= -\log (0.1395)$ $= 0.85$
Identify the acid having the higher $[H_3O^+]$.	Acid HA_2 has the lower pH and a higher $[H_3O^+]$.

Check Your Solution
The answers for the concentration of $[H_3O^+]$ seem reasonable for these K_a values.

Determining Concentrations from K_a and Initial Weak Acid Concentration
(Student textbook page 516)

60. Sulfamic acid, $NH_2SO_3H(aq)$, is the main ingredient in many household cleaners designed to remove calcium, lime, and rust stains. It has a K_a value of 0.089. What is the $[H_3O^+]$ of a 0.15 mol/L solution of the acid? The chemical equation is given below:

$$NH_2SO_3H(aq) + H_2O(\ell) \rightleftharpoons NH_2SO_3^-(aq) + H_3O^+(aq)$$
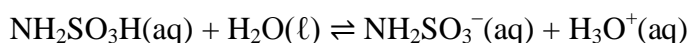
What Is Required?
You must determine the hydronium ion concentration in a given sulfamic acid solution.

What Is Given?
You know the initial concentration of sulfamic acid is 0.15 mol/L.
You know the sulfamic acid has a K_a of 0.089.
You have the balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of sulfamic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of sulfamic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$NH_2SO_3H(aq)$ +	$H_2O(\ell)$	$\rightleftharpoons$	$NH_2SO_3^-(aq)$ +	$H_3O^+(aq)$
	$[NH_2SO_3H]$ (mol/L)			$[NH_2SO_3^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.15			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.15 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][NH_2SO_3^-]}{[NH_2SO_3H]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][NH_2SO_3^-]}{[NH_2SO_3H]}$ $0.089 = \dfrac{(x)(x)}{0.15 - x}$

Re-arrange the expression into a quadratic equation.	$0.089 = \dfrac{(x)(x)}{0.15 - x}$ $0.089(0.15 - x) = x^2$ $1.335 \times 10^{-2} - 0.089x = x^2$ $x^2 + 0.089x - 1.335 \times 10^{-2} = 0$
Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-0.089 \pm \sqrt{(0.089)^2 - (4)(1)(-1.335 \times 10^{-2})}}{2(1)}$ $= \dfrac{-0.089 \pm 0.2476}{2}$ $x = 0.0793$ or $x = -0.1683$ Because it is impossible to have a negative concentration, x cannot equal -0.1683. Therefore, $[H_3O^+] = 7.9 \times 10^{-2}$ mol/L.

Check Your Solution
This hydronium ion concentration seems reasonable for a weak acid having this K_a.

61. Dichloroacetic acid, $C_2H_2Cl_2O_2(aq)$, is a weak organic monoprotic acid with $K_a = 0.045$. It is used in cancer research, where researchers at the University of Alberta used the acid on cancer cells in mice. They found that the acid restored the function of the mitochondria in the cancerous cells, allowing these cells to self-destruct and cause the tumour to shrink. What is the $[H_3O^+]$ of a 0.025 mol/L solution of the acid? The chemical equation is given below:
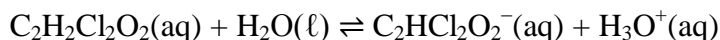
$$C_2H_2Cl_2O_2(aq) + H_2O(\ell) \rightleftharpoons C_2HCl_2O_2^-(aq) + H_3O^+(aq)$$

What Is Required?
You must determine the hydronium ion concentration in a given solution of dichloroacetic acid.

What Is Given?
You know the initial concentration of dichloroacetic acid is 0.025 mol/L.
You know the dichloroacetic acid has a K_a of 0.045.
You have the balanced equation for the acid-dissociation.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of dichloroacetic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of dichloroacetic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_2H_2Cl_2O_2(aq)$ +	$H_2O(\ell)$	$\rightleftharpoons$	$C_2HCl_2O_2^-(aq)$ +	$H_3O^+(aq)$
	$[C_2H_2Cl_2O_2]$ (mol/L)			$[C_2HCl_2O_2^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.025			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.025 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_2HCl_2O_2^-]}{[C_2H_2Cl_2O_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_2HCl_2O_2^-]}{[C_2H_2Cl_2O_2]}$ $0.045 = \dfrac{(x)(x)}{0.025 - x}$

Re-arrange the expression into a quadratic equation.	$$0.045 = \frac{(x)(x)}{0.025 - x}$$ $$(0.045)(0.025 - x) = x^2$$ $$1.125 \times 10^{-3} - 0.045x = x^2$$ $$x^2 + 0.045x - 1.125 \times 10^{-3} = 0$$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-0.045 \pm \sqrt{(0.045)^2 - (4)(1)(-1.125 \times 10^{-3})}}{2(1)}$$ $$= \frac{-0.045 \pm 0.08078}{2}$$ $$x = 1.789 \times 10^{-2} \text{ or } x = -1.258 \times 10^{-1}$$ Because it is impossible to have a negative concentration, x cannot equal -1.258×10^{-1}. Therefore, $[H_3O^+] = 1.8 \times 10^{-2}$ mol/L.

Check Your Solution
This hydronium ion concentration seems reasonable for a weak acid having this K_a. Working backward using the calculated $[H_3O^+]$ gives the K_a. The hydronium ion concentration is expressed to the correct number of significant digits.

62. Acetic acid, $C_2H_4O_2(aq)$, is the weak monoprotic acid found in vinegar. It has a K_a value of 1.75×10^{-5}. Determine the $[H_3O^+]$ in a 0.84 mol/L acetic acid solution.

What Is Required?

You must determine the hydronium ion concentration in a given solution of acetic acid.

What Is Given?

You know the initial concentration of acetic acid is 0.84 mol/L.
You know the acetic acid has a K_a of 1.75×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the acid-dissociation. Let x represent the change in concentration in mol/L of acetic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of acetic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_2H_4O_2(aq)$ +	$H_2O(\ell)$	$\rightleftharpoons$	$C_2H_3O_2^-(aq)$ +	$H_3O^+(aq)$
	$[C_2H_4O_2]$ (mol/L)			$[C_2H_3O_2^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.84			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.84 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_2H_3O_2^-]}{[C_2H_4O_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_2H_3O_2^-]}{[C_2H_4O_2]}$ $1.75 \times 10^{-5} = \dfrac{(x)(x)}{0.84 - x}$
Re-arrange the expression into a quadratic equation.	$1.75 \times 10^{-5} = \dfrac{(x)(x)}{0.84 - x}$ $1.75 \times 10^{-5}(0.84 - x) = x^2$ $1.47 \times 10^{-5} - 1.75 \times 10^{-5}x = x^2$ $x^2 + 1.75 \times 10^{-5}\,x - 1.47 \times 10^{-5} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-1.75 \times 10^{-5} \pm \sqrt{(1.75 \times 10^{-5})^2 - (4)(1)(-1.47 \times 10^{-5})}}{2(1)}$$ $$\frac{-1.75 \times 10^{-5} \pm 7.668 \times 10^{-3}}{2}$$ $x = 3.825 \times 10^{-3}$ or $x = -3.843 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -3.843×10^{-3}. Therefore, $[H_3O^+] = 3.8 \times 10^{-3}$ mol/L.

Check Your Solution
This hydronium ion concentration seems reasonable for a weak acid having this K_a. Working backward using the calculated $[H_3O^+]$ gives the K_a. The hydronium ion concentration is expressed to the correct number of significant digits.

63. What is the non-ionized periodic acid, $HIO_4(aq)$, concentration at equilibrium if the initial concentration of the weak acid is 0.10 mol/L? The K_a for periodic acid is 2.29×10^{-2} and the reaction is written as follows:

$$HIO_4(aq) + H_2O(\ell) \rightleftharpoons IO_4^-(aq) + H_3O^+(aq)$$
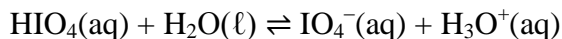
What Is Required?
You must determine the non-ionized concentration of periodic acid in a given solution.

What Is Given?
You know the initial concentration of periodic acid is 0.10 mol/L.
You know the periodic acid has a K_a of 2.29×10^{-2}.
You have the balanced equation for the reaction.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of periodic acid that ionize to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of periodic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$HIO_4(aq)$	$+$	$H_2O(\ell)$	$\rightleftharpoons$	$IO_4^-(aq)$	$+$	$H_3O^+(aq)$
	$[HIO_4]$ (mol/L)				$[IO_4^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.10				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.10 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][IO_4^-]}{[HIO_4]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][IO_4^-]}{[HIO_4]}$ $2.29 \times 10^{-2} = \dfrac{(x)(x)}{0.10 - x}$
Re-arrange the expression into a quadratic equation.	$2.29 \times 10^{-2} = \dfrac{(x)(x)}{0.10 - x}$ $2.29 \times 10^{-2} (0.10 - x) = x^2$ $2.29 \times 10^{-3} - 2.29 \times 10^{-2} x = x^2$ $x^2 + 2.29 \times 10^{-2} x - 2.29 \times 10^{-3} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-2.29 \times 10^{-2} \pm \sqrt{(2.29 \times 10^{-2})^2 - (4)(1)(-2.29 \times 10^{-3})}}{2(1)}$$ $$= \frac{-2.29 \times 10^{-2} \pm 9.841 \times 10^{-2}}{2}$$ $$x = 3.776 \times 10^{-2} \text{ or } x = -6.066 \times 10^{-2}$$ Because it is impossible to have a negative concentration, x cannot equal -6.066×10^{-2}. Therefore, $[H_3O^+] = 3.8 \times 10^{-2}$ mol/L.
Subtract the calculated value of $[H_3O^+]$ from the initial concentration of periodic acid.	[non-ionized periodic acid] $= 0.10$ mol/L $- 3.8 \times 10^{-2}$ mol/L $= 0.062$ mol/L

Check Your Solution

For a weak acid having this K_a this a reasonable amount of non-ionized acid. The answer shows the correct number of significant digits.

64. How does the concentration of the hydronium ion for acetic acid change in question 62, if the solution concentration increases by a factor of 2 to a value of 1.68 mol/L?

What Is Required?
You must determine the hydronium ion concentration in a given solution of acetic acid.

What Is Given?
You know the initial concentration of acetic acid is 1.68 mol/L.
You know the value of the K_a of the acid is 1.75×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the acid-dissociation. Let x represent the change in concentration mol/L of acetic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of acetic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_2H_4O_2$(aq)	+	H_2O(ℓ)	⇌	$C_2H_3O_2{}^-$(aq)	+	H_3O^+(aq)
	$[C_2H_4O_2]$ (mol/L)				$[C_2H_3O_2{}^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	1.68				0.000		0.000
C	$-x$				$+x$		$+x$
E	$1.68 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_2H_3O_2{}^-]}{[C_2H_4O_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_2H_3O_2{}^-]}{[C_2H_4O_2]}$ $1.75 \times 10^{-5} = \dfrac{(x)(x)}{1.68 - x}$
Re-arrange the expression into a quadratic equation.	$1.75 \times 10^{-5} = \dfrac{(x)(x)}{1.68 - x}$ $1.75 \times 10^{-5}(1.68 - x) = x^2$ $2.94 \times 10^{-5} - 1.75 \times 10^{-5}x = x^2$ $x^2 + 1.75 \times 10^{-5}\, x - 2.94 \times 10^{-5} = 0$

| Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ | $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-1.75 \times 10^{-5} \pm \sqrt{(1.75 \times 10^{-5})^2 - (4)(1)(-2.94 \times 10^{-5})}}{2(1)}$$ $$= \frac{-1.75 \times 10^{-5} \pm 1.084 \times 10^{-2}}{2}$$ $$x = 5.411 \times 10^{-3} \text{ or } x = -5.429 \times 10^{-3}$$ Because it is impossible to have a negative concentration, x cannot equal -5.429×10^{-3}. Therefore, $[H_3O^+] = 5.4 \times 10^{-3}$ mol/L. |
| Compare the concentration of the hydronium acid in question 62 to the calculated value of the concentration. | The hydronium ion concentration changes from 3.8×10^{-3} mol/L to 5.4×10^{-3} mol/L. |

Check Your Solution
The change in the hydronium ion concentraion is less than the concentration of the acid. This is expected for a weak acid.

65. Benzoic acid, $C_7H_6O_2$(aq), is a weak monoprotic organic acid used in food preservation because it inhibits the growth of mould, yeast, and some bacteria. It has a $K_a = 6.25 \times 10^{-5}$. If initially, a solution of the acid has a concentration of 0.074 mol/L, what will be the equilibrium concentration of the non-ionized benzoic acid? The chemical equation is given below:

$$C_7H_6O_2(aq) + H_2O(\ell) \rightleftharpoons C_7H_5O_2^-(aq) + H_3O^+(aq)$$
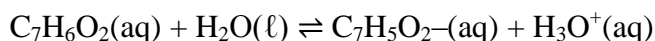
What Is Required?
You must determine the non-ionized concentration of benzoic acid in a given solution.

What Is Given?
You know the initial concentration of benzoic acid is 0.074 mol/L.
You know the benzoic acid has a K_a of 6.25×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of benzoic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of benzoic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_7H_6O_2$(aq) +	$H_2O(\ell)$	$\rightleftharpoons$	$C_7H_5O_2^-$(aq) +	H_3O^+(aq)
	$[C_7H_6O_2]$ (mol/L)			$[C_7H_5O_2^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.074			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.074 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_7H_5O_2^-]}{[C_7H_6O_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_7H_5O_2^-]}{[C_7H_6O_2]}$ $6.25 \times 10^{-5} = \dfrac{(x)(x)}{0.074 - x}$
Re-arrange the expression into a quadratic equation.	$6.25 \times 10^{-5} = \dfrac{(x)(x)}{0.074 - x}$ $6.25 \times 10^{-5} (0.074 - x) = x^2$ $4.625 \times 10^{-6} - 6.25 \times 10^{-5}x = x^2$ $x^2 + 6.25 \times 10^{-5}x - 4.625 \times 10^{-6} = 0$

Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-6.25\times10^{-5} \pm \sqrt{(6.25\times10^{-5})^2 - (4)(1)(-4.625\times10^{-6})}}{2(1)}$ $= \dfrac{-6.25\times10^{-5} \pm 4.302\times10^{-3}}{2}$ $x = 2.120 \times 10^{-3}$ or $x = -2.182 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -2.182×10^{-3}. Therefore, $[H_3O^+] = 2.12 \times 10^{-3}$ mol/L.
Subtract the calculated value of $[H_3O^+]$ from the initial concentration of benzoic acid.	[non-ionized benzoic acid] = 0.074 mol/L $- 2.12 \times 10^{-3}$ mol/L = 0.072 mol/L

Check Your Solution

For a weak acid having this K_a, this a reasonable amount of non-ionized acid. The answer shows the correct number of significant digits.

66. Heptanoic acid is used in the preparation of esters used in fragrances and as artificial flavours. It is a weak monoprotic acid with a K_a value of 1.29×10^{-5}. What is the hydronium ion concentration in a 0.12 mol/L solution of the acid? The chemical equation is given below:
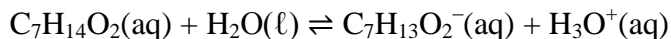
$$C_7H_{14}O_2(aq) + H_2O(\ell) \rightleftharpoons C_7H_{13}O_2^-(aq) + H_3O^+(aq)$$

What Is Required?
You must determine the hydronium ion concentration in a given solution of heptanoic acid.

What Is Given?
You know the initial concentration of heptanoic acid is 0.12 mol/L.
You know the acid has a K_a of 1.29×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of heptanoic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced chemical equation and the initial concentration of heptanoic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_7H_{14}O_2(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$C_7H_{13}O_2^-(aq)$	+	$H_3O^+(aq)$
	$[C_7H_{14}O_2]$ (mol/L)				$[C_7H_{13}O_2^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.12				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.12 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_7H_{13}O_2^-]}{[C_7H_{14}O_2]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_7H_{13}O_2^-]}{[C_7H_{14}O_2]}$ $1.29 \times 10^{-5} = \dfrac{(x)(x)}{0.12 - x}$
Re-arrange the expression into a quadratic equation.	$1.29 \times 10^{-5} = \dfrac{(x)(x)}{0.12 - x}$ $1.29 \times 10^{-5}(0.12 - x) = x^2$ $1.548 \times 10^{-6} - 1.29 \times 10^{-5}x = x^2$ $x^2 + 1.29 \times 10^{-5}x - 1.548 \times 10^{-6} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-1.29\times10^{-5} \pm \sqrt{(1.29\times10^{-5}) - (4)(1)(1.548\times10^{-6})}}{(2)(1)}$$ $$= \frac{-1.29\times10^{-5} \pm 2.488\times10^{-3}}{2}$$ $x = 1.238\times10^{-3}$ or $x = -1.250 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -1.250×10^{-3}. Therefore, $[H_3O^+] = 1.2 \times 10^{-3}$ mol/L.

Check Your Solution
This is a reasonable hydronium concentration for a weak acid having the given Ka. The correct number if significant digits are shown.

67. Mandelic acid has long been used by doctors as an antibacterial medication, primarily for the treatment of urinary tract infections. It has a molecular formula of C_8H_8O and a K_a value of 4.27×10^{-5}. Initially, 1.84 g of the solid is used to make a 1.0 L solution of the weak acid. The chemical equation is given below:

$$C_6H_5CH(OH)CO_2H \text{ (aq)} + H_2O(\ell) \rightleftharpoons C_6H_5CH(OH)CO_2^-(aq) + H_3O^+(aq)$$
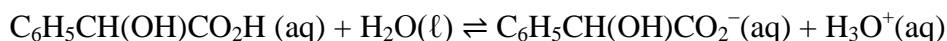
Determine the $[H_3O^+]$ in this solution.

What Is Required?

You must determine the hydronium ion concentration in a given solution of mandelic acid.

What Is Given?

You know the mass of mandelic acid is 1.84 g.
You know the volume of the solution is 1.0 L.
You know the solution has a K_a of 4.27×10^{-5}.

| Plan Your Strategy	Act on Your Strategy
Use the mass and molar mass, M, of C_8H_8O(aq) to determine the amount in moles, n, of the acid using the formula $n = \dfrac{m}{M}$.	$n_{C_8H_8O} = \dfrac{m}{M}$ $= \dfrac{1.84 \text{ g}}{120.16 \text{ g/mol}}$ $= 1.53 \times 10^{-2} \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to determine the initial concentration, c, of the mandelic acid.	$c_{C_8H_8O} = \dfrac{n}{V}$ $= \dfrac{1.53 \times 10^{-2} \text{ mol}}{1.0 \text{ L}}$ $= 1.53 \times 10^{-2} \text{ mol/L}$
Let x represent the change in concentration in mol/L of mandelic acid that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of mandelic acid, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_6H_5CH(OH)CO_2H$ (aq) +	$H_2O(\ell)$ $\rightleftharpoons$	$C_6H_5CH(OH)CO_2^-$(aq) +	H_3O^+(aq)
	$[C_6H_5CH(OH)CO_2H]$ (mol/L)		$[C_6H_5CH(OH)CO_2^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	1.53×10^{-2}		0.000	0.000
C	$-x$		$+x$	$+x$
E	$1.53 \times 10^{-2} - x$		x	x

Write the K_a expression for the reaction.	$$K_a = \frac{[H_3O^+][C_6H_5CH(OH)CO_2^-]}{[C_6H_5CH(OH)CO_2H]}$$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$$K_a = \frac{[H_3O^+][C_6H_5CH(OH)CO_2^-]}{[C_6H_5CH(OH)CO_2H]}$$ $$4.27 \times 10^{-5} = \frac{(x)(x)}{1.53 \times 10^{-2} - x}$$
Re-arrange the expression into a quadratic equation.	$$4.27 \times 10^{-5} = \frac{(x)(x)}{1.53 \times 10^{-2} - x}$$ $$4.27 \times 10^{-5}(1.53 \times 10^{-2} - x) = x^2$$ $$6.533 \times 10^{-7} - 4.27 \times 10^{-5}x = x^2$$ $$x^2 + 4.27 \times 10^{-5} x - 6.53 \times 10^{-7} = 0$$
Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-4.27 \times 10^{-5} \pm \sqrt{(4.27 \times 10^{-5})^2 - (4)(1)(-6.53 \times 10^{-7})}}{2(1)}$$ $$= \frac{-4.27 \times 10^{-5} \pm 1.617 \times 10^{-3}}{2}$$ $$x = 7.872 \times 10^{-4} \text{ or } x = -8.298 \times 10^{-4}$$ Because it is impossible to have a negative concentration, x cannot equal -8.298×10^{-4}. Therefore, $[H_3O^+] = 7.9 \times 10^{-4}$ mol/L.

Check Your Solution
This is a reasonable hydronium concentration for a weak acid having the given K_a. The correct number if significant digits are shown.

68. Tropine, $C_8H_{15}NO(aq)$, is a weak organic acid with a K_a value of 1.58×10^{-4}. It is used in the manufacturing of atropine, a chemical used by as a premedication for anaesthesia in some patients. Tropine ionizes according to the equation below:

$$C_8H_{15}NO(aq) + H_2O(\ell) \rightleftharpoons C_8H_{14}NO^-(aq) + H_3O^+(aq)$$
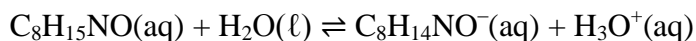
What Is the non-ionized tropine concentration at equilibrium in a 0.025 mol/L solution of the acid?

What Is Required?
You need to determine the non-ionized concentration of tropine in a given solution.

What Is Given?
You know the initial concentration of tropine is 0.025 mol/L.
You know that tropine has a K_a of 1.58×10^{-4}.

Plan Your Strategy	Act on Your Strategy
Let x represent the change in concentration in mol/L of tropine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of tropine, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_8H_{15}NO(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$C_8H_{14}NO^-(aq)$	+	$H_3O^+(aq)$
	$[C_8H_{15}NO]$ (mol/L)				$[C_8H_{14}NO^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.025				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.025 - x$				x		x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][C_8H_{14}NO^-]}{[C_8H_{15}NO]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][C_8H_{14}NO^-]}{[C_8H_{15}NO]}$ $1.58 \times 10^{-4} = \dfrac{(x)(x)}{0.025 - x}$
Re-arrange the expression into a quadratic equation.	$1.58 \times 10^{-4} = \dfrac{(x)(x)}{0.025 - x}$ $1.58 \times 10^{-4}(0.025 - x) = x^2$ $3.95 \times 10^{-6} - 1.58 \times 10^{-4}x = x^2$ $x^2 + 1.58 \times 10^{-4}\, x - 3.950 \times 10^{-6} = 0$

Use the quadratic formula to solve for x. $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$	$x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ $= \dfrac{-1.58 \times 10^{-4} \pm \sqrt{(1.58 \times 10^{-4})^2 - (4)(1)(-3.95 \times 10^{-6})}}{2(1)}$ $= \dfrac{-1.58 \times 10^{-4} \pm 3.978 \times 10^{-3}}{2}$ $x = 1.910 \times 10^{-3}$ or $x = -2.068 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -2.068×10^{-3}. Therefore, $[H_3O^+] = 1.9 \times 10^{-3}$ mol/L.
Subtract the calculated value of $[H_3O^+]$ from the initial concentration of tropine.	[non-ionized tropine] = 0.025 mol/L $-$ 1.92 $\times$ 10^{-3} mol/L = 0.023 mol/L

Check Your Solution
For a weak acid having this K_a , this a reasonable amount of non-ionized acid. The answer shows the correct number of significant digits.

69. Which solution has a higher concentration of hydronium ion: a 0.042 mol/L solution of a weak monoprotic acid with $K_a = 1.58 \times 10^{-4}$ or a 0.035 mol/L solution of a weak monoprotic acid with $K_a = 2.27 \times 10^{-4}$?

What Is Required?
You must determine which of two acids has a higher hydronium ion concentration.

What Is Given?
You know the initial concentration of the first acid is 0.042 mol/L.
You know the first acid has a K_a of 1.58×10^{-4}.
You know the second acid has an initial concentration of 0.035 mol/L.
You know the second acid has a K_a of 2.27×10^{-4}.

Plan Your Strategy	Act on Your Strategy
Write a balanced acid-dissociation equation for HA_1. Let x represent the change in concentration in mol/L of HA that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of HA_1, set up an ICE table for the equilibrium.	See the ICE table below.

	HA_1 +	$H_2O(\ell)$	$\rightleftharpoons$	A_1^- (aq) +	H_3O^+(aq)
	$[HA_1]$ (mol/L)			$[A_1^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.042			0.000	0.000
C	$-x$			$+x$	$+x$
E	$0.042 - x$			x	x

Write the K_a expression for the reaction.	$K_a = \dfrac{[H_3O^+][A_1^-]}{[HA_1]}$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$K_a = \dfrac{[H_3O^+][A_1^-]}{[HA_1]}$ $1.58 \times 10^{-4} = \dfrac{(x)(x)}{0.042 - x}$
Re-arrange the expression into a quadratic equation.	$1.58 \times 10^{-4} = \dfrac{(x)(x)}{0.042 - x}$ $1.58 \times 10^{-4}(0.042 - x) = x^2$ $6.636 \times 10^{-6} - 1.58 \times 10^{-4}x = x^2$ $x^2 + 1.58 \times 10^{-4}\,x - 6.636 \times 10^{-6} = 0$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-1.58 \times 10^{-4} \pm \sqrt{(1.58 \times 10^{-4})^2 - (4)(1)(6.636 \times 10^{-6})}}{(2)(1)}$$ $$= \frac{-1.58 \times 10^{-4} \pm 5.154 \times 10^{-3}}{2}$$ $x = 2.498 \times 10^{-3}$ or $x = -2.656 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -2.656×10^{-3}. Therefore, $[H_3O^+] = 2.5 \times 10^{-3}$ mol/L.

Plan Your Strategy	Act on Your Strategy
Write a balanced acid-dissociation equation for HA_2. Let x represent the concentration in mol/L of HA_2 that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of HA_2, set up an ICE table for the equilibrium.	See the ICE table below.

	HA_2	$+$	$H_2O(\ell)$	$\rightleftharpoons$	A_2^- (aq)	$+$	H_3O^+(aq)
	$[HA_2]$ (mol/L)				$[A_1^-]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.035				0.000		0.000
C	$-x$				$+x$		$+x$
E	$0.035 - x$				x		x

Write the K_a expression for the reaction.	$$K_a = \frac{[H_3O^+][A_2^-]}{[HA_2]}$$
Substitute the equilibrium concentrations and the value of K_a into the expression for K_a.	$$K_a = \frac{[H_3O^+][A_2^-]}{[HA_2]}$$ $$2.27 \times 10^{-4} = \frac{(x)(x)}{0.035 - x}$$
Re-arrange the expression into a quadratic equation.	$$2.27 \times 10^{-4} = \frac{(x)(x)}{0.035 - x}$$ $$2.27 \times 10^{-4}\,(0.035 - x) = x^2$$ $$7.945 \times 10^{-6} - 2.27 \times 10^{-4}x = x^2$$ $$x^2 + 2.27 \times 10^{-4}x - 7.945 \times 10^{-6} = 0$$

Use the quadratic formula to solve for x. $$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$	$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$ $$= \frac{-2.27 \times 10^{-4} \pm \sqrt{(2.27 \times 10^{-4})^2 - (4)(1)(7.94 \times 10^{-6})}}{(2)(1)}$$ $$= \frac{-2.27 \times 10^{-4} \pm 5.640 \times 10^{-3}}{2}$$ $x = 2.706 \times 10^{-3}$ or $x = -2.934 \times 10^{-3}$ Because it is impossible to have a negative concentration, x cannot equal -2.934×10^{-3}. Therefore, $[H_3O^+] = 2.7 \times 10^{-3}$ mol/L.
Determine which acid has the higher concentraion of H_3O^+(aq).	The hydronium ion concentration is higher in the second acid.

Check Your Solution
The hydronium concentratins have the correct number of significant digits. the concentrations are very nearly the same which would be expected since their concentrations and K_a values are very close.

Calculating the pH of a Weak Base
(Student textbook page 523)

70. Write the chemical equation for each base ionizing in an aqueous solution:
   a. ammonia, $NH_3(aq)$
   b. trimethylamine, $(CH_3)_3N(aq)$
   c. hydrogen sulfite ion, $HSO_3^-(aq)$
   d. carbonate ion, $CO_3^{2-}(aq)$

What Is Required?
You have to write the balanced chemical equation for the ionization of a base.

What Is Given?
The name and chemical formula of the base is given.

Plan Your Strategy	Act on Your Strategy
A base is proton-acceptor. Donate a proton, H^+, from H_2O to the base.	**a.** $NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ **b.** $(CH_3)_3N(aq) + H_2O(\ell) \rightleftharpoons (CH_3)_3NH^+(aq) + OH^-(aq)$ **c.** $HSO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_2SO_3(aq) + OH^-(aq)$ **d.** $CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

Check Your Solution
In each case $OH^-(aq)$ is produced as expected from a hydrolysis reaction with a base. There is conservation of charge in each case.

71. Write the K_b expression for each base in question 70.

What Is Required?
You have to write the expression for the base-ionization constant.

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

Plan Your Strategy	Act on Your Strategy
Follow the pattern for the writing K_b for the generic reaction. $$B(aq) + H_2O(\ell) \rightleftharpoons BH^+(aq) + OH^-(aq)$$ $$K_b = \frac{[BH^+][OH^-]}{[B]}$$	**a.** $K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$ **b.** $K_b = \dfrac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$ **c.** $K_b = \dfrac{[H_2SO_3][OH^-]}{[HSO_3^-]}$ **d.** $K_b = \dfrac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$

Check Your Solution
Each expression follows the generic example.

72. When a 0.25 mol/L aqueous solution of methylamine was prepared, the pH of the solution was found to be 10.04. What percentage of methylamine molecules ionized in the solution?

What Is Required?
You must determine the percentage of methylamine molecules that have ionized in the solution.

What Is Given?
You know the initial concentration of methylamine is 0.25 mol/L.
You know the methylamine has a pH of 10.04.

Plan Your Strategy	Act on Your Strategy
Use the expression pH + pOH = 14.00 to determine the pOH of the methylamine solution.	$\begin{aligned} \text{pOH} &= 14.00 - \text{pH} \\ &= 14.00 - 10.04 \\ &= 3.96 \end{aligned}$
Calculate [OH⁻] using the expression $[\text{OH}^-] = 10^{-\text{pOH}}$.	$\begin{aligned} [\text{OH}^-] &= 10^{-\text{pOH}} \\ &= 10^{-3.96} \\ &= 1.1 \times 10^{-4} \text{ mol/L} \end{aligned}$
Express [OH⁻] as a percentage of the original concentration, c, of methylamine.	$\begin{aligned} \text{percentage } c_{\text{methylamine ionized}} &= \frac{[\text{OH}^-]}{[\text{methylamine}]} \times 100\% \\ &= \frac{1.1 \times 10^{-4} \text{ mol/L}}{0.25 \text{ mol/L}} \times 100\% \\ &= 0.044\% \end{aligned}$

Check Your Solution
The answer has the correct number of significant digits. This low a percentage ionization is reasonable for a weak base.

73. Codeine, $C_{18}H_{21}NO_3(s)$, is added to some cough medicines. When a 0.020 mol/L aqueous solution of codeine was prepared, the pH of the solution was found to be 10.26. Calculate K_b for codeine.

What Is Required?
You need to calculate the K_b for a given solution of the base codeine.

What Is Given?
You know the initial concentration of codeine is 0.020 mol/L.
You know the pH of the solution is 10.26.

Plan Your Strategy	Act on Your Strategy
Use the expression pH + pOH = 14.00 to determine the pOH of the codeine solution.	pH + pOH = 14.00 pOH = 14.00 − 10.26 = 3.74
Use the expression $[OH^-] = 10^{-pOH}$ to calculate $[OH^-]$.	$[OH^-] = 10^{-pOH}$ $= 10^{-3.74}$ $= 1.82 \times 10^{-4}$ mol/L (carry 3 sig. digits)
Write the balanced chemical equation for the base-ionization of codeine. Let x represent the change in concentration in mol/L of codeine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of codeine, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_{18}H_{21}NO_3(aq)$ +	$H_2O(\ell)$ ⇌	$C_{18}H_{22}NO_3^+(aq)$ +	$OH^-(aq)$
	$[C_{18}H_{21}NO_3]$ (mol/L)		$[C_{18}H_{22}NO_3^+]$ (mol/L)	$[OH^-]$ (mol/L)
I	0.020		0	0
C	-1.82×10^{-4}		$+1.82 \times 10^{-4}$	$+ 1.82 \times 10^{-4}$
E	$0.020 - 1.82 \times 10^{-4}$		$+1.82 \times 10^{-4}$	$+ 1.82 \times 10^{-4}$

| Write the K_b expression for the reacton and calculate the K_b. | $K_b = \dfrac{[C_{18}H_{22}NO_3^+][OH^-]}{[C_{18}H_{21}NO_3]}$

 $= \dfrac{(1.82 \times 10^{-4})^2}{1.98 \times 10^{-2}}$

 $= 1.7 \times 10^{-6}$ |

Check Your Solution
The answer shows the correct number of significant digits and is a reasonable answer for a base of this pH.

74. A Material Safety Data Sheet (MSDS) describes pyridine, $C_5H_5N(\ell)$, as a clear liquid with a putrid odour. A 16 g/L solution of pyridine has pH = 9.23. Use these data to calculate K_b for pyridine.

What Is Required?
You need to calculate the K_b for a given solution of the base pyridine.

What Is Given?
You know that 16 g of pyridine is dissolved in 1.0 L of solution.
You know the pH of the solution is 9.23.

Plan Your Strategy	Act on Your Strategy
Use the mass and molar mass, M, of $C_5H_5N(\ell)$ to determine the amount in moles, n, of pyridine using the formula $n = \dfrac{m}{M}$.	$n_{C_5H_5N} = \dfrac{m}{M}$ $= \dfrac{16 \text{ g}}{79.11 \text{ g/mol}}$ $= 0.202 \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to determine the initial concentration, c, of pyridine.	$c_{C_5H_5N} = \dfrac{n}{V}$ $= \dfrac{0.202 \text{ mol}}{1.0 \text{ L}}$ $= 0.202 \text{ mol/L}$
Use the expression pH + pOH = 14.00 to determine the pOH of the pyridine solution.	pOH = 14.00 − 9.23 = 4.77
Use the expression $[OH^-] = 10^{-pOH}$ to calculate $[OH^-]$.	$[OH^-] = 10^{-pOH}$ $= 10^{-4.77}$ $= 1.7 \times 10^{-5} \text{ mol/L}$
Write the balanced equation for the base-ionization of pyridine. Let x represent the change in concentration in mol/L of pyridine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the calculated initial concentration of pyridine, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_5H_5N(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$C_5H_6N^+(aq)$	+	$OH^-(aq)$
	$[C_5H_5N]$ (mol/L)				$[C_5H_6N^+]$ (mol/L)		$[OH^-]$ (mol/L)
I	0.202				0		0
C	-1.7×10^{-5}				$+1.7 \times 10^{-5}$		$+1.7 \times 10^{-5}$
E	$0.202 - 1.7 \times 10^{-5}$				$+1.7 \times 10^{-5}$		$+1.7 \times 10^{-5}$

Write the K_b expression for the reaction and calculate the K_b.	$K_b = \dfrac{[C_5H_6N^+][OH^-]}{[C_5H_5N]}$ $= \dfrac{(1.7 \times 10^{-5})^2}{0.202}$ $= 1.4 \times 10^{-9}$

Check Your Solution
The answer shows the same number of significant digits as is found to the right of the decimal place in the given pH. The answer is reasonable for a base of this pH.

75. An aqueous solution of household ammonia, $NH_3(aq)$, has a molar concentration of 0.105 mol/L. Calculate the pH of the solution. K_b for ammonia is 1.8×10^{-5}.

What Is Required?
You need to calculate the pH for a given solution of the ammonia.

What Is Given?
You know the initial concentration of ammonia is 0.105 mol/L.
You know the K_b for ammonia is 1.8×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the base-ionization of ammonia. Let x represent the change in concentration in mol/L of ammonia that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of ammonia, set up an ICE table for the equilibrium.	See the ICE table below.

	$NH_3(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$NH_4^+(aq)$	+	$OH^-(aq)$
	$[NH_3]$ (mol/L)				$[NH_4^+]$ (mol/L)		$[OH^-]$ (mol/L)
I	0.105				0		0
C	$-x$				$+x$		$+x$
E	$0.105 - x$				$+x$		$+x$

Write the K_b expression for the reaction.	$K_b = \dfrac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_3\right]}$
Use the K_b expression to determine the $[OH^-]$.	$K_b = \dfrac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_3\right]}$ $= \dfrac{(x)(x)}{0.105 - x}$ $\approx \dfrac{(x)(x)}{0.105}$ $1.8 \times 10^{-5} = \dfrac{(x)(x)}{0.105}$ $x^2 = 1.89 \times 10^{-6}$ $x = 1.37 \times 10^{-3}$ $[OH^-] = 1.37 \times 10^{-3}$

Calculate pOH using the expression $pOH = -\log [OH^-]$.	$\begin{aligned} pOH &= -\log [OH^-] \\ &= -\log (1.37 \times 10^{-3}) \\ &= 2.86 \end{aligned}$
determine pH using the relationship $pH + pOH = 14.00$.	$\begin{aligned} pH + pOH &= 14.00 \\ pH &= 14.00 - 2.86 \\ &= 11.14 \end{aligned}$

Check Your Solution

The answer shows the correct number of significant digits to the right of the decimal as is given in the K_b. This is a reasonable pH for an ammonia solution of the given concentration.

76. Hydrazine, $N_2H_4(\ell)$, has been used as a rocket fuel. The concentration of an aqueous solution of hydrazine is 5.9×10^{-2} mol/L. Calculate the pH of the solution. K_b for hydrazine is 1.3×10^{-6}.

What Is Required?

You need to calculate the pH for a given solution of the hydrazine.

What Is Given?

You know the initial concentration of hydrazine is 5.9×10^{-2} mol/L.
You know the K_b for hydrazine is 1.3×10^{-6}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the base-ionization of hydrazine. Let x represent the change in concentration in mol/L of hydrazine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of hydrazine, set up an ICE table for the equilibrium.	See the ICE table below.

	$N_2H_4(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$N_2H_5^+(aq)$	+	$OH^-(aq)$
	$[N_2H_4]$ (mol/L)				$[N_2H_5^+]$ (mol/L)		$[OH^-]$ (mol/L)
I	5.9×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$0.\,5.9 \times 10^{-2} - x$				$+x$		$+x$

Write the K_b expression for the reaction.	$K_b = \dfrac{[N_2H_5^+][OH^-]}{[N_2H_4]}$
Use the K_b expression to determine the $[OH^-]$.	$\begin{aligned} K_b &= \dfrac{[N_2H_5^+][OH^-]}{[N_2H_4]} \\ &= \dfrac{(x)(x)}{5.9 \times 10^{-2} - x} \\ &\approx \dfrac{x^2}{5.9 \times 10^{-2}} \end{aligned}$ $\begin{aligned} 1.3 \times 10^{-6} &= \dfrac{(x)(x)}{5.9 \times 10^{-2}} \\ x^2 &= 7.67 \times 10^{-8} \\ x &= 2.769 \times 10^{-4} \end{aligned}$ $[OH^-] = 2.769 \times 10^{-4}$

Calculate pOH using the expression $pOH = -\log [OH^-]$.	$pOH = -\log [OH^-]$ $\quad = -\log (2.769 \times 10^{-4})$ $\quad = 3.56$
determine pH using the relationship $pH + pOH = 14.00$.	$pH + pOH = 14.00$ $\quad pH = 14.00 - 3.56$ $\quad\quad = 10.44$

Check Your Solution

The answer shows the correct number of significant digits to the right of the decimal as is given in the initial concentration of hydrazine. This is a reasonable pH for a hydrazine solution of the given concentration.

77. Morphine, $C_{17}H_{19}NO_3(s)$, is a naturally occurring base found in opium poppy plants that is used to control pain. A 4.5×10^{-3} mol/L solution has a pH of 9.93. Calculate K_b for morphine.

What Is Required?
You need to calculate the K_b for a given solution of the base morphine.

What Is Given?
You know the initial concentration of morphine is 4.5×10^{-3} mol/L.
You know the pH of the solution is 9.93.

Plan Your Strategy	Act on Your Strategy
Use the expression pH + pOH = 14.00 to determine the pOH of the morphine solution.	pH + pOH = 14.00 pOH = 14.00 − 9.93 = 4.07
Use the expression $[OH^-] = 10^{-pOH}$ to calculate $[OH^-]$.	$[OH^-] = 10^{-pOH}$ $= 10^{-4.07}$ $= 8.51 \times 10^{-5}$ mol/L (carry 3 sig. digits)
Write the balanced equation for the base-ionization of morphine. Let x represent the change in concentration in mol/L of codeine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of morphine, set up an ICE table for the equilibrium.	See the ICE table below.

	$C_{17}H_{19}NO_3$ (aq) +	$H_2O(\ell)$	⇌	$C_{17}H_{20}NO_3^+$(aq) +	OH^-(aq)
	$[C_{17}H_{19}NO_3]$ (mol/L)			$[C_{17}H_{20}NO_3^+]$ (mol/L)	$[OH^-]$ (mol/L)
I	4.5×10^{-3}			0	0
C	-8.51×10^{-5}			$+8.51 \times 10^{-5}$	$+8.51 \times 10^{-5}$
E	$4.5 \times 10^{-3} - 8.51 \times 10^{-5}$			$+8.51 \times 10^{-5}$	$+8.51 \times 10^{-5}$

Write the K_b expression for the reaction and calculate the K_b.	$K_b = \dfrac{[C_{17}H_{20}NO_3^+][OH^-]}{[C_{17}H_{19}NO_3]}$ $= \dfrac{(8.51\times10^{-5})^2}{4.41\times10^{-3}}$ $= 1.6 \times 10^{-6}$

Check Your Solution
The answer shows the correct number of significant digits and is a reasonable answer for a base of this pH.

78. Methylamine, $CH_3NH_2(g)$, is a fishy-smelling gas at room temperature. It is used in the manufacture of several prescription drugs, including methamphetamine. Calculate $[OH^-]$ and pOH of a 1.5 mol/L aqueous solution of methylamine. K_b for methylamine is 4.6×10^{-4}.

What Is Required?
You need to calculate the $[OH^-]$ and pOH for a given solution of the methylamine.

What Is Given?
You know the initial concentration of methylamine is 1.5 mol/L.
You know the K_b for methylamine is 4.6×10^{-4}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the base-ionization of methylamine. Let x represent the change in concentration in mol/L of methylamine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of methylamine, set up an ICE table for the equilibrium.	See the ICE table below.

	$CH_3NH_2(aq)$	$+$	$H_2O(\ell)$	$\rightleftharpoons$	$CH_3NH_3^+(aq)$	$+$	$OH^-(aq)$
	$[CH_3NH_2]$ (mol/L)				$[CH_3NH_3^+]$ (mol/L)		$[OH^-]$ (mol/L)
I	1.5				0		0
C	$-x$				$+x$		$+x$
E	$1.5 - x$				$+x$		$+x$

Write the K_b expression for the reaction.	$K_b \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$
Use the K_b and the information in the ICE table to determine the $[OH^-]$.	$K_b = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$ $= \dfrac{(x)(x)}{1.5-x}$ $\approx \dfrac{(x)(x)}{1.5}$ $4.6 \times 10^{-4} = \dfrac{x^2}{1.5}$ $x^2 = 6.9 \times 10^{-4}$ $x = 2.6 \times 10^{-2}$ $[OH^-] = 2.6 \times 10^{-2}$

Calculate pOH using the expression $pOH = -\log [OH^-]$.	$pOH = -\log [OH^-]$ $= -\log (2.6 \times 10^{-2})$ $= 1.58$

Check Your Solution
The answer shows the correct number of significant digits to the right of the decimal as is given in the initial concentration of methylamine. This is a reasonable pH for a methylamine solution of the given concentration.

79. At room temperature, trimethylamine, $(CH_3)_3N(g)$, is a gas with a strong ammonia-like odour. Calculate $[OH^-]$ and the percentage of trimethylamine molecules that react with water in a 0.22 mol/L aqueous solution. K_b for trimethylamine is 6.3×10^{-5}.

What Is Required?
You need to calculate the $[OH^-]$ and the percentage of trimethylamine molecules that ionize.

What Is Given?
You know the initial concentration of hydrazine is 0.22 mol/L.
You know the K_b for hydrazine is 6.3×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the base-ionization of trimethylamine. Let x represent the change in concentration in mol/L of trimethylamine that ionizes to reach equilibrium conditions. Using the stoichiometric relationships in the balanced equation and the initial concentration of trimethylamine, set up an ICE table for the equilibrium.	See the ICE table below.

	$(CH_3)_3N(aq)$	+	$H_2O(\ell)$	$\rightleftharpoons$	$(CH_3)_3NH^+(aq)$	+	$OH^-(aq)$
	$[(CH_3)_3N]$				$[(CH_3)_3NH^+]$		$[OH^-]$
I	0.22				0		0
C	$-x$				$+x$		$+x$
E	$0.22 - x$				$+x$		$+x$

Write the K_b expression for the reaction.	$K_b = \dfrac{[(CH_3)_3 NH^+][OH^-]}{[(CH_3)_3 N]}$

Use the K_b expression to determine the [OH⁻].	$$K_b = \frac{[(CH_3)_3 NH^+][OH^-]}{[(CH_3)_3 N]}$$ $$= \frac{(x)(x)}{0.22 - x}$$ $$\approx \frac{x^2}{0.22}$$ $$6.3 \times 10^{-5} = \frac{x^2}{0.22}$$ $$x^2 = 1.386 \times 10^{-5}$$ $$x = 3.7 \times 10^{-3}$$ $$[OH^-] = 3.7 \times 10^{-3}$$
Calculate the percentage of trimethylamine molecules that have ionized.	percentage of trimethylamine that ionizes $$= \frac{3.7 \times 10^{-3}}{0.22} \times 100\%$$ $$= 1.7\%$$

Check Your Solution

The answer shows the correct number of significant digits. The percentage of trimethylamine that ionizes is reasonable for a weak base.

80. An aqueous solution of ammonia has a pH of 10.85. What is the concentration of the solution?

What Is Required?
You need to calculate the concentration of an ammonia solution at equilibrium.

What Is Given?
You know the ammonia solution has a pH of 10.85.
From Appendix B, you know the K_b for ammonia is 1.8×10^{-5}.

Plan Your Strategy	Act on Your Strategy
Use the expression pH + pOH = 14.00 to determine the pOH of the morphine solution.	pH + pOH = 14.00 pOH = 14.00 − 10.85 = 3.15
Use the expression $[OH^-] = 10^{-pOH}$ to calculate $[OH^-]$.	$[OH^-] = 10^{-pOH}$ $= 10^{-3.15}$ $= 7.08 \times 10^{-4}$ (carry 3 sig. digits)
Write the balanced equation for the base-ionization of ammonia. Let x represent the initial concentration of ammonia at equilibrium. Using x and the stoichiometric relationships in the balanced equation , set up an ICE table for the equilibrium.	See the ICE table below.

	$NH_3(aq)$	+	$H_2O(\ell)$	⇌	$NH_4^+(aq)$	+	$OH^-(aq)$
	$[NH_3]$ (mol/L)				$[NH_4^+]$ (mol/L)		$[OH^-]$ (mol/L)
I	?				0		0
C	7.08×10^{-4}				$+7.08 \times 10^{-4}$		$+7.08 \times 10^{-4}$
E	x				7.08×10^{-4}		7.08×10^{-4}

Write the K_b expression for the reaction and use this information to determine the $[x]$.	$K_b = \dfrac{\left[NH_4^+\right]\left[OH^-\right]}{\left[NH_3\right]}$ $\dfrac{(7.08\times10^{-4})^2}{x} = 1.8\times10^{-5}$ $x = 0.028$ mol/L
The initial concentration of ammonia is equal to the equilibrium concentration + the $[OH^-]$.	$c_{initial} = 0.028$ mol/L $+ 7.08 \times 10^{-4}$ mol/L $= 0.029$ mol/L

Check Your Solution
Working backward with concentration of ammonia gives a solution of the given pH. The answer seems reasonable.

Equilibrium Constants for Conjugate Acid-Base Pairs
(Student textbook page 526)

81. List the conjugate bases for each acid below:
   a. periodic acid, HIO_4(aq)
   b. cyanic acid, $HCNO$(aq)
   c. nitrous acid, HNO_2(aq)
   d. hydrofluoric acid, HF(aq)

What Is Required?
You need to write the formula for the conjugate base.

What Is Given?
You know the name and formula for the acid.

Plan Your Strategy	Act on Your Strategy
The acid acts as a proton donor for its conjugate base. Take a proton, H^+, from the acid and what remains is the conjugate base.	**a.** IO_4^-(aq) **b.** CNO^-(aq) **c.** NO_2^-(aq) **d.** F^-(aq)

Check Your Solution
Each base has one less H^+ than its conjugate acid.

82. Write the K_b expressions for each conjugate base in question 81.

What Is Required?
You need to write the expression for the K_b expressions for the given bases.

What Is Given?
You know the formula for the base.

Plan Your Strategy	Act on Your Strategy
Write the equation for the reaction between each base with water. Write the expression for the equilibrium constant, K_b, for each reaction.	**a.** $IO_4{}^-(aq) + H_2O(\ell) \rightleftharpoons HIO_4(aq) + OH^-(aq)$ $K_b = \dfrac{[HIO_4][OH^-]}{[IO_4{}^-]}$ **b.** $CNO^-(aq) + H_2O(\ell) \rightleftharpoons HCNO(aq) + OH^-(aq)$ $K_b = \dfrac{[HCNO][OH^-]}{[CNO^-]}$ **c.** $NO_2{}^-(aq) + H_2O(\ell) \rightleftharpoons HNO_2(aq) + OH^-(aq)$ $K_b = \dfrac{[HNO_2][OH^-]}{[NO_2{}^-]}$ **d.** $F^-(aq) + H_2O(\ell) \rightleftharpoons HF(aq) + OH^-(aq)$ $K_b = \dfrac{[HF][OH^-]}{[F^-]}$

Check Your Solution
Each base is properly shown accepting a proton from water. The K_b expressions each properly show products in the numerator and reactants in the denominator.

83. Calculate the K_b value for the conjugate base of hydrosulfuric acid.

What Is Required?
You have to calculate the ionization constant for the conjugate base of hydrosulfuric acid.

What Is Given?
You have the name of the acid, hydrosulfuric acid.
From Appendix B, the K_a for this acid is 8.9×10^{-8}.
The ion-product constant of water , K_w, is 1.0×10^{-14}.

Plan Your Strategy	Act on Your Strategy
Calculate K_b using the expression $K_w = K_a\,K_b = 1.0 \times 10^{-14}$.	$\begin{aligned} K_b &= \dfrac{K_w}{K_a} \\ &= \dfrac{1.0 \times 10^{-14}}{8.9 \times 10^{-8}} \\ &= 1.1 \times 10^{-7} \end{aligned}$

Check Your Solution
The answer has the same number of significant digits as the given data and seems reasonable for the K_b.

84. Calculate the K_b value for the conjugate base of the hydrogen citrate ion.

What Is Required?
You have to calculate the ionization constant for the conjugate base of hydrogen citrate ion.

What Is Given?
You have the name of the acid, hydrogen citrate.
From Appendix B, the K_a for this acid is 4.0×10^{-7}.
The ion-product constant of water constant of water, K_w, is 1.0×10^{-14}.

Plan Your Strategy	Act on Your Strategy
Calculate K_b using the expression $K_w = K_a\,K_b = 1.0 \times 10^{-14}$.	$\begin{aligned} K_b &= \dfrac{K_w}{K_a} \\ &= \dfrac{1.0 \times 10^{-14}}{4.0 \times 10^{-7}} \\ &= 2.5 \times 10^{-8} \end{aligned}$

Check Your Solution
The answer has the same number of significant digits as the given data and seems reasonable for the K_b.

85. Calculate the K_b value for the conjugate base of acetic acid.

What Is Required?
You have to calculate the ionization constant for the conjugate base of acetic acid.

What Is Given?
You have the name of the acid, acetic acid.
From Appendix B, the K_a for this acid is 1.8×10^{-5}.
The ion-product constant of water constant of water, K_w, is 1.0×10^{-14}.

Plan Your Strategy	Act on Your Strategy
Calculate K_b using the expression $K_w = K_a K_b = 1.0 \times 10^{-14}$.	$K_b = \dfrac{K_w}{K_a}$ $= \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$ $= 5.6 \times 10^{-8}$

Check Your Solution
The answer has the same number of significant digits as the given data and seems reasonable for the K_b.

86. Calculate the K_a value for the conjugate acid of dimethylamine.

What Is Required?
You have to calculate the ionization constant for the conjugate acid of dimethylamine.

What Is Given?
You have the name of the base, dimethylamine.
From Appendix B, the K_b for this base is 5.4×10^{-4}.
The ion-product constant of water constant of water, K_w, is 1.0×10^{-14}.

Plan Your Strategy	Act on Your Strategy
Calculate K_a using the expression $K_w = K_a K_b = 1.0 \times 10^{-14}$.	$K_a = \dfrac{K_w}{K_b}$ $= \dfrac{1.0 \times 10^{-14}}{5.4 \times 10^{-4}}$ $= 1.9 \times 10^{-11}$

Check Your Solution
The answer has the same number of significant digits as the given data and seems reasonable for the K_a of a weak acid.

87. Calculate the K_a value for the conjugate acid of hydrazine.

What Is Required?
You have to calculate the ionization constant for the conjugate acid of hydrazine.

What Is Given?
You have the name of the base, hydrazine.
From Appendix B, the K_b for this base is 1.3×10^{-6}.
The ion-product constant of water constant of water, K_w, is 1.0×10^{-14}.

Plan Your Strategy	Act on Your Strategy
Calculate K_a using the expression $K_w = K_a K_b = 1.0 \times 10^{-14}$.	$K_a = \dfrac{K_w}{K_b}$ $= \dfrac{1.0 \times 10^{-14}}{1.3 \times 10^{-6}}$ $= 7.7 \times 10^{-9}$

Check Your Solution
The answer has the same number of significant digits as the given data and seems reasonable for the K_a of a weak acid.

88. Calculate the K_a value for the conjugate acid of hydroxylamine.

What Is Required?
You have to calculate the ionization constant for the conjugate acid of hydroxylamine.

What Is Given?
You have the name of the base, hyrdroxylamine.
From Appendix B, the K_b for this base is 8.8×10^{-9}.
The ion-product constant of water constant of water, K_w, is 1.0×10^{-14}.

Plan Your Strategy	Act on Your Strategy
Calculate K_a using the expression $K_w = K_a K_b = 1.0 \times 10^{-14}$.	$K_a = \dfrac{K_w}{K_b}$ $= \dfrac{1.0 \times 10^{-14}}{8.8 \times 10^{-9}}$ $= 1.1 \times 10^{-6}$

Check Your Solution
The answer has the same number of significant digits as the given data and seems reasonable for the K_a of a weak acid.

89. "The K_b expression for a conjugate base is simply the reciprocal expression of the K_a expression for the acid of that conjugate base." Is this statement true or false? Explain your reasoning.

Answer
This statement is false. The reactions of the acid and conjugate base either losing or gaining a proton are the reverse processes of each other, but not in terms of their equations for K_a and K_b, as these both involve the combination of the acid or conjugate base with water. In the K_a situation, the hydronium ion is produced and in the K_b situation, the hydroxide ion is produced.

90. Verify your answer to question 89 by writing the K_a and K_b expressions, using ammonia as the example.

What Is Required?
You need to illustrate that the expression for the conjugate base is not the reciprocal of the expression for the acid of that conjugate base.

What Is Given?
You have ammonia as an example.

Plan Your Strategy	Act on Your Strategy
Write the equation for ammonia acting as a base and as an acid. Write the respective expressions for K_a and K_b.	$NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ $K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$ $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $K_a = \dfrac{[NH_3][H_3O^+]}{[NH_4^+]}$

Check Your Solution
The expressions for K_a and K_b are not the reciprocal of each other.

Calculating pH During a Strong Acid-Strong Base Titration
(Student textbook page 535)

91. What is the pH of 35.00 mL of a 0.0225 mol/L hydrochloric acid solution, HCl(aq)?

What Is Required?
You have to calculate the pH of a hydrochloric acid solution.

What Is Given?
You know the volume of HCl(aq) is 35.00 mL and its concentration is 0.0225 mol/L

Plan Your Strategy	Act on Your Strategy
Since HCl(aq) is a strong acid, $[H_3O^+] = [HCl]$ Use the equation $pH = -\log [H_3O^+]$.	$[H_3O^+] = [HCl]$ $\qquad\quad = 0.0225$ mol/L $pH = -\log [H_3O^+]$ $\qquad = -\log 0.0225$ $\qquad = 1.648$ (3 sig. digits)

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the given concentration. The answer is reasonable.

92. During a titration using sodium hydroxide, NaOH(aq), and hydroiodic acid, HI(aq), 35.00 mL of 0.15 mol/L sodium hydroxide was added to 14.40 mL of 0.12 mol/L hydroiodic acid. At this point in the titration, what is the pH of the solution?

What Is Required?
You need to find the pH at a certain point in an acid-base titration.

What Is Given?
You know that 35.00 mL of NaOH(aq) having a concentration of 0.15 mol /L has been added to 14.40 mL of HI(aq) having a concentration of 0.12 mol/L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of HI(aq) to NaOH(aq).	$HI(aq) + NaOH(aq) \rightarrow NaI(aq) + H_2O(\ell)$ The balanced equation for the reaction gives a mole ratio $\dfrac{[HI(aq)]}{[NaOH(aq)]} = \dfrac{1}{1}$.
Use the formula $n = cV$ to calculate the amount in moles, n, of NaOH(aq) and HI(aq).	$\begin{aligned} n_{NaOH} &= cV \\ &= (0.15 \text{ mol}/\cancel{L})(0.03500\,\cancel{L}) \\ &= 0.00525 \text{ mol} \end{aligned}$ $\begin{aligned} n_{HI} &= cV \\ &= (0.12 \text{ mol}/\cancel{L})(0.01440\,\cancel{L}) \\ &= 0.00173 \text{ mol} \end{aligned}$
Determine which reagent is in excess. If HI(aq) is in excess, $[HI] = [H_3O^+]$. If NaOH(aq) is in excess, $[NaOH] = [OH^-]$.	$n_{NaOH} = 0.00525 \text{ mol}$ $n_{HI} = 0.00173 \text{ mol}$ Since $n_{NaOH} > n_{HI}$, NaOH(aq) is the reagent in excess.
Determine the amount in moles, n, of the excess of NaOH(aq).	$\begin{aligned} n &= 0.00525 \text{ mol} - 0.00173 \text{ mol} \\ &= 0.00352 \text{ mol} \end{aligned}$
Determine the the total volume, V, of the mixture.	$\begin{aligned} V &= 35.00 \text{ mL} + 14.40 \text{ mL} \\ &= 49.40 \text{ mL} \\ &= 0.0494 \text{ L} \end{aligned}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration, c, of the excess reagent.	$\begin{aligned} c_{NaOH} &= \dfrac{n}{V} \\ &= \dfrac{0.003\,52 \text{ mol}}{0.0494 \text{ L}} \\ &= 0.0712 \text{ mol/L} \end{aligned}$

Since NaOH(aq) is in excess, [NaOH] = [OH⁻]. Use the equations pOH = –log [OH⁻] and pH = 14.00 – pOH to determine the pH.	$[OH^-] = 0.0712$ mol/L $pOH = - \log [OH^-]$ $\qquad = - \log 0.0712$ $\qquad = 1.15$ $pH = 14.00 - pOH$ $\qquad = 14.00 - 1.15$ $\qquad = 12.85$

Check Your Solution

The resulting pH is greater than 7 indicating an excess of base (NaOH). The number of significant digits to the right of the decimal in the pH is the same as the number of significant digits in the given data. The answer seems reasonable.

93. At the point in a titration where 5.00 mL of 0.225 mol/L hydrochloric acid, HCl(aq), is added to 40.00 mL of 0.175 mol/L potassium hydroxide, KOH(aq), what is the pH of the solution?

What Is Required?
You need to find the pH at a certain point in an acid-base titration.

What Is Given?
You know that 5.00 mL of HCl(aq) having a concentration of 0.225 mol/L has been added to 40.00 mL of KOH(aq) having a concentration of 0.175 mol /L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of HCl(aq) to KOH(aq).	$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(\ell)$ The balanced equation for the reaction gives a mole ratio $\dfrac{[HCl(aq)]}{[KOH(aq)]} = \dfrac{1}{1}$.
Use the formula $n = cV$ to calculate the amount in moles, n, of HCl(aq) and KOH(aq).	$\begin{aligned} n_{KOH} &= cV \\ &= (0.175 \text{ mol}/\cancel{L})(0.0400\,\cancel{L}) \\ &= 0.007000 \text{ mol} \end{aligned}$ $\begin{aligned} n_{HCl} &= cV \\ &= (0.225 \text{ mol}/\cancel{L})(0.00500\,\cancel{L}) \\ &= 0.001125 \text{ mol} \end{aligned}$
Determine which reagent is in excess. If HCl(aq) is in excess, $[HCl] = [H_3O^+]$. If KOH(aq) is in excess, $[KOH] = [OH^-]$.	$n_{KOH} = 0.007000 \text{ mol}$ $n_{HCl} = 0.01125 \text{ mol}$ Since $n_{KOH} > n_{HCl}$, KOH(aq) is the reagent in excess.
Determine the amount in moles, n, of the excess of KOH(aq).	$\begin{aligned} n &= 0.007000 \text{ mol} - 0.001125 \text{ mol} \\ &= 0.005875 \text{ mol} \end{aligned}$
Determine the the total volume, V, of the mixture.	$\begin{aligned} V &= 5.00 \text{ mL} + 40.00 \text{ mL} \\ &= 45.00 \text{ mL} \\ &= 0.04500 \text{ L} \end{aligned}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration, c, of the excess reagent.	$\begin{aligned} c_{KOH} &= \dfrac{n}{V} \\ &= \dfrac{0.005875 \text{ mol}}{0.04500 \text{ L}} \\ &= 0.1305 \text{ mol/L} \end{aligned}$

| Since KOH(aq) is in excess, [KOH] = [OH$^-$]. Use the equations pOH = $-$log [OH$^-$] and pH = 14.00 $-$ pOH to determine the pH. | pOH = $-$log [OH$^-$]
$\qquad$ = $-$log 0.1305
$\qquad$ = 0.884

pH = 14.00 $-$ pOH
$\qquad$ = 14.00 $-$ 0.884
$\qquad$ = 13.116 (3 sig. digits) |

Check Your Solution
The resulting pH is greater than 7 indicating an excess of base (KOH). The number of significant digits to the right of the decimal in the pH is the same as the number of significant digits in the given data. The answer seems reasonable.

94. A beaker containing 31.50 mL of a 0.015 mol/L hydrobromic acid solution, HBr(aq), has been titrated with 24.53 mL of 0.012 mol/L potassium hydroxide, KOH(aq). Determine the pH of the solution at this point in the titration.

What Is Required?
You need to find the pH at a certain point in an acid-base titration.

What Is Given?
You know that 24.53 mL of KOH(aq) having a concentration of 0.012 mol /L has been added to 31.50 mL of HBr(aq) having a concentration of 0.015 mol/L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of HBr(aq) to KOH(aq).	HBr(aq) + KOH(aq) → KBr(aq) + H₂O(ℓ) The balanced chemical equation for the reaction gives a mole ratio $\dfrac{[\text{HBr(aq)}]}{[\text{KOH(aq)}]} = \dfrac{1}{1}$.
Use the formula $n = cV$ to calculate the amount in moles, n, of HBr(aq)and KOH(aq).	$n_{\text{HBr}} = cV$ $\quad = (0.015 \text{ mol}/\cancel{L})(0.03150 \cancel{L})$ $\quad = 4.725 \times 10^{-4} \text{ mol}$ $n_{\text{KOH}} = cV$ $\quad = (0.012 \text{ mol}/\cancel{L})(0.02453 \cancel{L})$ $\quad = 2.9436 \times 10^{-4} \text{ mol}$
Determine which reagent is in excess. If HBr(aq) is in excess, [HBr] = [H₃O⁺]. If KOH(aq) is in excess, [KOH] = [OH⁻].	$n_{\text{HBr}} = 4.725 \times 10^{-4} \text{ mol}$ $n_{\text{KOH}} = 2.9436 \times 10^{-4} \text{ mol}$ Since $n_{\text{HBr}} > n_{\text{KOH}}$, HBr(aq) is the reagent in excess.
Determine the amount in moles, n, of the excess of HBr(aq).	$n = 4.725 \times 10^{-4} \text{ mol} - 2.9436 \times 10^{-4} \text{ mol}$ $\quad = 1.7814 \times 10^{-4} \text{ mol}$
Determine the the total volume, V, of the mixture.	$V = 31.50 \text{ mL} + 24.53 \text{ mL}$ $\quad = 56.03 \text{ mL}$ $\quad = 0.05603 \text{ L}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration, c, of the excess reagent.	$c_{\text{HBr}} = \dfrac{n}{V}$ $\quad = \dfrac{1.7814 \times 10^{-4} \text{ mol}}{0.05603 \text{ L}}$ $\quad = 0.00318 \text{ mol/L}$

Since HBr(aq) is in excess, [HBr] = $[H_3O^+]$. Calculate pH using pH = $-\log [H_3O^+]$.	pH = $-\log [H_3O^+]$
	    = $-\log 0.00318$
	    = 2.50 (2 sig. digits)

Check Your Solution

The resulting pH is less than 7 indicating an excess of acid (HBr). The number of significant digits to the right of the decimal in the pH is the same as the number of significant digits in the given data. The answer seems reasonable.

95. Determine the pH of a solution during a titration when 25.00 mL of 0.74 mol/L nitric acid, $HNO_3(aq)$, has been added to 26.05 mL of 0.71 mol/L sodium hydroxide, $NaOH(aq)$.

What Is Required?
You need to find the pH at a certain point in an acid-base titration.

What Is Given?
You know that 25.00 mL of $HNO_3(aq)$ having a concentration of 0.74 mol/L has been added to 26.05 mL of $NaOH(aq)$ having a concentration of 0.71 mol /L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of $HNO_3(aq)$ to $NaOH(aq)$.	$HNO_3(aq) + NaOH(aq) \rightarrow$ $$NaNO_3(aq) + H_2O(\ell)$$ The balanced chemical equation for the reaction gives a mole ratio $\dfrac{[HNO_3(aq)]}{[NaOH(aq)]} = \dfrac{1}{1}$.
Use the formula $n = cV$ to calculate the amount in moles, n, of $HNO_3(aq)$ and $NaOH(aq)$.	$n_{HNO_3} = cV$ $\quad = (0.74 \text{ mol/} \cancel{L})(0.02500 \cancel{L})$ $\quad = 0.01850 \text{ mol}$ $n_{NaOH} = cV$ $\quad = (0.71 \text{ mol/} \cancel{L})(0.02605 \cancel{L})$ $\quad = 0.0184955 \text{ mol}$
Determine which reagent is in excess. If $HNO_3(aq)$ is in excess, $[HNO_3] = [H_3O^+]$. If $NaOH(aq)$ is in excess, $[NaOH] = [OH^-]$.	$n_{HNO_3} = 0.01850 \text{ mol}$ $n_{NaOH} = 0.0184955 \text{ mol}$ Since $n_{HNO_3} > n_{NaOH}$, $HBr(aq)$ is the reagent in excess.
Determine the amount in moles, n, of the excess of $HNO_3(aq)$.	$n = 0.1850 \text{ mol} - 0.184955 \text{ mol}$ $\quad = 4.500 \times 10^{-6} \text{ mol}$
Determine the the total volume, V, of the mixture.	$V = 25.00 \text{ mL} + 26.05 \text{ mL}$ $\quad = 51.05 \text{ mL}$ $\quad = 0.05105 \text{ L}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration, c, of the excess reagent.	$c_{HNO_3} = \dfrac{n}{V}$ $\quad = \dfrac{4.500 \times 10^{-6} \text{ mol}}{0.05105 \text{ L}}$ $\quad = 8.81 \times 10^{-5} \text{ mol/L}$

Since HNO_3(aq) is in excess, $[HNO_3] = [H_3O^+]$. Calculate pH using pH = $-\log [H_3O^+]$.	pH = $-\log [H_3O^+]$ $= -\log (8.81 \times 10^{-5})$ $= 4.05$ (2 sig. digits)

Check Your Solution
The resulting pH is less than 7 indicating an excess of acid (HNO_3). The number of significant digits to the right of the decimal in the pH is the same as the number of significant digits in the given data. The answer seems reasonable.

96. What are the pOH and pH of 25.00 mL of a 0.075 mol/L potassium hydroxide solution, KOH(aq)?

What Is Required?
You have to calculate the pH and pOH of a potassium hydroxide solution.

What Is Given?
You know the volume of KOH(aq) is 25.00 mL.
You know the concentration of KOH(aq)is 0.075 mol/L.

Plan Your Strategy	Act on Your Strategy
Since KOH(aq) is a strong base, $[KOH] = [OH^-]$.	$[OH^-] = [KOH]$ $\quad = 0.075$ mol/L
Use the equation $pOH = -\log [OH^-]$ to calculate pOH.	$pOH = -\log [OH^-]$ $\quad = -\log 0.075$ $\quad = 1.12$ (2 sig. digits)
Use the expression $pH + pOH = 14.00$ to determine pH.	$pH + pOH = 14.00$ $\quad\quad pH = 14.00 - 1.12$ $\quad\quad\quad = 12.88$ (2 sig. digits)

Check Your Solution
The number of digits to the right of the decimal in the pH is the same as the number of significant digits in the given concentration. The answers are reasonable.

97. What volume of a 0.22 mol/L hydrochloric acid solution, HCl(aq) is needed to cause the pH of 25.00 mL of a 0.15 mol/L lithium hydroxide, LiOH(aq), solution to have a pH that is equal to 7.00?

What Is Required?
You need to calculate the volume of HCl(aq) required to bring a solution of LiOH(aq) to a pH that is equal to 7.

What Is Given?
You know the HCl(aq) has a concentration of 0.22 mol/L.
You know the LiOH(aq) has a volume of 25.00 mL and a concentration of 0.15 mol/L.

Plan Your Strategy	Act on Your Strategy
For the pH of the solution to equal 7, the HCl(aq) must exactly neutralize the LiOH(aq). Write the balanced chemical equation for the reaction and determine the mole ratio of HCl(aq) to LiOH(aq).	$HCl(aq) + LiOH(aq) \rightarrow LiCl(aq) + H_2O(\ell)$ The balanced chemical equation for the reaction gives a mole ratio $\dfrac{[HCl(aq)]}{[LiOH(aq)]} = \dfrac{1}{1}$.
In a neutral solution $n_{H_3O^+} = n_{OH^-}$ Since LiOH(aq) is a strong base, $[LiOH] = [OH^-]$. Since HCl(aq) is a strong acid, $[HCl] = [H_3O^+]$.	$n_{H_3O^+} = n_{OH^-}$ $n_{HCl} = n_{LiOH}$
Determine the volume of HCl(aq) using the formula $n = cV$ and the relationship $n_{HCl} = n_{LiOH}$.	$n_{HCl} = n_{LiOH}$ $c_{HCl}V_{HCl} = c_{LiOH}V_{LiOH}$ $(0.22 \text{ mol/L})V_{HCl} = (0.15 \text{ mol/L})(0.02500 \text{ L})$ $V_{HCl} = \dfrac{(0.15 \ \cancel{\text{mol/L}})(0.02500 \text{ L})}{(0.22 \ \cancel{\text{mol/L}})}$ $= 0.01705 \text{ L}$ $= 17.05 \text{ mL}$ $= 17 \text{ mL}$

Check Your Solution
Since the concentration of HCl(aq) is greater than the concentration of LiOH(aq), it is reasonable that a smaller volume of HCl(aq) than LiOH(aq) is required.

98. What is the volume of 2.45 mol/L hydrochloric acid, HCl(aq), needed to completely neutralize 125.00 mL of 1.14 mol/L lithium hydroxide, LiOH(aq)?

What Is Required?
You need to calculate the volume of HCl(aq) required to neutralize a solution of LiOH(aq).

What Is Given?
You know the HCl(aq) has a concentration of 2.45 mol/L.
You know the LiOH(aq) has a volume of 125.00 mL and a concentration of 1.14 mol/L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of HCl(aq) to LiOH(aq).	$HCl(aq) + LiOH(aq) \rightarrow LiCl(aq) + H_2O(\ell)$ The balanced chemical equation for the reaction gives a mole ratio $\dfrac{[HCl(aq)]}{[LiOH(aq)]} = \dfrac{1}{1}$.
In a neutral solution $n_{H_3O^+} = n_{OH^-}$. Since LiOH(aq) is a strong base, $[LiOH] = [OH^-]$. Since HCl(aq) is a strong acid, $[HCl] = [H_3O^+]$.	$n_{H_3O^+} = n_{OH^-}$ $n_{HCl} = n_{LiOH}$
Determine the volume of HCl(aq) using the formula $n = cV$ and the relationship $n_{HCl} = n_{LiOH}$.	$n_{HCl} = n_{LiOH}$ $c_{HCl}V_{HCl} = c_{LiOH}V_{LiOH}$ $(2.45 \text{ mol/L})V_{HCl} = (1.14 \text{ mol/L})(0.12500 \text{ L})$ $V_{HCl} = \dfrac{(1.14 \ \cancel{\text{mol/L}})(0.12500 \text{ L})}{(0.22 \ \cancel{\text{mol/L}})}$ $= 0.05816 \text{ L}$ $= 58.16 \text{ mL}$ $= 58.2 \text{ mL}$

Check Your Solution
Since the concentration of HCl(aq) is greater than the concentration of LiOH(aq) , it is reasonable that a smaller volume of HCl(aq) than LiOH(aq) is required.

99. At the equivalence point, 18.76 mL of 1.75 mol/L potassium hydroxide, KOH(aq), has been titrated with 28.65 mL of hydrochloric acid. What is the concentration of the original hydrochloric acid solution?

What Is Required?
You must find the concentration of hydrochloric acid used in a neutralization reaction.

What Is Given?
At the equivalence point, the volume of hydrochloric acid used is 28.65 mL.
You know the KOH(aq) used has a volume of 18.76 mL and a concentration of 1.75 mol/L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of HCl(aq) to KOH(aq).	HCl(aq) + KOH(aq) $\rightarrow$ KCl(aq) + H$_2$O(ℓ) The balanced chemical equation for the reaction gives a mole ratio $\dfrac{[\text{HCl(aq)}]}{[\text{KOH(aq)}]} = \dfrac{1}{1}$.
At the equivalence point in a titration, $n_{\text{H}_3\text{O}^+} = n_{\text{OH}^-}$. Since KOH(aq) is a strong base, [KOH] = [OH$^-$]. Since HCl(aq) is a strong acid, [HCl] = [H$_3$O$^+$].	$n_{\text{H}_3\text{O}^+} = n_{\text{OH}^-}$ $n_{\text{HCl}} = n_{\text{KOH}}$
Determine the concentration of HCl(aq) using the relationship $n_{\text{HCl}} = n_{\text{KOH}}$ and the formula $n = cV$.	$n_{\text{HCl}} = n_{\text{KOH}}$ $c_{\text{HCl}}V_{\text{HCl}} = c_{\text{KOH}}V_{\text{KOH}}$ $c_{\text{HCl}}(0.02865 \text{ L}) = (1.75 \text{ mol/L})(0.01876 \text{ L})$ $c_{\text{HCl}} = \dfrac{(1.75 \text{ mol/L})(0.01876 \text{ \cancel{L}})}{(0.02865 \text{ \cancel{L}})}$ $= 1.1458 \text{ mol/L}$ $= 1.15 \text{ mol/L}$

Check Your Solution
Since the volume of HCl(aq) is greater than the volume of KOH(aq), it is reasonable that the concentration of HCl(aq) will be less than the concentration of KOH(aq).

100. If 17.50 mL of 0.25 mol/L nitric acid, $HNO_3(aq)$, is titrated to a pH of 7.00 with 11.27 mL of sodium hydroxide, $NaOH(aq)$, what was the original concentration of the sodium hydroxide solution used in the titration?

What Is Required?
You must find the concentration of sodium hydroxide used in a neutralization reaction.

What Is Given?
At the equivalence point, the volume of $NaOH(aq)$ used is 11.27 mL. The $HNO_3(aq)$ used has a volume of 17.50 mL and a concentration of 0.25 mol/L.

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction and determine the mole ratio of $HNO_3(aq)$ to $NaOH(aq)$.	$HNO_3(aq) + NaOH(aq) \rightarrow$ $NaNO_3(aq) + H_2O(\ell)$ The balanced chemical equation for the reaction gives a mole ratio $\dfrac{[HNO_3(aq)]}{[NaOH(aq)]} = \dfrac{1}{1}$.
Since $NaOH(aq)$ is a strong base, $[NaOH] = [OH^-]$. Since $HCl(aq)$ is a strong acid, $[HCl] = [H_3O^+]$ At the equivalence point in a titration, $n_{H_3O^+} = n_{OH^-}$.	$n_{HNO_3} = n_{NaOH}$ $n_{H_3O^+} = n_{OH^-}$
Determine the concentration of $NaOH(aq)$ using the relationship $n_{NaOH} = n_{HNO_3}$ and the formula $n = cV$.	$n_{NaOH} = n_{HNO_3}$ $c_{NaOH}V_{NaOH} = c_{HNO_3}V_{HNO_3}$ $c_{NaOH}(0.01127 \text{ L}) = (0.25 \text{ mol/L})(0.01750 \text{ L})$ $c_{NaOH} = \dfrac{(0.25 \text{ mol/L})(0.01750 \, \cancel{L})}{(0.01127 \, \cancel{L})}$ $= 0.3882 \text{ mol/L}$ $= 0.39 \text{ mol/L}$

Check Your Solution
Since the volume of $HNO_3(aq)$ is greater than the volume of $NaOH(aq)$, it is reasonable that the concentration of $NaOH(aq)$ will be greater than the concentration $HNO_3(aq)$.

Calculating pH During a Weak Acid-Strong Base Titration
(Student textbook page 539)

101. Determine the pH at a point in the titration when 20.00 mL of 0.2000 mol/L acetic acid, $CH_3COOH(aq)$, has been added to 17.00 mL of 0.1500 mol/L potassium hydroxide, $KOH(aq)$.

What Is Required?
You need to find the pH at a certain point in the titration using acetic acid and potassium hydroxide solutions.

What Is Given?
You know the volume of the acetic acid: $V_{CH_3COOH} = 20.00$ mL

You know the concentration of the acetic acid: $c_{CH_3COOH} = 0.2000$ mol/L

You know the volume of the potassium hydroxide: $V_{KOH} = 17.00$ mL

You know the concentration of the potassium hydroxide: $c_{KOH} = 0.1500$ mol/L
From Appendix B, you can find that $K_a = 1.8 \times 10^{-5}$.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between $CH_3COOH(aq)$ and $KOH(aq)$.	Because KOH is a strong base, the reaction essentially goes to completion. $CH_3COOH(aq)\ (aq) + KOH(aq) \rightarrow$ $\qquad\qquad\qquad CH_3COOK(aq) + H_2O(\ell)$
Use the formula $n = cV$ to calculate the amount in moles, n, of $CH_3COOH(aq)$ present before any base was added.	$n_{i(CH_3COOH)} = cV$ $= \left(0.200\ \dfrac{mol}{L}\right)(20.00\ mL)\left(\dfrac{1\ L}{1000\ mL}\right)$ $= 4.000 \times 10^{-3}\,mol$
Use the formula $n = cV$ to calculate the amount in moles, n, of $KOH(aq)$.	$n_{KOH} = cV$ $= \left(0.1500\ \dfrac{mol}{L}\right)(17.00\ mL)\left(\dfrac{1\ L}{1000\ mL}\right)$ $= 2.550 \times 10^{-3}\,mol$

Calculate the amount in moles, n, of $CH_3COO^-(aq)$ present after the base was added.	Because the reaction goes to completion and because the ratio of CH_3COO^- to KOH is 1:1, for every mole of KOH added, one mole of CH_3COOH will be converted into CH_3COO^-. Therefore: $n_{CH_3COO^-} = n_{KOH}$ $= 2.550 \times 10^{-3}$ mol
Calculate the amount in moles of unreacted $CH_3COOH(aq)$.	$n_{(\text{excess } CH_3COOH)} = n_{(i\ CH_3COOH)} - n_{(CH_3COO^-\ \text{formed})}$ $= 4.000 \times 10^{-3}$ mol $- 2.550 \times 10^{-3}$ mol $= 1.450 \times 10^{-3}$ mol
Determine the final volume of the solution.	$V = V_{CH_3COOH} + V_{KOH}$ $= 20.00$ mL $+ 17.00$ mL $= 37.00$ mL $\times \dfrac{1\ L}{1000\ mL}$ $= 0.03700$ L
Use $c = \dfrac{n}{V}$ to determine the concentrations of $CH_3COOH(aq)$ and $CH_3COO^-(aq)$ in mol/L.	$c_{CH_3COOH} = \dfrac{n}{V}$ $= \dfrac{1.450 \times 10^{-3}\ \text{mol}}{0.0370\ \text{L}}$ $= 3.9189 \times 10^{-2}$ mol/L $c_{CH_3COO^-} = \dfrac{n}{V}$ $= \dfrac{2.550 \times 10^{-3}\ \text{mol}}{0.0370\ \text{L}}$ $= 6.8919 \times 10^{-2}$ mol/L

Write equations for any equilibrium reactions occurring and determine whether the solution has buffering capacity.	$CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons$ $\qquad\qquad CH_3COO^-(aq) + H_3O^+(aq)$ $CH_3COO^-(aq) + H_2O(\ell) \rightleftharpoons$ $\qquad\qquad CH_3COOH\ (aq) + OH^-(aq)$ The solution has buffering capacity. The reactions of acetic acid and acetate reacting with water balance each other and very little change actually occurs. Therefore, you can use the amounts of acetic acid and acetate present after the reaction with KOH and the equilibrium constant to find the concentration of hydronium ion, H_3O^+ present.
Write the equilibrium expression, substitute in the calculated concentrations and solve for the concentration of the hydronium ion.	$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$ $$1.8 \times 10^{-5} = \frac{(6.8919 \times 10^{-2})(x)}{3.9189 \times 10^{-2}}$$ $$x = \frac{(1.8 \times 10^{-5})(3.9189 \times 10^{-2})}{6.8919 \times 10^{-2}}$$ $$= 1.0235 \times 10^{-5}$$ $[H_3O^+] = x$ $\qquad = 1.0235 \times 10^{-5}$ mol/L
Calculate pH using: $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $\qquad = -\log (1.0235 \times 10^{-5})$ $\qquad = 4.98989$ $\qquad = 4.99$

Check Your Solution
The pH is consistent with having unreacted CH_3COOH present. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answers seem reasonable.

102. When hydrogen cyanide, HCN(g), is dissolved in water, it forms a very weak acid, $K_a = 6.2 \times 10^{-10}$. If a 50.00 mL sample of 0.1000 mol/L hydrocyanic acid, HCN(aq), is titrated with 8.00 mL of 0.1000 mol/L potassium hydroxide, KOH(aq), what is the pH of the solution after the base has been added?

What Is Required?
You need to find the pH at a certain point in the titration using hydrocyanic acid and potassium hydroxide solutions.

What Is Given?
You know the volume of the hydrocyanic acid: $V_{HCN} = 50.00$ mL

You know the concentration of the hydrocyanic acid: $c_{HCN} = 0.1000$ mol/L

You know the volume of the potassium hydroxide: $V_{KOH} = 8.00$ mL

You know the concentration of the potassium hydroxide: $c_{KOH} = 0.1000$ mol/L

You know the acid dissociation constant: $K_a = 6.2 \times 10^{-10}$

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between HCN(aq) and KOH(aq).	Because KOH is a strong base, the reaction essentially goes to completion. $HCN(aq) + KOH(aq) \rightarrow KCN(aq) + H_2O(\ell)$
Use the formula $n = cV$ to calculate the amount in moles, n, of $CH_3COOH(aq)$ present before any base was added.	$n_{HCN} = cV$ $= \left(0.100 \ \frac{mol}{L}\right)(50.00 \ mL)\left(\frac{1 \ L}{1000 \ mL}\right)$ $= 5.00 \times 10^{-3}$ mol
Use the formula $n = cV$ to calculate the amount in moles, n, of KOH(aq).	$n_{KOH} = cV$ $= \left(0.100 \ \frac{mol}{L}\right)(8.00 \ mL)\left(\frac{1 \ L}{1000 \ mL}\right)$ $= 8.00 \times 10^{-4}$ mol
Calculate the amount in moles, n, of CN^-(aq) present after the base was added.	Because the reaction goes to completion and because the ratio of CN^- to KOH is 1:1, for every mole of KOH added, one mole of HCN will be converted into CN^-. Therefore: $n_{CN^-} = n_{KOH}$ $= 8.00 \times 10^{-4}$ mol
Calculate the amount in moles of unreacted HCN(aq).	$n_{(excess \ HCN)} = n_{(i \ HCN)} - n_{(CN^- \ formed)}$ $= 5.00 \times 10^{-3}$ mol $- 8.00 \times 10^{-4}$ mol $= 4.20 \times 10^{-3}$ mol

Determine the final volume of the solution.	$V = V_{HCN} + V_{KOH}$ $= 50.00 \text{ mL} + 8.00 \text{ mL}$ $= 58.00 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}$ $= 0.05800 \text{ L}$
Use $c = \dfrac{n}{V}$ to determine the concentrations of HCN(aq) and CN$^-$(aq) in mol/L.	$c_{HCN} = \dfrac{n}{V}$ $= \dfrac{4.20 \times 10^{-3} \text{ mol}}{0.0580 \text{ L}}$ $= 7.241 \times 10^{-2} \text{ mol/L}$ $c_{CN^-} = \dfrac{n}{V}$ $= \dfrac{8.00 \times 10^{-4} \text{ mol}}{0.0580 \text{ L}}$ $= 1.379 \times 10^{-2} \text{ mol/L}$
Write equations for any equilibrium reactions occurring and determine whether the solution has buffering capacity.	$HCN(aq) + H_2O(\ell) \rightleftharpoons CN^-(aq) + H_3O^+(aq)$ $CN^-(aq) + H_2O(\ell) \rightleftharpoons HCN(aq) + OH^-(aq)$ The solution has buffering capacity. The reactions of HCN and CN$^-$ reacting with water balance each other and very little change actually occurs. Therefore, you can use the amounts of HCN and CN$^-$ present after the reaction with KOH and the equilibrium constant to find the concentration of hydronium ion, H$_3$O$^+$ present.
Write the equilibrium expression, substitute in the calculated concentrations and solve for the concentration of the hydronium ion.	$K_a = \dfrac{[CN^-][H_3O^+]}{[HCN]}$ $6.2 \times 10^{-10} = \dfrac{(1.379 \times 10^{-2})(x)}{7.241 \times 10^{-2}}$ $x = \dfrac{(6.2 \times 10^{-10})(7.241 \times 10^{-2})}{1.379 \times 10^{-2}}$ $= 3.2557 \times 10^{-9}$ $[H_3O^+] = x$ $= 3.2557 \times 10^{-9} \text{ mol/L}$
Calculate pH using: $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $= -\log (3.2557 \times 10^{-9})$ $= 8.487$ $= 8.49$

Check Your Solution

The pH is consistent with having unreacted HCN present. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

103. What is the pH of the titration in question 102 at the equivalence point?

What Is Required?
You need to find the pH at the equivalence point in the titration using hydrocyanic acid and potassium hydroxide solutions.

What Is Given?
You know the volume of the hydrocyanic acid: $V_{HCN} = 50.00$ mL

You know the concentration of the hydrocyanic acid: $c_{HCN} = 0.1000$ mol/L

You know the concentration of the potassium hydroxide: $c_{KOH} = 0.1000$ mol/L

You know the acid dissociation constant: $K_a = 6.2 \times 10^{-10}$

Determine the amount in moles, n, of KOH(aq) that reacted.	At the equivalence point, the amount of KOH(aq) added is equal to the amount of HCN(aq) initially present in the solution. $$n_{KOH} = n_{i(HCN)}$$ $$= c_{i(HCN)}V_{i(HCN)}$$ $$= (0.1000\ \frac{mol}{\cancel{L}})(50.00\ \cancel{mL})\left(\frac{1\cancel{L}}{1000\ \cancel{mL}}\right)$$ $$= 5.000 \times 10^{-3}\ mol$$
Determine the volume of KOH added.	$$V = \frac{n}{c}$$ $$= \frac{5.000 \times 10^{-3}\ \cancel{mol}}{0.1000\ \frac{\cancel{mol}}{L}}$$ $$= 0.050\ \cancel{L} \times \frac{1000\ mL}{\cancel{L}}$$ $$= 50.00\ mL$$
Determine the total volume, V, at the equivalence point.	$$V_{total} = V_{HCN} + V_{KOH}$$ $$= 50.00\ mL + 50.00\ mL$$ $$= 100.0\ mL$$
Determine the amount in moles, n, of CN$^-$ present in the solution at the equivalence point.	Because you know that the amount of KOH added was equal to the amount of HCN originally present, and you know that the reaction went to completion, you know that the amount of CN$^-$ present when the equivalence is reached is: $$n_{CN^-} = n_{i(HCN)}$$ $$= 5.000 \times 10^{-3}\ mol$$

Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[CN^-]$.	$\left[CN^-\right] = \dfrac{n}{V}$ $= \dfrac{5.000 \times 10^{-3}\ \text{mol}}{100.0\ \text{mL}} \times \dfrac{1000\ \text{mL}}{\text{L}}$ $= 5.000 \times 10^{-2}\ \text{mol/L}$
Write the equation for the acid-dissociation of CN^- (aq). Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table on the following page.

	CN^-(aq)	+	$H_2O(\ell)$	$\leftrightharpoons$	HCN(aq)	+	OH^- (aq)
	$[CN^-]$ (mol/L)		[H2O] (mol/L)		[HCN] (mol/L)		$[OH^-]$ (mol/L)
I	5.000×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$5.000 \times 10^{-2} - x$				x		x

Write the expression for the ionization constant.	$K_b = \dfrac{[HCN][OH^-]}{[CN^-]}$
Calculate K_b from the known K_b	$K_b = \dfrac{K_w}{K_a}$ $= \dfrac{1 \times 10^{-14}}{6.2 \times 10^{-10}}$ $= 1.61 \times 10^{-5}$
Test to determine whether the approximation can be applied.	$1000 \times K_b = 1000 \times 1.61 \times 10^{-5} = 1.61 \times 10^{-2}$ $1.61 \times 10^{-2} < 5.00 \times 10^{-2}$ The approximation is valid. $5.000 \times 10^{-2} - x \approx 5.000 \times 10^{-2}$
Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$K_b = \dfrac{[HCN][OH^-]}{[CN^-]}$ $1.61 \times 10^{-5} = \dfrac{(x)(x)}{5.000 \times 10^{-2}}$ $x^2 = \left(5.000 \times 10^{-2}\right)\left(1.61 \times 10^{-5}\right)$ $x^2 = 8.065 \times 10^{-7}$ $x = 8.98 \times 10^{-4}$ $[OH^-] = x$ $ = 8.98 \times 10^{-4}\ \text{mol/L}$

Calculate pOH using the expression $pH = -\log [OH^-]$.	$pH = -\log [OH^-]$ $ = -\log (9.98 \times 10^{-4})$ $ = 3.467$
Calculate the pH from the pOH.	$pH = 14 - pOH$ $ = 10.95$

Check Your Solution
The pH at the equialence point for the titration of a weak acid with a strong base is expected to be greater than 7. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

104. A chemist titrated 35.00 mL of a 0.150 mol/L solution of hypobromous acid, HBrO(aq), $K_a = 2.8 \times 10^{-9}$. Calculate the resulting pH after the addition of 15.00 mL of 0.1000 mol/L sodium hydroxide solution, NaOH(aq).

What Is Required?
You need to find the pH at a certain point in the titration using hypobromous acid and potassium hydroxide solutions.

What Is Given?
You know the volume of the hypobromous acid: $V_{HBrO} = 35.00$ mL

You know the concentration of the hypobromous acid: $c_{HBrO} = 0.1500$ mol/L

You know the volume of the sodium hydroxide: $V_{NaOH} = 15.00$ mL

You know the concentration of the sodium hydroxide: $c_{NaOH} = 0.1000$ mol/L

You are given the acid dissociation constant for hypobromous acid: $K_a = 2.8 \times 10^{-9}$.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between HBrO (aq) and NaOH(aq).	Because NaOH is a strong base, the reaction essentially goes to completion. $\text{HBrO(aq) (aq)} + \text{NaOH(aq)} \rightarrow$ $\quad\quad\quad\quad\quad\quad\quad\quad \text{NaBrO(aq)} + \text{H}_2\text{O}(\ell)$
Use the formula $n = cV$ to calculate the amount in moles, n, of HBrO(aq) present before any base was added.	$n_{HBrO} = cV$ $\quad\quad = \left(0.1500 \ \dfrac{\text{mol}}{L}\right)(35.00 \ \text{mL})\left(\dfrac{1 \ L}{1000 \ \text{mL}}\right)$ $\quad\quad = 5.250 \times 10^{-3} \text{ mol}$
Use the formula $n = cV$ to calculate the amount in moles, n, of NaOH(aq).	$n_{NaOH} = cV$ $\quad\quad = \left(0.1000 \ \dfrac{\text{mol}}{L}\right)(15.00 \ \text{mL})\left(\dfrac{1 \ L}{1000 \ \text{mL}}\right)$ $\quad\quad = 1.500 \times 10^{-3} \text{ mol}$
Calculate the amount in moles, n, of BrO^-(aq) present after the base was added.	Because the reaction goes to completion and because the ratio of BrO^- to NaOH is 1:1, for every mole of NaOH added, one mole of HBrO will be converted into BrO^-. Therefore: $n_{BrO^-} = n_{NaOH}$ $\quad\quad = 1.500 \times 10^{-3} \text{ mol}$
Calculate the amount in moles of unreacted HBrO(aq).	$n_{(\text{excess HBrO})} = n_{(i \ \text{HBrO})} - n_{(BrO^- \ \text{formed})}$ $\quad\quad = 5.250 \times 10^{-3} \text{ mol} - 1.500 \times 10^{-3} \text{ mol}$ $\quad\quad = 3.750 \times 10^{-3} \text{ mol}$

Determine the final volume of the solution.	$V = V_{\text{HBrO}} + V_{\text{NaOH}}$ $= 35.00 \text{ mL} + 15.00 \text{ mL}$ $= 50.00 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}$ $= 0.05000 \text{ L}$
Use $c = \dfrac{n}{V}$ to determine the concentrations of HBrO(aq) and BrO⁻(aq) in mol/L.	$c_{\text{HBrO}} = \dfrac{n}{V}$ $= \dfrac{3.750 \times 10^{-3} \text{ mol}}{0.0500 \text{ L}}$ $= 7.500 \times 10^{-2} \text{ mol/L}$ $c_{\text{BrO}^-} = \dfrac{n}{V}$ $= \dfrac{1.500 \times 10^{-3} \text{ mol}}{0.0500 \text{ L}}$ $= 3.000 \times 10^{-2} \text{ mol/L}$
Write equations for any equilibrium reactions occurring and determine whether the solution has buffering capacity.	$\text{HBrO(aq)} + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{BrO}^-\text{(aq)} + \text{H}_3\text{O}^+\text{(aq)}$ $\text{BrO}^-\text{(aq)} + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HBrO (aq)} + \text{OH}^-\text{(aq)}$ The solution has buffering capacity. The reactions of HBrO and BrO⁻ reacting with water balance each other and very little change actually occurs. Therefore, you can use the amounts of HBrO and BrO⁻ present after the reaction with NaOH and the equilibrium constant to find the concentration of hydronium ion, H_3O^+ present.
Write the equilibrium expression, substitute in the calculated concentrations and solve for the concentration of the hydronium ion.	$K_a = \dfrac{[\text{BrO}^-][\text{H}_3\text{O}^+]}{[\text{HBrO}]}$ $2.8 \times 10^{-9} = \dfrac{(3.000 \times 10^{-2})(x)}{7.500 \times 10^{-2}}$ $x = \dfrac{(2.8 \times 10^{-9})(7.500 \times 10^{-2})}{3.000 \times 10^{-2}}$ $= 7.000 \times 10^{-9}$ $[\text{H}_3\text{O}^+] = x$ $= 7.000 \times 10^{-9} \text{ mol/L}$

Calculate pH using: $\text{pH} = -\log [H_3O^+].$	$\text{pH} = -\log [H_3O^+]$ $= -\log (7.000 \times 10^{-9})$ $= 8.1549$ $= 8.15$

Check Your Solution

The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a.

105. A solution of 50.00 mL of a 0.120 mol/L nitrous acid, $HNO_2(aq)$, is titrated with 0.1000 mol/L of potassium hydroxide, $KOH(aq)$. Determine the resulting pH of the solution after 11.25 mL of the base has been added. For the acid, $K_a = 5.6 \times 10^{-4}$.

What Is Required?

You need to find the pH at a certain point in the titration using nitrous acid and potassium hydroxide solutions.

What Is Given?

You know the volume of the nitrous acid: $V_{HNO_2} = 50.00$ mL

You know the concentration of the nitrous acid: $c_{HNO_2} = 0.120$ mol/L

You know the volume of the potassium hydroxide: $V_{KOH} = 11.50$ mL

You know the concentration of the potassium hydroxide: $c_{KOH} = 0.1000$ mol/L

You know the acid dissociation constant: $K_a = 5.6 \times 10^{-4}$

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between $HNO_2(aq)$ and $KOH(aq)$.	Because KOH is a strong base, the reaction essentially goes to completion. $$HNO_2(aq) + KOH(aq) \rightarrow KNO_2(aq) + H_2O(\ell)$$
Use the formula $n = cV$ to calculate the amount in moles, n, of $HNO_2(aq)$ present before any base was added.	$n_{HNO_2} = cV$ $$= \left(0.120\ \frac{mol}{L} \right) (50.00\ mL) \left(\frac{1\ L}{1000\ mL} \right)$$ $$= 6.00 \times 10^{-3}\ mol$$
Use the formula $n = cV$ to calculate the amount in moles, n, of $KOH(aq)$.	$n_{KOH} = cV$ $$= \left(0.100\ \frac{mol}{L} \right) (11.25\ mL) \left(\frac{1\ L}{1000\ mL} \right)$$ $$= 1.125 \times 10^{-3}\ mol$$
Calculate the amount in moles, n, of $NO_2^-(aq)$ present after the base was added.	Because the reaction goes to completion and because the ratio of NO_2^- to KOH is 1:1, for every mole of KOH added, one mole of HNO_2 will be converted into NO_2^-. Therefore: $n_{NO_2^-} = n_{KOH}$ $$= 1.125 \times 10^{-3}\ mol$$
Calculate the amount in moles of unreacted $HNO_2(aq)$.	$n_{(excess\ HNO_2)} = n_{(i\ HNO_2)} - n_{(NO_2^-\ formed)}$ $$= 6.000 \times 10^{-3}\ mol - 1.125 \times 10^{-3}\ mol$$ $$= 4.875 \times 10^{-3}\ mol$$

Determine the final volume of the solution.	$V = V_{HNO_2} + V_{KOH}$ $= 50.00 \text{ mL} + 11.25 \text{ mL}$ $= 61.25 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}$ $= 0.06125 \text{ L}$
Use $c = \dfrac{n}{V}$ to determine the concentrations of $HNO_2(aq)$ and $NO_2^-(aq)$ in mol/L.	$c_{HNO_2} = \dfrac{n}{V}$ $= \dfrac{4.875 \times 10^{-3} \text{ mol}}{0.06125 \text{ L}}$ $= 7.9592 \times 10^{-2} \text{ mol/L}$ $c_{NO_2^-} = \dfrac{n}{V}$ $= \dfrac{1.125 \times 10^{-3} \text{ mol}}{0.06125 \text{ L}}$ $= 1.8367 \times 10^{-2} \text{ mol/L}$
Write equations for any equilibrium reactions occurring and determine whether the solution has buffering capacity.	$HNO_2(aq) + H_2O(\ell) \rightleftharpoons NO_2^-(aq) + H_3O^+(aq)$ $NO_2^-(aq) + H_2O(\ell) \rightleftharpoons HNO_2(aq) + OH^-(aq)$ The solution has buffering capacity. The reactions of HNO_2 and NO_2^- reacting with water balance each other and very little change actually occurs. Therefore, you can use the amounts of HNO_2 and NO_2^- present after the reaction with KOH and the equilibrium constant to find the concentration of hydronium ion, H_3O^+ present.
Write the equilibrium expression, substitute in the calculated concentrations and solve for the concentration of the hydronium ion.	$K_a = \dfrac{[NO_2^-][H_3O^+]}{[HNO_2]}$ $5.6 \times 10^{-4} = \dfrac{(1.8367 \times 10^{-2})(x)}{7.9592 \times 10^{-2}}$ $x = \dfrac{(5.6 \times 10^{-4})(7.9592 \times 10^{-2})}{1.8367 \times 10^{-2}}$ $= 2.43267 \times 10^{-3}$ $[H_3O^+] = x$ $= 2.43267 \times 10^{-3} \text{ mol/L}$

Calculate pH using: $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $= -\log (2.43267 \times 10^{-3})$ $= 2.615$ $= 2.61$

Check Your Solution

The pH is consistent with having unreacted HNO_2 present. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

106. If 100.00 mL of a 0.400 mol/L hydrofluoric acid solution, HF(aq), is titrated with a 0.2000 mol/L sodium hydroxide solution, NaOH(aq), determine the pH that results when 20.00 mL of the base is added. For the hydrofluoric acid solution, $K_a = 6.3 \times 10^{-4}$.

What Is Required?

You need to find the pH at a certain point in the titration of hydrofluoric acid with sodium hydroxide solution.

What Is Given?

You know the volume of the hydrofluoric acid: $V_{HF} = 100.00$ mL

You know the concentration of the hydrofluoric acid: $c_{HF} = 0.4000$ mol/L

You know the volume of the sodium hydroxide: $V_{NaOH} = 20.00$ mL

You know the concentration of the sodium hydroxide: $c_{NaOH} = 0.2000$ mol/L

You are given the acid dissociation constant for hydrofluoric acid: $K_a = 6.3 \times 10^{-4}$

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between HF (aq) and NaOH(aq).	Because NaOH is a strong base, the reaction essentially goes to completion. HF(aq) (aq) + NaOH(aq) → NaF(aq) + H₂O(ℓ)
Use the formula $n = cV$ to calculate the amount in moles, n, of HF(aq) present before any base was added.	$n_{HF} = cV$ $= \left(0.4000 \ \dfrac{\text{mol}}{\cancel{L}}\right)\left(100.00 \ \cancel{mL}\right)\left(\dfrac{1 \ \cancel{L}}{1000 \ \cancel{mL}}\right)$ $= 4.000 \times 10^{-2}$ mol
Use the formula $n = cV$ to calculate the amount in moles, n, of NaOH(aq).	$n_{NaOH} = cV$ $= \left(0.2000 \ \text{mol}/\cancel{L}\right)\left(20.00 \ \cancel{mL}\right)\left(\dfrac{1 \ \cancel{L}}{1000 \ \cancel{mL}}\right)$ $= 4.000 \times 10^{-3}$ mol
Calculate the amount in moles, n, of F⁻(aq) present after the base was added.	Because the reaction goes to completion and because the ratio of F⁻ to NaOH is 1:1, for every mole of NaOH added, one mole of HF will be converted into F⁻. Therefore: $n_{F^-} = n_{NaOH}$ $= 4.000 \times 10^{-3}$ mol
Calculate the amount in moles of unreacted HF(aq).	$n_{(\text{excess HF})} = n_{(i \ HF)} - n_{(F^- \ \text{formed})}$ $= 4.000 \times 10^{-2}$ mol $- 4.000 \times 10^{-3}$ mol $= 3.600 \times 10^{-2}$ mol

Determine the final volume of the solution.	$V = V_{HF} + V_{NaOH}$ $= 100.00 \text{ mL} + 20.00 \text{ mL}$ $= 120.00 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}$ $= 0.1200 \text{ L}$
Use $c = \dfrac{n}{V}$ to determine the concentrations of HF(aq) and F⁻(aq) in mol/L.	$c_{HF} = \dfrac{n}{V}$ $= \dfrac{3.600 \times 10^{-2} \text{ mol}}{0.1200 \text{ L}}$ $= 0.300 \times 10^{-1} \text{ mol/L}$ $c_{F^-} = \dfrac{n}{V}$ $= \dfrac{4.000 \times 10^{-3} \text{ mol}}{0.1200 \text{ L}}$ $= 3.333 \times 10^{-2} \text{ mol/L}$
Write equations for any equilibrium reactions occurring and determine whether the solution has buffering capacity.	$HF(aq) + H_2O(\ell) \rightleftharpoons F^-(aq) + H_3O^+(aq)$ $F^-(aq) + H_2O(\ell) \rightleftharpoons HF(aq) + OH^-(aq)$ The solution has buffering capacity. The reactions of HF and F⁻ reacting with water balance each other and very little change actually occurs. Therefore, you can use the amounts of HF and F⁻ present after the reaction with NaOH and the equilibrium constant to find the concentration of hydronium ion, H_3O^+ present.
Write the equilibrium expression, substitute in the calculated concentrations and solve for the concentration of the hydronium ion.	$K_a = \dfrac{[F^-][H_3O^+]}{[HF]}$ $6.3 \times 10^{-4} = \dfrac{(3.333 \times 10^{-2})(x)}{3.000 \times 10^{-1}}$ $x = \dfrac{(6.3 \times 10^{-4})(3.000 \times 10^{-1})}{3.333 \times 10^{-2}}$ $= 5.6706 \times 10^{-3}$ $[H_3O^+] = x$ $= 5.6706 \times 10^{-3} \text{ mol/L}$

Calculate pH using: $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $ = -\log (5.6706 \times 10^{-3})$ $ = 2.2464$ $ = 2.25$

Check Your Solution

The pH is consistent with having unreacted HF present. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

107. A chemist titrates 50.00 mL of a 0.350 mol/L solution of hypochlorous acid, HOCl(aq), with $K_a = 4.0 \times 10^{-8}$, with 0.150 mol/L of potassium hydroxide, KOH(aq). Calculate the pH after the addition of 5.00 mL of the base.

What Is Required?

You need to find the pH at a certain point in the titration of hypochlorous acid, HOCl(aq), with potassium hydroxide, KOH(aq) solution.

What Is Given?

You know the volume of the hypochlorous acid: $V_{HOCl} = 50.00$ mL

You know the concentration of the hypochlorous acid: $c_{HOCl} = 0.350$ mol/L

You know the volume of the potassium hydroxide: $V_{KOH} = 5.00$ mL

You know the concentration of the potassium hydroxide: $c_{KOH} = 0.150$ mol/L

You know the acid dissociation constant: $K_a = 4.0 \times 10^{-8}$

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between HOCl(aq) and KOH(aq).	Because KOH is a strong base, the reaction essentially goes to completion. $$HOCl(aq) + KOH(aq) \rightarrow KOCl(aq) + H_2O(\ell)$$
Use the formula $n = cV$ to calculate the amount in moles, n, of HOCl(aq) present before any base was added.	$$n_{HOCl} = cV$$ $$= \left(0.350\ \frac{mol}{\cancel{L}}\right)(50.00\ \cancel{mL})\left(\frac{1\ \cancel{L}}{1000\ \cancel{mL}}\right)$$ $$= 1.750 \times 10^{-2}\ mol$$
Use the formula $n = cV$ to calculate the amount in moles, n, of KOH(aq).	$$n_{KOH} = cV$$ $$= \left(0.150\ \frac{mol}{\cancel{L}}\right)(5.00\ \cancel{mL})\left(\frac{1\ \cancel{L}}{1000\ \cancel{mL}}\right)$$ $$= 7.500 \times 10^{-4}\ mol$$
Calculate the amount in moles, n, of OCl⁻(aq) present after the base was added.	Because the reaction goes to completion and because the ratio of OCl⁻ to KOH is 1:1, for every mole of KOH added, one mole of HOCl will be converted into OCl⁻. Therefore: $$n_{OCl^-} = n_{KOH}$$ $$= 7.500 \times 10^{-4}\ mol$$
Calculate the amount in moles of unreacted HOCl(aq).	$$n_{(excess\ HOCl)} = n_{(i\ HOCl)} - n_{(OCl^-\ formed)}$$ $$= 1.750 \times 10^{-2}\ mol - 7.500 \times 10^{-4}\ mol$$ $$= 1.675 \times 10^{-2}\ mol$$

Determine the final volume of the solution.	$V = V_{HOCl} + V_{KOH}$ $= 50.00 \text{ mL} + 5.00 \text{ mL}$ $= 55.00 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}$ $= 0.0550 \text{ L}$
Use $c = \dfrac{n}{V}$ to determine the concentrations of HOCl(aq) and OCl⁻(aq) in mol/L.	$c_{HOCl} = \dfrac{n}{V}$ $= \dfrac{1.675 \times 10^{-2} \text{ mol}}{0.0550 \text{ L}}$ $= 0.3045 \text{ mol/L}$ $c_{OCl^-} = \dfrac{n}{V}$ $= \dfrac{7.500 \times 10^{-4} \text{ mol}}{0.0550 \text{ L}}$ $= 1.3636 \times 10^{-2} \text{ mol/L}$
Write equations for any equilibrium reactions occurring and determine whether the solution has buffering capacity.	$HOCl(aq) + H_2O(\ell) \rightleftharpoons OCl^-(aq) + H_3O^+(aq)$ $OCl^-(aq) + H_2O(\ell) \rightleftharpoons HOCl(aq) + OH^-(aq)$ The solution has buffering capacity. The reactions of HOCl and OCl⁻ reacting with water balance each other and very little change actually occurs. Therefore, you can use the amounts of HOCl and OCl⁻ present after the reaction with KOH and the equilibrium constant to find the concentration of hydronium ion, H_3O^+ present.
Write the equilibrium expression, substitute in the calculated concentrations and solve for the concentration of the hydronium ion.	$K_a = \dfrac{[OCl^-][H_3O^+]}{[HOCl]}$ $4.0 \times 10^{-8} = \dfrac{(1.3636 \times 10^{-2})(x)}{0.3045}$ $x = \dfrac{(4.0 \times 10^{-8})(0.3045)}{1.3636 \times 10^{-2}}$ $= 8.9322 \times 10^{-7}$ $[H_3O^+] = x$ $= 8.9322 \times 10^{-7} \text{ mol/L}$

Calculate pH using: $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $= -\log (8.9322 \times 10^{-7})$ $= 6.049$ $= 6.05$

Check Your Solution
The pH is consistent with having unreacted HOCl present. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

108. Determine the pH of 25.00 mL of a solution of 0.100 mol/L hypobromous acid, HBrO(aq), at the equivalence point, when titrated with a 0.100 mol/L sodium hydroxide solution, NaOH(aq).

What Is Required?

You need to find the pH at the equivalence point in the titration of hypobromous acid with sodium hydroxide solution.

What Is Given?

You know the volume of the hypobromous acid: $V_{HBrO} = 25.00$ mL

You know the concentration of the hypobromous acid: $c_{HBrO} = 0.100$ mol/L

You know the concentration of the sodium hydroxide: $c_{NaOH} = 0.100$ mol/L

From question 104, you know that, for hypobromous acid: $K_a = 2.8 \times 10^{-9}$.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between HBrO (aq) and NaOH(aq).	Because NaOH is a strong base, the reaction essentially goes to completion. $$HBrO(aq)\ (aq) + NaOH(aq) \rightarrow$$ $$NaBrO(aq) + H_2O(\ell)$$
Use the formula $n = cV$ to calculate the amount in moles, n, of HBrO(aq) present before any base was added.	$n_{HBrO} = cV$ $$= \left(0.100\ \frac{\text{mol}}{\cancel{L}}\right)(25.00\ \cancel{\text{mL}})\left(\frac{1\ \cancel{L}}{1000\ \cancel{\text{mL}}}\right)$$ $$= 2.50 \times 10^{-3}\,\text{mol}$$
Determine the amount in moles, n, of NaOH(aq).	Because you know that the reaction reaches the equivalence point, you know that the amount of NaOH must be equal to the amount of HBrO. $n_{NaOH} = n_{HBrO}$ $$= 2.50 \times 10^{-3}\ \text{mol}$$
Calculate the amount in moles, n, of BrO⁻(aq) present after the base was added.	Because you know that the reaction reaches the equivalence point, you know that all of the HBrO was converted into BrO⁻ $n_{BrO^-} = n_{HBrO}$ $$= 2.50 \times 10^{-3}\ \text{mol}$$

Determine the volume of the NaOH that had to be added to bring the reaction to the equivalence point.	$V = \dfrac{n}{c}$ $= \dfrac{2.5 \times 10^{-3} \text{ mol}}{0.100 \dfrac{\text{mol}}{\text{L}}}$ $= 2.5 \times 10^{-2} \text{ L} \times \dfrac{1000 \text{ mL}}{\text{L}}$ $= 25.00 \text{ mL}$
Determine the final volume of the solution.	$V = V_{\text{HBrO}} + V_{\text{NaOH}}$ $= 25.00 \text{ mL} + 25.00 \text{ mL}$ $= 50.00 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}}$ $= 0.05000 \text{ L}$
Use $c = \dfrac{n}{V}$ to determine the final concentration of BrO$^-$(aq) in mol/L.	$c_{\text{BrO}^-} = \dfrac{n}{V}$ $= \dfrac{2.500 \times 10^{-3} \text{ mol}}{0.0500 \text{ L}}$ $= 5.00 \times 10^{-2} \text{ mol/L}$
Write the equation for the reaction of BrO$^-$ with water. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the [OH$^-$].	See the ICE table below.

	BrO$^-$(aq)	+	H$_2$O(ℓ)	$\leftrightarrows$	HBrO(aq)	+	OH$^-$ (aq)
	[BrO$^-$] (mol/L)		[H$_2$O] (mol/L)		[HBrO] (mol/L)		[OH$^-$] (mol/L)
I	5.00×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$5.00 \times 10^{-2} - x$				x		x

Write the expression for the ionization constant.	$K_b = \dfrac{[\text{HBrO}][\text{OH}^-]}{[\text{BrO}^-]}$
Calculate K_b from the known K_b	$K_b = \dfrac{K_w}{K_a}$ $= \dfrac{1 \times 10^{-14}}{2.8 \times 10^{-9}}$ $= 3.5714 \times 10^{-6}$

Test to determine whether the approximation can be applied.	$1000 \times K_b = 1000 \times 3.5714 \times 10^{-6}$ $\qquad\qquad = 3.5714 \times 10^{-3}$ $3.5714 \times 10^{-3} < 5.00 \times 10^{-2}$ The approximation is valid. $5.00 \times 10^{-2} - x \approx 5.00 \times 10^{-2}$
Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$K_b = \dfrac{[\text{HBrO}][\text{OH}^-]}{[\text{BrO}^-]}$ $3.5714 \times 10^{-6} = \dfrac{(x)(x)}{5.000 \times 10^{-2}}$ $x^2 = \left(5.000 \times 10^{-2}\right)\left(3.5714 \times 10^{-6}\right)$ $x^2 = 1.7857 \times 10^{-7}$ $x = 4.2258 \times 10^{-4}$ $[\text{OH}^-] = x$ $\qquad = 4.2258 \times 10^{-4}$ mol/L
Calculate pOH using the expression pH = –log [OH⁻].	pH = –log [OH⁻] $\qquad = -\log(4.2258 \times 10^{-4})$ $\qquad = 3.374$
Calculate the pH from the pOH.	pH = 14 – pOH $\qquad = 10.626$ $\qquad = 10.63$

Check Your Solution
The pH at the equialence point for the titration of a weak acid with a strong base is expected to be greater than 7. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

109. A chemist titrated 50.00 mL of a 0.150 mol/L acetic acid solution, $CH_3COOH(aq)$, with 0.300 mol/L sodium hydroxide, $NaOH(aq)$. Determine the pH of the resulting solution when 25.00 mL of the base has been added.

What Is Required?

You need to find the pH after adding acetic acid, $CH_3COOH(aq)$, to sodium hydroxide, $NaOH(aq)$.

What Is Given?

You know the volume of the acetic acid: $V_{CH_3COOH} = 50.00$ mL

You know the concentration of the acetic acid: $c_{CH_3COOH} = 0.150$ mol/L

You know the volume of the sodium hydroxide: $V_{NaOH} = 25.00$ mL

You know the concentration of the sodium hydroxide: $c_{NaOH} = 0.300$ mol/L

From Appendix B, you can find that $K_a = 1.8 \times 10^{-5}$.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between $CH_3COOH(aq)$ and $NaOH(aq)$.	Because NaOH is a strong base, the reaction essentially goes to completion. $CH_3COOH(aq) (aq) + NaOH(aq) \rightarrow$ $\qquad\qquad CH_3COONa(aq) + H_2O(\ell)$
Use the formula $n = cV$ to calculate the amount in moles, n, of $CH_3COOH(aq)$ present before any base was added.	$n_{i(CH_3COOH)} = cV$ $= \left(0.150 \dfrac{mol}{L}\right)(50.00 \text{ mL})\left(\dfrac{1 \text{ L}}{1000 \text{ mL}}\right)$ $= 7.500 \times 10^{-3} \text{ mol}$
Use the formula $n = cV$ to calculate the amount in moles, n, of $NaOH(aq)$.	$n_{NaOH} = cV$ $= \left(0.300 \dfrac{mol}{L}\right)(25.00 \text{ mL})\left(\dfrac{1 \text{ L}}{1000 \text{ mL}}\right)$ $= 7.500 \times 10^{-3} \text{ mol}$
Calculate the amount in moles, n, of $CH_3COO^-(aq)$ present after the base was added.	Because the reaction goes to completion and because the ratio of CH_3COO^- to NaOH is 1:1, for every mole of NaOH added, one mole of CH_3COOH will be converted into CH_3COO^-. Therefore: $n_{CH_3COO^-} = n_{NaOH}$ $\qquad = 7.50 \times 10^{-3} \text{ mol}$

Calculate the amount in moles of unreacted $CH_3COOH(aq)$.	$n_{(\text{excess } CH_3COOH)} = n_{(i\ CH_3COOH)} - n_{(CH_3COO^-\ \text{formed})}$ $= 7.50 \times 10^{-3}$ mol $- 7.50 \times 10^{-3}$ mol $= 0.0$ mol All of the acetic acid was converted into acetate ion so this reaction is at its equivalence point.
Determine the final volume of the solution.	$V = V_{CH_3COOH} + V_{NaOH}$ $= 50.00$ mL $+ 25.00$ mL $= 75.00 \text{ mL} \times \dfrac{1\text{ L}}{1000 \text{ mL}}$ $= 0.0750$ L
Use $c = \dfrac{n}{V}$ to determine the concentrations of $CH_3COO^-(aq)$ in mol/L.	$c_{CH_3COO^-} = \dfrac{n}{V}$ $= \dfrac{7.50 \times 10^{-3} \text{ mol}}{0.0750 \text{ L}}$ $= 0.100$ mol/L
Set up an ICE table to calculate the $[OH^-]$.	See the ICE table below.

	CH_3COO^- (aq)	$+$	$H_2O(\ell)$	$\leftrightharpoons$	$CH_3COOH(aq)$	$+$	OH^- (aq)
	$[CH_3COO^-]$ (mol/L)				$[CH_3COOH]$ (mol/L)		$[OH^-]$(mol/L)
I	1.00×10^{-1}				0		0
C	$-x$				$+x$		$+x$
E	$1.00 \times 10^{-1} - x$				$+x$		$+x$

Write the expression for the ionization constant.	$K_b = \dfrac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$
Calculate K_b from the known K_b	$K_b = \dfrac{K_w}{K_a}$ $= \dfrac{1 \times 10^{-14}}{1.8 \times 10^{-5}}$ $= 5.555 \times 10^{-10}$
Test to determine whether the approximation can be applied.	$1000 \times K_b = 1000 \times 5.555 \times 10^{-10}$ $\qquad\qquad = 5.555 \times 10^{-7}$ $5.555 \times 10^{-7} \ll 1.00 \times 10^{-1}$ The approximation is valid. $1.00 \times 10^{-1} - x \approx 1.00 \times 10^{-1}$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$$ $$5.555 \times 10^{-10} = \frac{(x)(x)}{1.000 \times 10^{-1}}$$ $$x^2 = (1.000 \times 10^{-1})(5.555 \times 10^{-10})$$ $$x^2 = 5.555 \times 10^{-10}$$ $$x = 7.4536 \times 10^{-6}$$ $$[OH^-] = x$$ $$= 7.4536 \times 10^{-6} \text{ mol/L}$$
Calculate pOH using the expression pH = –log [OH⁻].	$$pH = -\log [OH^-]$$ $$= -\log (7.4536 \times 10^{-6})$$ $$= 5.1276$$
Calculate the pH from the pOH.	$$pH = 14 - pOH$$ $$= 14 - 5.1276$$ $$= 8.8724$$ $$= 8.87$$

Check Your Solution
The pH at the equialence point for the titration of a weak acid with a strong base is expected to be greater than 7. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given data. The answer seems reasonable.

110. If 50.00 mL of a 0.100 mol/L nitrous acid solution, HNO_2(aq), with $K_a = 5.6 \times 10^{-4}$, is titrated with 0.100 mol/L potassium hydroxide solution, KOH(aq), determine the resulting pH of the solution at the equivalence point in the titration.

What Is Required?

You need to find the pH at the equivalence point in the titration using nitrous acid and potassium hydroxide solutions.

What Is Given?

You know the volume of the nitrous acid: $V_{HNO_2} = 50.00$ mL

You know the concentration of the nitrous acid: $c_{HNO_2} = 0.100$ mol/L

You know the concentration of the potassium hydroxide: $c_{KOH} = 0.1000$ mol/L

You know the acid dissociation constant: $K_a = 5.6 \times 10^{-4}$

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the reaction between HNO_2(aq) and KOH(aq).	Because KOH is a strong base, the reaction essentially goes to completion. HNO_2(aq) + KOH(aq) → KNO_2(aq) + H_2O(ℓ)
Use the formula $n = cV$ to calculate the amount in moles, n, of HNO_2(aq) present before any base was added.	$n_{HNO_2} = cV$ $= \left(0.100 \dfrac{mol}{L}\right)(50.00 \ mL)\left(\dfrac{1 \ L}{1000 \ mL}\right)$ $= 5.00 \times 10^{-3}$ mol
Use the formula $n = cV$ to calculate the amount in moles, n, of KOH(aq) that reacted to bring the reaction to the equivalence point.	$n_{KOH} = n_{HNO_2}$ $= 5 \times 10^{-3}$ mol
Determine the volume of KOH added.	$V = \dfrac{n}{c}$ $= \dfrac{5 \times 10^{-3} \ mol}{0.100 \ \dfrac{mol}{L}}$ $= 5 \times 10^{-2} \ L \times \dfrac{1000 \ mL}{1 \ L}$ 50.00 mL
Determine the total volume of the solution.	$V = V_{HNO_2} + V_{KOH}$ $= 50.00$ mL $+$ 50.00 mL $= 100.0 \ mL \times \dfrac{1 \ L}{1000 \ mL}$ $= 0.100$ L

Determine the amount of NO_2^- present at the equivalence point.	Because you know that all of the HNO_2 was converted into NO_2^-, you know that the amount of NO_2^- present is equal to the amount of HNO_2 present before the KOH was added. $n_{NO_2^-} = n_{HNO_2}$ $= 5.00 \times 10^{-3}$ mol
Use $c = \dfrac{n}{V}$ to determine the concentration of $NO_2^-(aq)$ in mol/L.	$c = \dfrac{n}{V}$ $= \dfrac{5.00 \times 10^{-3} \text{ mol}}{0.100 \text{ L}}$ $= 5.00 \times 10^{-2} \dfrac{\text{mol}}{\text{L}}$
Set up an ICE table to calculate the $[OH^-]$.	See the ICE table below.

	NO_2^- (aq)	+	$H_2O(\ell)$	$\leftrightharpoons$	HNO_2(aq)	+	OH^- (aq)
	$[NO_2^-]$ (mol/L)				$[HNO_2]$ (mol/L)		$[OH^-]$ (mol/L)
I	5.00×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$5.00 \times 10^{-2} - x$				x		x

Write the expression for the ionization constant.	$K_b = \dfrac{[HNO_2][OH^-]}{[HNO_2^-]}$
Calculate K_b from the known K_b	$K_b = \dfrac{K_w}{K_a}$ $= \dfrac{1 \times 10^{-14}}{5.6 \times 10^{-4}}$ $= 1.7857 \times 10^{-11}$
Test to determine whether the approximation can be applied.	$1000 \times K_b = 1000 \times 1.7857 \times 10^{-11}$ $\quad\quad = 1.7857 \times 10^{-8}$ $1.7857 \times 10^{-8} \ll 1.00 \times 10^{-1}$ The approximation is valid. $5.00 \times 10^{-2} - x \approx 5.00 \times 10^{-2}$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_b = \frac{[HNO_2][OH^-]}{[HNO_2^-]}$$ $$1.7857 \times 10^{-10} = \frac{(x)(x)}{5.00 \times 10^{-2}}$$ $$x^2 = (5.00 \times 10^{-2})(1.7857 \times 10^{-11})$$ $$x^2 = 8.9286 \times 10^{-13}$$ $$x = 9.4491 \times 10^{-7}$$ $$[OH^-] = x$$ $$= 9.4491 \times 10^{-7} \text{ mol/L}$$
Calculate pOH using the expression pH = $-$log [OH$^-$].	pH = $-$log [OH$^-$] $= -$log (9.4491×10^{-7}) $= 6.0246$
Calculate the pH from the pOH.	pH = 14 $-$ pOH = 14 $-$ 6.0246 = 7.9754 = 7.98

Check Your Solution
The pH at the equivalence point for the titration of a weak acid with a strong base is expected to be greater than 7. The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

Calculating the pH at the Equivalence Point
(Student textbook page 543)

111. Determine the pH at the equivalence point when 50.00 mL of a 0.150 mol/L ammonia solution, $NH_3(aq)$, is titrated with 0.150 mol/L hydrochloric acid, $HCl(aq)$.

What Is Required?
You need to find the pH at the equivalence point in a titration of ammonia with hydrochloric acid.

What Is Given?
You know that the reaction reached the equivalence point.

You know the concentration of the hydrochloric acid: $c_{HCl} = 0.150$ mol/L

You know the volume of the ammonia solution: $V_{NH_3} = 50.00$ mL

You know the concentration of the ammonia solution: $c_{NH_3} = 0.150$ mol/L

From Appendix B, you can find that $K_a = 5.6 \times 10^{-10}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HCl is a strong acid, the reaction essentially goes to completion. $$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$
Determine the amount in moles, n, of $HCl(aq)$ that reacted.	At the equivalence point, the amount of $HCl(aq)$ added is equal to the amount of $NH_3(aq)$ initially present in the solution. $$n_{HCl} = n_{i(NH_3)}$$ $$= c_{i(NH_3)}V_{i(NH_3)}$$ $$= (0.150 \ \frac{\text{mol}}{\cancel{L}})(50.00 \ \cancel{mL})\left(\frac{1\cancel{L}}{1000 \ \cancel{mL}}\right)$$ $$= 7.50 \times 10^{-3} \text{ mol}$$
Determine the volume of HCl added.	$$V = \frac{n}{c}$$ $$= \frac{7.50 \times 10^{-3} \ \cancel{\text{mol}}}{0.150 \frac{\cancel{\text{mol}}}{\text{L}}}$$ $$= 0.050 \ \cancel{L} \times \frac{1000 \text{ mL}}{\cancel{L}}$$ $$= 50.00 \text{ mL}$$

Determine the total volume, V, at the equivalence point.	$V_{total} = V_{NH_3} + V_{HCl}$ $= 50.00 \text{ mL} + 50.00 \text{ mL}$ $= 100.0 \text{ mL}$
Determine the amount in moles, n, of NH_4^+ present in the solution at the equivalence point.	Because you know that the amount of HCl added was equal to the amount of NH_3 originally present, and you know that the reaction went to completion, you know that the amount of NH_4^+ present when the equivalence is reached is: $n_{NH_4^+} = n_{i(NH_3)}$ $= 7.50 \times 10^{-3} \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[NH_4^+]$.	$\left[NH_4^+ \right] = \dfrac{n}{V}$ $= \dfrac{7.50 \times 10^{-3} \text{ mol}}{100.0 \text{ mL}} \times \dfrac{1000 \text{ mL}}{L}$ $= 7.50 \times 10^{-2} \text{ mol/L}$
Write the equation for the acid-dissociation of $NH_4^+(aq)$. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table on the following page.

	$NH_4^+(aq)$	+	$H_2O(\ell)$	⇋	$NH_3(aq)$	+	$H_3O^+(aq)$
	$[NH_4^+]$ (mol/L)				$[NH_3]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	7.50×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$7.5 \times 10^{-2} - x$				x		x

Write the expression for the ionization constant.	$K_a = \dfrac{[NH_3][H_3O^+]}{[NH_4^+]}$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 5.6 \times 10^{-10} = 5.6 \times 10^{-7}$ $5.6 \ 10^{-7} << 7.5 \times 10^{-2}$ The approximation is valid. $7.5 \times 10^{-2} - x \approx 7.5 \times 10^{-2}$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$ $$5.6 \times 10^{-10} = \frac{(x)(x)}{7.5 \times 10^{-2}}$$ $$x^2 = (7.5 \times 10^{-2})(5.6 \times 10^{-10})$$ $$x^2 = 4.2 \times 10^{-11}$$ $$x = 6.48 \times 10^{-6}$$ $$[H_3O^+] = x$$ $$= 6.48 \times 10^{-6} \text{ mol/L}$$
Calculate pH using the expression $pH = -\log [H_3O^+]$.	$$pH = -\log [H_3O^+]$$ $$= -\log (6.48 \times 10^{-6})$$ $$= 5.19$$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

112. If 125.00 mL of a 0.100 mol/L hydrazine solution, $N_2H_4(aq)$, $K_a = 7.94 \times 10^{-9}$, is titrated to the equivalence point with 0.100 mol/L nitric acid, $HNO_3(aq)$, what is the pH at this point?

What Is Required?

You need to find the pH at the equivalence point in a titration of hydrazine with nitric acid.

What Is Given?

You know that the reaction reached the equivalence point.

You know the concentration of the nitric acid: $c_{HNO_3} = 0.100$ mol/L

You know the volume of the hydrazine: $V_{H_2N_2} = 125.00$ mL

You know the concentration of the hydrazine: $c_{H_2N_2} = 0.100$ mol/L

You know that, for the conjugate acid of hydrazine, $K_a = 7.94 \times 10^{-9}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because nitric acid is a strong acid, the reaction essentially goes to completion. $$N_2H_4(aq) + HNO_3(aq) \rightarrow N_2H_5^+(aq) + NO_3^-(aq)$$
Determine the amount in moles, n, of $HNO_3(aq)$ that reacted.	At the equivalence point, the amount of $HNO_3(aq)$ added is equal to the amount of $N_2H_4(aq)$ initially present in the solution. $$n_{HNO_3} = n_{i(N_2H_4)}$$ $$= c_{i(N_2H_4)}V_{i(N_2H_4)}$$ $$= (0.100 \; \frac{mol}{\cancel{L}})(125.00 \; \cancel{mL})\left(\frac{1\cancel{L}}{1000 \; \cancel{mL}}\right)$$ $$= 1.25 \times 10^{-2} \, mol$$
Determine the volume of HNO_3 added.	$$V = \frac{n}{c}$$ $$= \frac{1.25 \times 10^{-2} \; \cancel{mol}}{0.100 \frac{\cancel{mol}}{L}}$$ $$= 0.125 \; \cancel{L} \times \frac{1000 \; mL}{\cancel{L}}$$ $$= 125.0 \; mL$$
Determine the total volume, V, at the equivalence point.	$$V_{total} = V_{N_2H_4} + V_{HNO_3}$$ $$= 125.00 \; mL + 125.00 \; mL$$ $$= 250.0 \; mL$$

Determine the amount in moles, n, of $N_2H_5^+$ present in the solution at the equivalence point.	Because you know that the amount of HNO_3 added was equal to the amount of N_2H_4 originally present, and you know that the reaction went to completion, you know that the amount of $N_2H_5^+$ present when the equivalence is reached is: $$n_{N_2H_5^+} = n_{i(N_2H_4)}$$ $$= 1.25 \times 10^{-2} \text{ mol}$$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[N_2H_5^+]$.	$$\left[N_2H_5^+ \right] = \frac{n}{V}$$ $$= \frac{1.250 \times 10^{-2} \text{ mol}}{250.00 \text{ mL}} \times \frac{1000 \text{ mL}}{L}$$ $$= 5.00 \times 10^{-2} \frac{\text{mol}}{L}$$
Write the equation for the acid-dissociation of $N_2H_5^+$ (aq). Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

$N_2H_5^+(aq)$	+	$H_2O(\ell)$	$\leftrightharpoons$	$N_2H_4(aq)$	+	$H_3O^+(aq)$
$[N_2H_5^+]$ (mol/L)				$[N_2H_4]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	5.000×10^{-2}				0	0
C	$-x$				$+x$	$+x$
E	$5.000 \times 10^{-2} - x$				x	x

Write the expression for the ionization constant.	$$K_a = \frac{[N_2H_4][H_3O^+]}{[N_2H_5^+]}$$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 7.94 \times 10^{-9} = 7.94 \times 10^{-6}$ $7.94 \times 10^{-6} \ll 5.0 \ 10^{-2}$ The approximation is valid. $5.000 \times 10^{-2} - x \approx 5.000 \times 10^{-2}$

Substitute the equilibrium values into the equilibrium expression. and solve for x. Use x th find the concentration of the hydronium ion.	$$K_a = \frac{[N_2H_4][H_3O^+]}{[N_2H_5^+]}$$ $$7.94 \times 10^{-9} = \frac{(x)(x)}{5.000 \times 10^{-2}}$$ $$x^2 = \left(5.000 \times 10^{-2}\right)\left(7.94 \times 10^{-9}\right)$$ $$x^2 = 3.9700 \times 10^{-10}$$ $$x = 1.9925 \times 10^{-5}$$ $$[H_3O^+] = x$$ $$= 1.9925 \times 10^{-5} \text{ mol/L}$$
Calculate pH using: pH = $-\log$ [H$_3$O$^+$].	$$\text{pH} = -\log [H_3O^+]$$ $$= -\log (1.9925 \times 10^{-5})$$ $$= 4.701$$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

113. A 25.00 mL sample of a 0.200 mol/L pyridine solution, $C_5H_5N(aq)$, $K_b = 1.7 \times 10^{-9}$, is titrated to the equivalence point using 0.200 mol/L hydrochloric acid, $HCl(aq)$. What is the resulting pH of the solution?

What Is Required?
You need to find the pH at the equivalence point in a titration of pyridine with hydrochloric acid.

What Is Given?
You know that the reaction reached the equivalence point.
You know the concentration of the hydrochloric acid: $c_{HCl} = 0.200$ mol/L
You know the volume of the pyridine solution: $V_{C_5H_5N} = 25.00$ mL
You know the concentration of the pyridine solution: $c_{C_5H_5N} = 0.200$ mol/L
You know that, for pyridine, $K_b = 1.7 \times 10^{-9}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HCl is a strong acid, the reaction essentially goes to completion. $C_5H_5N(aq) + HCl(aq) \rightarrow$ $\quad\quad\quad\quad\quad\quad C_5H_5NH^+(aq) + Cl^-(aq)$
Determine the amount in moles, n, of $HCl(aq)$ that reacted.	At the equivalence point, the amount of $HCl(aq)$ added is equal to the amount of $C_5H_5N(aq)$ initially present in the solution. $n_{HCl} = n_{i(C_5H_5N)}$ $\quad = c_{i(C_5H_5N)}V_{i(C_5H_5N)}$ $\quad = (0.200 \frac{mol}{\cancel{L}})(25.00 \cancel{mL})\left(\frac{1\cancel{L}}{1000 \cancel{mL}}\right)$ $\quad = 5.00 \times 10^{-3}$ mol
Determine the volume of HCl added.	$V = \dfrac{n}{c}$ $\quad = \dfrac{5.00 \times 10^{-3} \cancel{mol}}{0.200 \frac{\cancel{mol}}{L}}$ $\quad = 0.0250 \cancel{L} \times \dfrac{1000 \text{ mL}}{\cancel{L}}$ $\quad = 25.00$ mL

Determine the total volume, V, at the equivalence point.	$$V_{\text{total}} = V_{C_5H_5N} + V_{HCl}$$ $$= 25.00 \text{ mL} + 25.00 \text{ mL}$$ $$= 50.00 \text{ mL}$$
Determine the amount in moles, n, of $C_5H_5NH^+(aq)$ present in the solution at the equivalence point.	Because you know that the amount of HCl added was equal to the amount of $C_5H_5N(aq)$ originally present, and you know that the reaction went to completion, you know that the amount of $C_5H_5NH^+(aq)$ present when the equivalence is reached is: $$n_{C_5H_5NH^+}(aq) = n_{i(C_5H_5N)}$$ $$= 5.00 \times 10^{-3} \text{ mol}$$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[C_5H_5NH^+]$.	$$[C_5H_5NH^+] = \frac{n}{V}$$ $$= \frac{5.00 \times 10^{-3} \text{ mol}}{0.05000 \text{ L}}$$ $$= 0.100 \text{ mol/L}$$
Calculate the K_a for $C_5H_5NH^+(aq)$ using the formula $K_a = \dfrac{K_w}{K_b}$.	$$K_a = \frac{K_w}{K_b}$$ $$= \frac{1 \times 10^{-14}}{1.7 \times 10^{-9}}$$ $$= 5.9 \times 10^{-6}$$
Write the balanced equation for the acid-dissociation of $C_5H_5NH^+$ (aq). Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	$C_5H_5NH^+$ (aq)	+	$H_2O(\ell)$	$\leftrightharpoons$	$C_5H_5N(aq)$	+	$H_3O^+(aq)$
	$[C_5H_5NH^+]$ (mol/L)				$[C_5H_5N]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.100				0		0
C	$-x$				$+x$		$+x$
E	$0.100 - x$				x		x

Write the expression for the ionization constant	$$K_a = \frac{[C_5H_5N][H_3O^+]}{[C_5H_5NH^+]}$$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 5.9 \ 10^{-6} = 5.9 \times 10^{-3}$ $5.9 \times 10^{-3} << 0.100$ The approximation is valid. $0.100 - x \approx 0.100$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_a = \frac{[C_5H_5N][H_3O^+]}{[C_5H_5NH^+]}$$ $$5.9 \times 10^{-6} = \frac{(x)(x)}{0.100}$$ $$x^2 = \left(5.9 \times 10^{-6}\right)(0.100)$$ $$x^2 = 5.9 \times 10^{-7}$$ $$x = 7.68 \times 10^{-4}$$ $$[H_3O^+] = x$$ $$= 7.68 \times 10^{-4} \text{ mol/L}$$
Calculate pH using the expression $pH = -\log [H_3O^+]$.	$$pH = -\log [H_3O^+]$$ $$= -\log (7.68 \times 10^{-4})$$ $$= 3.12$$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_b. The answer seems reasonable.

114. Determine the pH at the equivalence point when 0.104 g of sodium acetate, $NaC_2H_3O_2(s)$, $K_b = 5.6 \times 10^{-10}$, is added to enough water to make 25.00 mL of solution and is titrated with 0.100 mol/L of hydrochloric acid, $HCl(aq)$.

What Is Required?
You need to find the pH at the equivalence point in a titration of sodium acetate with hydrochloric acid.

What Is Given?
You know that the reaction reached the equivalence point.
You know the concentration of the hydrochloric acid: $c_{HCl} = 0.100$ mol/L
You know the volume of the sodium acetate: $V_{NaC_2H_3O_2} = 25.00$ mL
You know the mass of the sodium acetate: $c_{NaC_2H_3O_2} = 0.104$ g
You know that, for acetate, $K_b = 5.6 \times 10^{-10}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HCl is a strong acid, the reaction essentially goes to completion. $NaC_2H_3O_2(aq) + HCl(aq) \rightarrow$ $\qquad HC_2H_3O_2(aq) + Na^+(aq) + Cl^-(aq)$
Use mass and molar mass, M, to calculate the amount in moles, n, of $NaC_2H_3O_2$, using the formula $n = \dfrac{m}{M}$.	$\begin{aligned} n_{NaC_2H_3O_2} &= \frac{m}{M} \\ &= \frac{0.104 \text{ g}}{82.04 \text{ g/mol}} \\ &= 1.2677 \times 10^{-3} \text{ mol} \end{aligned}$
Determine the amount in moles, n, of $HCl(aq)$ that reacted.	At the equivalence point, the amount of $HCl(aq)$ added is equal to the amount of $NH_3(aq)$ initially present in the solution. $\begin{aligned} n_{HCl} &= n_{i(NaC_2H_3O_2)} \\ &= 1.2677 \times 10^{-3} \text{ mol} \end{aligned}$
Determine the volume of HCl added.	$\begin{aligned} V &= \frac{n}{c} \\ &= \frac{1.2677 \times 10^{-3} \text{ mol}}{0.100 \dfrac{\text{mol}}{\text{L}}} \\ &= 1.275 \times 10^2 \text{ L} \times \frac{1000 \text{ mL}}{\text{L}} \\ &= 12.677 \text{ mL} \end{aligned}$

Determine the total volume, V, at the equivalence point.	$\begin{aligned} V_{\text{total}} &= V_{\text{NaC}_3\text{H}_3\text{O}_2} + V_{\text{HCl}} \\ &= 25.00 \text{ mL} + 12.677 \text{ mL} \\ &= 37.677 \text{ mL} \end{aligned}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[HC_2H_3O_2]$. Notice that $HC_2H_3O_2$ is acetic acid, CH_3COOH.	$\begin{aligned} [CH_3COOH] &= \frac{n}{V} \\ &= \frac{1.2677 \times 10^{-3} \text{ mol}}{37.677 \text{ mL}} \times \frac{1000 \text{ mL}}{L} \\ &= 3.3647 \times 10^{-2} \text{ mol/L} \end{aligned}$
Calculate the K_a for $CH_3COOH(aq)$ using the formula $K_a = \dfrac{K_w}{K_b}$.	$\begin{aligned} K_a &= \frac{K_w}{K_b} \\ &= \frac{1 \times 10^{-14}}{5.6 \times 10^{-10}} \\ &= 1.8 \times 10^{-5} \end{aligned}$
Write the balanced chemical equation for the acid-dissociation of $CH_3COOH(aq)$. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	CH_3COOH (aq) +	$H_2O(\ell)$	$\leftrightarrows$	CH_3COO^-(aq) +	H_3O^+(aq)
	$[CH_3COOH]$ (mol/L)			$[CH_3COO^-]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	3.3647×10^{-2}			0	0
C	$-x$			$+x$	$+x$
E	$3.3647 \times 10^{-2} - x$			x	x

Write the expression for the equilibrium constant, K_a.	$K_a = \dfrac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 1.8 \times 10^{-5} = 1.8 \times 10^{-2}$ $1.8 \times 10^{-2} < 3.3647 \times 10^{-2}$ The approximation is valid. $3.3647 \times 10^{-2} - x \approx 3.3647 \times 10^{-2}$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$K_a = \dfrac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$ $1.8 \times 10^{-5} = \dfrac{(x)(x)}{3.3647 \times 10^{-2}}$ $x^2 = (3.3647 \times 10^{-2})(1.8 \times 10^{-5})$ $x^2 = 6.0564 \times 10^{-7}$ $x = 7.7823 \times 10^{-4}$ $[H_3O^+] = x$ $\qquad = 7.7823 \times 10^{-4}$ mol/L
Calculate pH using the expression $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $\qquad = -\log (7.7823 \times 10^{-4})$ $\qquad = 3.1089$ $\qquad = 3.11$

Check Your Solution

The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_b. The answer seems reasonable.

115. Determine the pH at the equivalence point of the titration involving 30.00 mL of a 0.150 mol/L solution of trimethylamine, $(CH_3)_3N(aq)$, and a 0.150 mol/L solution of the strong monoprotic acid perchloric acid, $HClO_4(aq)$. The K_b for trimethylamine is 6.3×10^{-5}.

What Is Required?
You need to find the pH at the equivalence point in a titration of trimethylamine with perchloric acid.

What Is Given?
You know that the reaction reached the equivalence point.

You know the concentration of the perchloric acid: $c_{HClO_4} = 0.150$ mol/L

You know the volume of the trimethylamine: $V_{(CH_3)_3N} = 30.00$ mL

You know the concentration of the trimethylamine: $c_{(CH_3)_3N} = 0.150$ mol/L

You know that, for trimethylamine, $K_b = 6.3 \times 10^{-5}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because $HClO_4$ is a strong acid, the reaction essentially goes to completion. $$(CH_3)_3N(aq) + HClO_4(aq) \rightarrow$$ $$(CH_3)_3NH^+(aq) + ClO_4^-(aq)$$
Determine the amount in moles, n, of $HClO_4(aq)$ that reacted.	At the equivalence point, the amount of $HClO_4(aq)$ added is equal to the amount of $(CH_3)_3N(aq)$ initially present in the solution. $$n_{HClO_4} = n_{i((CH_3)_3N)}$$ $$= c_{i((CH_3)_3N)} V_{i((CH_3)_3N)}$$ $$= (0.150 \ \frac{mol}{\cancel{L}})(30.00 \ \cancel{mL})\left(\frac{1\cancel{L}}{1000 \ \cancel{mL}}\right)$$ $$= 4.50 \times 10^{-3} \ mol$$
Determine the volume of $HClO_4$ added.	$$V = \frac{n}{c}$$ $$= \frac{4.50 \times 10^{-3} \ \cancel{mol}}{0.150 \ \frac{\cancel{mol}}{L}}$$ $$= 0.030 \ \cancel{L} \times \frac{1000 \ mL}{\cancel{L}}$$ $$= 30.00 \ mL$$

Determine the total volume, V, at the equivalence point.	$V_{\text{total}} = V_{(CH_3)_3N} + V_{HClO_4}$ $= 30.00 \text{ mL} + 30.00 \text{ mL}$ $= 60.0 \text{ mL}$
Determine the amount in moles, n, of $(CH_3)_3NH^+$ present in the solution at the equivalence point.	Because you know that the amount of $HClO_4$ added was equal to the amount of $(CH_3)_3N$ originally present, and you know that the reaction went to completion, you know that the amount of $(CH_3)_3NH^+$ present when the equivalence is reached is: $n_{(CH_3)_3NH^+} = n_{i((CH_3)_3N)}$ $= 4.50 \times 10^{-3} \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[(CH_3)_3NH^+]$.	$\left[(CH_3)_3NH^+\right] = \dfrac{n}{V}$ $= \dfrac{4.50 \times 10^{-3} \text{ mol}}{60.0 \text{ mL}} \times \dfrac{1000 \text{ mL}}{L}$ $= 7.50 \times 10^{-2} \text{ mol/L}$
Calculate K_a for the trimethylammonium ion, $(CH_3)_3NH^+$: $K_a = \dfrac{K_w}{K_b}$	$K_a = \dfrac{K_w}{K_b}$ $= \dfrac{1 \times 10^{-14}}{6.3 \times 10^{-5}}$ $= 1.587 \times 10^{-10}$
Write the equation for the acid-dissociation of $(CH_3)_3NH^+(aq)$. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	$(CH_3)_3NH^+(aq)$	+	$H_2O(\ell)$	$\leftrightharpoons$	$(CH_3)_3N(aq)$	+	$H_3O^+(aq)$
	$[(CH_3)_3NH^+]$ (mol/L)				$[(CH_3)_3N]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	7.50×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$7.5 \times 10^{-2} - x$				x		x

Write the expression for the equilibrium constant, K_a.	$$K_a = \frac{[(CH_3)N][H_3O^+]}{[(CH_3)NH^+]}$$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 1.587 \times 10^{-10}$ $\qquad\qquad = 1.587 \times 10^{-7}$ $1.587 \times 10^{-7} << 7.5 \times 10^{-2}$ The approximation is valid. $7.5 \times 10^{-2} - x \approx 7.5 \times 10^{-2}$
Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_a = \frac{[(CH_3)N][H_3O^+]}{[(CH_3)NH^+]}$$ $1.587 \times 10^{-10} = \dfrac{(x)(x)}{7.5 \times 10^{-2}}$ $x^2 = \left(7.5 \times 10^{-2}\right)\left(1.587 \times 10^{-10}\right)$ $x^2 = 1.1905 \times 10^{-11}$ $x = 3.4503 \times 10^{-6}$ $[H_3O^+] = x$ $\qquad\quad = 3.4503 \times 10^{-6} \text{ mol/L}$
Calculate pH using the expression $pH = -\log [H_3O^+]$.	$pH = -\log [H_3O^+]$ $\quad\ = -\log (3.4503 \times 10^{-6})$ $\quad\ = 5.46$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_b. The answer seems reasonable.

116. If 18.00 mL of a 0.040 mol/L solution of methanamine, $CH_3NH_2(aq)$, a weak base with $K_b = 4.6 \times 10^{-4}$, is titrated to the equivalence point using 0.40 mol/L hydrochloric acid, $HCl(aq)$, what is the pH of this solution at the equivalence point?

What Is Required?
You need to find the pH at the equivalence point in a titration of methylamine using hydrochloric acid.

What Is Given?
You know that the reaction reached the equivalence point.
You know the concentration of the hydrochloric acid: $c_{HCl} = 0.40$ mol/L
You know the volume of the methanamine: $V_{CH_3NH_2} = 18.00$ mL
You know the concentration of the methanamine: $c_{CH_3NH_2} = 0.040$ mol/L
You know that, for trimethyl amine, $K_b = 4.6 \times 10^{-4}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HCl is a strong acid, the reaction essentially goes to completion. $$CH_3NH_2(aq) + HCl(aq) \rightarrow$$ $$CH_3NH_3^+(aq) + Cl^-(aq)$$
Determine the amount in moles, n, of HCl (aq) that reacted.	At the equivalence point, the amount of HCl(aq) added is equal to the amount of NH_3(aq) initially present in the solution. $$n_{HCl} = n_{i(CH_3NH_2)}$$ $$= c_{i(CH_3NH_2)} V_{i(CH_3NH_2)}$$ $$= (0.040 \ \frac{mol}{L})(18.00 \ mL)\left(\frac{1 \ L}{1000 \ mL}\right)$$ $$= 7.20 \times 10^{-4} \ mol$$
Determine the volume of HCl added.	$$V = \frac{n}{c}$$ $$= \frac{7.20 \times 10^{-4} \ mol}{0.40 \frac{mol}{L}}$$ $$= 1.80 \times 10^{-3} \ L \times \frac{1000 \ mL}{L}$$ $$= 1.80 \ mL$$

Determine the total volume, V, at the equivalence point.	$V_{total} = V_{CH_3NH_2} + V_{HCl}$ $\quad = 18.00 \text{ mL} + 1.80 \text{ mL}$ $\quad = 19.80 \text{ mL}$
Determine the amount in moles, n, of $CH_3NH_2^+$ present in the solution at the equivalence point.	Because you know that the amount of HCl added was equal to the amount of CH_3NH_2 originally present, and you know that the reaction went to completion, you know that the amount of $CH_3NH_2^+$ present when the equivalence is reached is: $n_{(CH_3)_3NH^+} = n_{i(CH_3NH_2)}$ $\quad = 7.20 \times 10^{-4} \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[CH_3NH_2^+]$.	$\left[CH_3NH_2^+\right] = \dfrac{n}{V}$ $\quad = \dfrac{7.20 \times 10^{-4} \text{ mol}}{19.8 \text{ mL}} \times \dfrac{1000 \text{ mL}}{L}$ $\quad = 3.636 \times 10^{-2} \text{ mol/L}$
Calculate K_a for methylamine: $K_a = \dfrac{K_w}{K_b}$	$K_a = \dfrac{K_w}{K_b}$ $\quad = \dfrac{1 \times 10^{-14}}{4.6 \times 10^{-4}}$ $\quad = 2.1739 \times 10^{-11}$
Write the balanced equation for the acid-dissociation of $CH_3NH_3^+(aq)$. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	$CH_3NH_3^+(aq)$	+	$H_2O(\ell)$	$\leftrightharpoons$	$CH_3NH_2(aq)$	+	$H_3O^+(aq)$
	$[(CH_3)_3NH^+]$ (mol/L)				$[(CH_3)_3N]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	3.636×10^{-3}				0		0
C	$-x$				$+x$		$+x$
E	$3.636 \times 10^{-3} - x$				x		x

Write the expression for the ionization constant.	$K_a = \dfrac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$
Test to determine whether the approximation can be applied.	$\begin{aligned} 1000 \times K_a &= 1000 \times 2.1739 \times 10^{-11} \\ &= 2.1739 \times 10^{-8} \end{aligned}$ $2.1739 \times 10^{-8} \ll 3.636 \times 10^{-3}$ The approximation is valid. $3.636 \times 10^{-3} - x \approx 3.636 \times 10^{-3}$
Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$K_a = \dfrac{[CH_3NH_2][H_3O^+]}{[CH_3NH_3^+]}$ $2.1739 \times 10^{-11} = \dfrac{(x)(x)}{3.636 \times 10^{-3}}$ $x^2 = (3.636 \times 10^{-3})(2.1739 \times 10^{-11})$ $x^2 = 7.9043 \times 10^{-14}$ $x = 2.8115 \times 10^{-7}$ $\begin{aligned} [H_3O^+] &= x \\ &= 2.8115 \times 10^{-7} \text{ mol/L} \end{aligned}$
Calculate pH using the expression $pH = -\log [H_3O^+]$.	$\begin{aligned} pH &= -\log [H_3O^+] \\ &= -\log (2.8115 \times 10^{-7}) \\ &= 6.55 \end{aligned}$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_b. The answer seems reasonable.

117. Calculate the pH at the equivalence point in the titration of 112.5 mL of a 0.075 mol/L aniline solution, $C_6H_5NH_2(aq)$, with a 0.075 mol/L solution of hydrobromic acid, HBr(aq). Aniline has a K_b value of 7.4×10^{-10}.

What Is Required?

You need to find the pH at the equivalence point in a titration of aniline with hydrobromic acid.

What Is Given?

You know that the reaction reached the equivalence point.

You know the concentration of the hydrobromic acid: $c_{HBr} = 0.075$ mol/L

You know the volume of the aniline solution: $V_{C_6H_5NH_2} = 112.5$ mL

You know the concentration of the aniline solution: $c_{C_6H_5NH_2} = 0.075$ mol/L

You know that, for aniline, $K_b = 7.4 \times 10^{-10}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HBr is a strong acid, the reaction essentially goes to completion. $C_6H_5NH_2(aq) + HBr(aq) \rightarrow$ $\qquad\qquad\qquad C_6H_5NH_3{}^+(aq) + Br^-(aq)$
Determine the amount in moles, n, of HBr (aq) that reacted.	At the equivalence point, the amount of HBr(aq) added is equal to the amount of $C_6H_5NH_2(aq)$ initially present in the solution. $n_{HBr} = n_{i(C_6H_5NH_2)}$ $\quad\quad = c_{i(C_6H_5NH_2)} V_{i(C_6H_5NH_2)}$ $\quad\quad = (0.075 \frac{\text{mol}}{\cancel{L}})(112.5 \cancel{mL})\left(\frac{1\cancel{L}}{1000 \cancel{mL}}\right)$ $\quad\quad = 8.4375 \times 10^{-3} \text{ mol}$
Determine the volume of HBr added.	$V = \dfrac{n}{c}$ $\quad = \dfrac{8.4375 \times 10^{-3} \cancel{mol}}{0.075 \dfrac{\cancel{mol}}{L}}$ $\quad = 1.125 \times 10^{-1} \cancel{L} \times \dfrac{1000 \text{ mL}}{\cancel{L}}$ $\quad = 112.5 \text{ mL}$
Determine the total volume, V, at the equivalence point.	$V_{total} = V_{C_6H_5NH_2} + V_{HBr}$ $\quad = 112.5 \text{ mL} + 112.5 \text{ mL}$ $\quad = 225 \text{ mL}$

Determine the amount in moles, n, of $C_6H_5NH_3^+$ present in the solution at the equivalence point.	Because you know that the amount of HBr added was equal to the amount of $C_6H_5NH_2$ originally present, and you know that the reaction went to completion, you know that the amount of $C_6H_5NH_3^+$ present when the equivalence is reached is: $n_{C_6H_5NH_3^+} = n_{i(C_6H_5NH_2)}$ $\qquad = 8.4375 \times 10^{-3}\,mol$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[CH_3NH_2^+]$.	$\left[C_6H_5NH_3^+\right] = \dfrac{n}{V}$ $\qquad = \dfrac{8.4375 \times 10^{-3}\,mol}{225.0\,\cancel{mL}} \times \dfrac{1000\,\cancel{mL}}{L}$ $\qquad = 3.75 \times 10^{-2}\,mol/L$
Calculate K_a for the conjugate acid of aniline: $K_a = \dfrac{K_w}{K_b}$	$K_a = \dfrac{K_w}{K_b}$ $\qquad = \dfrac{1 \times 10^{-14}}{7.4 \times 10^{-10}}$ $\qquad = 1.3514 \times 10^{-5}$
Write the balancee equation for the acid-dissociation of $C_6H_5NH_3^+(aq)$. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table on the following page.

	$C_6H_5NH_3^+$ (aq) $\quad +$	$H_2O(\ell)$ $\quad \leftrightharpoons$	$C_6H_5NH_2$ (aq) $\quad +$	$H_3O^+(aq)$
	$[C_6H_5NH_3^+]$ (mol/L)		$[C_6H_5NH_2]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	3.75×10^{-2}		0	0
C	$-x$		$+x$	$+x$
E	$3.75 \times 10^{-2} - x$		x	x

Write the expression for the equilibrium constant, K_a.	$K_a = \dfrac{[C_6H_5NH_2][H_3O^+]}{[C_6H_5NH_3^+]}$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 1.3514 \times 10^{-5}$ $\qquad\qquad = 1.3514 \times 10^{-2}$ $1.3514 \times 10^{-2} < 3.75 \times 10^{-2}$ The approximation is valid. $3.75 \times 10^{-2} - x \approx 3.75 \times 10^{-2}$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$K_a = \dfrac{[C_6H_5NH_2][H_3O^+]}{[C_6H_5NH_3^+]}$ $1.3514 \times 10^{-5} = \dfrac{(x)(x)}{3.75 \times 10^{-2}}$ $x^2 = (3.75 \times 10^{-2})(1.3514 \times 10^{-5})$ $x^2 = 5.0676 \times 10^{-7}$ $x = 7.3775 \times 10^{-4}$ $[H_3O^+] = x$ $\quad\quad = 7.3775 \times 10^{-4}$ mol/L
Calculate pH using the expression $pH = -\log [H_3O^+]$	$pH = -\log [H_3O^+]$ $\quad\quad = -\log (7.3775 \times 10^{-4})$ $\quad\quad = 3.13$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_b. The answer seems reasonable.

118. Determine the pH at the equivalence point when 100.00 mL of a 0.200 mol/L ammonia solution, $NH_3(aq)$, is titrated with 0.400 mol/L hydrochloric acid solution, $HCl(aq)$.

What Is Required?
You need to find the pH at the equivalence point in a titration of an ammonia solution with hydrochloric acid.

What Is Given?
You know that the reaction reached the equivalence point.
You know the concentration of the hydrochloric acid: $c_{HCl} = 0.400$ mol/L
You know the volume of the y: $V_{NH_3} = 100.00$ mL
You know the concentration of the y: $c_{NH_3} = 0.200$ mol/L
From Appendix B, you can find that, for the ammonium ion, $K_a = 5.6 \times 10^{-10}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HCl is a strong acid, the reaction essentially goes to completion. $$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$
Determine the amount in moles, n, of HCl (aq) that reacted.	At the equivalence point, the amount of $HCl(aq)$ added is equal to the amount of $C_6H_5NH_2(aq)$ initially present in the solution. $n_{HCl} = n_{i(NH_3)}$ $= c_{i(NH_3)}V_{i(NH_3)}$ $= (0.200 \frac{mol}{\cancel{L}})(100.0 \cancel{mL})\left(\frac{1\cancel{L}}{1000 \cancel{mL}}\right)$ $= 0.0200$ mol
Determine the volume of HCl added.	$V = \dfrac{n}{c}$ $= \dfrac{0.0200 \cancel{mol}}{0.400 \dfrac{\cancel{mol}}{L}}$ $= 0.0500 \cancel{L} \times \dfrac{1000 \text{ mL}}{\cancel{L}}$ $= 50$ mL
Determine the total volume, V, at the equivalence point.	$V_{total} = V_{NH_3} + V_{HCl}$ $= 100.0$ mL $+ 50.00$ mL $= 150.00$ mL

Determine the amount in moles, n, of NH_4^+ present in the solution at the equivalence point.	Because you know that the amount of HCl added was equal to the amount of NH_3 originally present, and you know that the reaction went to completion, you know that the amount of NH_4^+ present when the equivalence is reached is: $n_{NH_4^+} = n_{i(NH_3)}$ $= 0.0200$ mol
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[NH_4^+]$.	$\left[NH_4^+ \right] = \dfrac{n}{V}$ $= \dfrac{0.0200 \text{ mol}}{150.00 \text{ mL}} \times \dfrac{1000 \text{ mL}}{L}$ $= 0.13333$ mol/L
Write the balanced chemical equation for the acid-dissociation of NH_4^+(aq). Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	NH_4^+(aq)	+	$H_2O(\ell)$	$\leftrightharpoons$	NH_3(aq)	+	H_3O^+(aq)
	$[NH_4^+]$ (mol/L)				$[NH_3]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	0.133				0		0
C	$-x$				$+x$		$+x$
E	$0.133 - x$				x		x

Write the expression for the ioization constant.	$K_a = \dfrac{[NH_3][H_3O^+]}{[NH_4^+]}$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 5.6 \times 10^{-10}$ $\qquad\qquad = 5.6 \times 10^{-7}$ $5.6 \times 10^{-7} \ll 0.13333$ The approximation is valid. $0.13333 - x \approx 0.13333$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$ $$5.6 \times 10^{-10} = \frac{(x)(x)}{0.13333}$$ $$x^2 = (0.13333)(5.6 \times 10^{-10})$$ $$x^2 = 7.4665 \times 10^{-11}$$ $$x = 8.6409 \times 10^{-6}$$ $$[H_3O^+] = x$$ $$= 8.6409 \times 10^{-6} \text{ mol/L}$$
Calculate pH using $pH = -\log [H_3O^+]$.	$$pH = -\log [H_3O^+]$$ $$= -\log (8.6409 \times 10^{-6})$$ $$= 5.06$$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

119. If 60.00 mL of a 0.100 mol/L hydrazine solution $N_2H_4(aq)$, $K_a = 7.94 \times 10^{-9}$, is titrated to the equivalence point with 0.200 mol/L hydrobromic acid solution, HBr(aq), what is the pH at this point?

What Is Required?
You need to find the pH at the equivalence point in a titration of hydrazine using hydrobromic acid.

What Is Given?
You know that the reaction reached the equivalence point.
You know the concentration of the hydrobromic acid: $c_{HBr} = 0.200$ mol/L
You know the volume of the hydrazine solution: $V_{N_2H_4} = 60.00$ mL
You know the concentration of the hydrazine solution: $c_{N_2H_4} = 0.100$ mol/L
You know that, for the conjugate acid of hydrazine, $K_a = 7.94 \times 10^{-9}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HBr is a strong acid, the reaction essentially goes to completion. $$N_2H_4(aq) + HBr(aq) \rightarrow N_2H_5^+(aq) + Br^-(aq)$$
Determine the amount in moles, n, of HBr (aq) that reacted.	At the equivalence point, the amount of HBr(aq) added is equal to the amount of $N_2H_4(aq)$ initially present in the solution. $$n_{HBr} = n_{i(N_2H_4)}$$ $$= c_{i(N_2H_4)}V_{i(N_2H_4)}$$ $$= \left(0.100 \ \frac{mol}{\cancel{L}}\right)(60.0 \ \cancel{mL})\left(\frac{1\cancel{L}}{1000 \ \cancel{mL}}\right)$$ $$= 6.00 \times 10^{-3} \ mol$$
Determine the volume of HBr added.	$$V = \frac{n}{c}$$ $$= \frac{6.00 \times 10^{-3} \ \cancel{mol}}{0.200 \ \frac{\cancel{mol}}{L}}$$ $$= 0.0300 \ \cancel{L} \times \frac{1000 \ mL}{\cancel{L}}$$ $$= 30 \ mL$$

Determine the total volume, V, at the equivalence point.	$V_{total} = V_{N_2H_4} + V_{HBr}$ $= 60.0 \text{ mL} + 30.00 \text{ mL}$ $= 90.00 \text{ mL}$
Determine the amount in moles, n, of $N_2H_5^+$ present in the solution at the equivalence point.	Because you know that the amount of HBr added was equal to the amount of N_2H_4 originally present, and you know that the reaction went to completion, you know that the amount of $N_2H_5^+$ present when the equivalence is reached is: $n_{N_2H_5^+} = n_{i(N_2H_4)}$ $= 6.00 \times 10^{-3} \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[N_2H_5^+]$.	$\left[N_2H_5^+ \right] = \dfrac{n}{V}$ $= \dfrac{6.00 \times 10^{-3} \text{ mol}}{90.00 \text{ mL}} \times \dfrac{1000 \text{ mL}}{L}$ $= 6.6667 \times 10^{-2} \text{ mol/L}$
Write the balanced chemical equation for the acid-dissociation of $N_2H_5^+(aq)$. Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	$N_2H_5^+$ (aq)	+	$H_2O(\ell)$	⇌	N_2H_4 (aq)	+	H_3O^+(aq)
	$[N_2H_5^+]$ (mol/L)				$[N_2H_4]$ (mol/L)		$[H_3O^+]$ (mol/L)
I	6.6667×10^{-2}				0		0
C	$-x$				$+x$		$+x$
E	$6.6667 \times 10^{-2} - x$				x		x

Write the expression for the ionization constant.	$K_a = \dfrac{[N_2H_4][H_3O^+]}{[N_2H_5^+]}$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 7.94 \times 10^{-9}$ $= 7.94 \times 10^{-6}$ $7.94 \times 10^{-6} \ll 6.6667 \times 10^{-2}$ The approximation is valid. $6.6667 \times 10^{-2} - x \approx 6.6667 \times 10^{-2}$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_a = \frac{[N_2H_4][H_3O^+]}{[N_2H_5^+]}$$ $$7.94 \times 10^{-9} = \frac{(x)(x)}{6.6667 \times 10^{-2}}$$ $$x^2 = (6.6667 \times 10^{-2})(7.94 \times 10^{-9})$$ $$x^2 = 5.2933 \times 10^{-10}$$ $$x = 2.3007 \times 10^{-5}$$ $$[H_3O^+] = x$$ $$= 2.3007 \times 10^{-5} \text{ mol/L}$$
Calculate pH using: pH = $-\log$ [H$_3$O$^+$].	pH = $-\log$ [H$_3$O$^+$] $= -\log (2.3007 \times 10^{-5})$ $= 4.64$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_a. The answer seems reasonable.

120. A 35.00 mL sample of a 0.250 mol/L pyridine solution, $C_2H_5N(aq)$, $K_b = 1.7 \times 10^{-9}$, is titrated to the equivalence point using 0.300 mol/L hydrochloric acid solution, $HCl(aq)$. What is the resulting pH of the solution?

What Is Required?

You need to find the pH at the equivalence point in a titration of pyridine with hydrochloric acid.

What Is Given?

You know that the reaction reached the equivalence point.

You know the concentration of the hydrochloric acid: $c_{HCl} = 0.300$ mol/L

You know the volume of the pyridine solution: $V_{C_2H_5N} = 35.00$ mL

You know the concentration of the pyridine solution: $c_{C_2H_5N} = 0.250$ mol/L

You know that, for the conjugate acid of pyridine, $K_a = 1.7 \times 10^{-9}$

Plan Your Strategy	Act on Your Strategy
Write the balanced chemical equation for the reaction.	Because HCl is a strong acid, the reaction essentially goes to completion. $$C_5H_5N(aq) + HCl(aq) \rightarrow$$ $$C_5H_5NH^+(aq) + Cl^-(aq)$$
Determine the amount in moles, n, of HCl (aq) that reacted.	At the equivalence point, the amount of HCl(aq) added is equal to the amount of C_5H_5N (aq) initially present in the solution. $$n_{HCl} = n_{i(C_5H_5N)}$$ $$= c_{i(C_5H_5N)}V_{i(C_5H_5N)}$$ $$= \left(0.250 \frac{\text{mol}}{\cancel{L}} \right)\left(35.0 \cancel{\text{mL}} \right)\left(\frac{1\cancel{L}}{1000 \cancel{\text{mL}}} \right)$$ $$= 8.75 \times 10^{-3} \text{ mol}$$
Determine the volume of HCl added.	$$V = \frac{n}{c}$$ $$= \frac{8.75 \times 10^{-3} \cancel{\text{mol}}}{0.300 \frac{\cancel{\text{mol}}}{L}}$$ $$= 2.92 \times 10^{-2} \cancel{L} \times \frac{1000 \text{ mL}}{\cancel{L}}$$ $$= 29.2 \text{ mL}$$

Determine the total volume, V, at the equivalence point.	$V_{total} = V_{C_5H_5N} + V_{HCl}$ $= 35.0 \text{ mL} + 29.20 \text{ mL}$ $= 64.20 \text{ mL}$
Determine the amount in moles, n, of $N_2H_5^+$ present in the solution at the equivalence point.	Because you know that the amount of HBr added was equal to the amount of C_5H_5N originally present, and you know that the reaction went to completion, you know that the amount of $C_5H_5NH^+$ present when the equivalence is reached is: $n_{C_5H_5N} = n_{i(C_5H_5NH^+)}$ $= 8.75 \times 10^{-3} \text{ mol}$
Use the formula $c = \dfrac{n}{V}$ to calculate the concentration of $[C_5H_5NH^+]$.	$\left[C_5H_5NH^+\right] = \dfrac{n}{V}$ $= \dfrac{8.75 \times 10^{-3} \text{ mol}}{64.20 \text{ mL}} \times \dfrac{1000 \text{ mL}}{L}$ $= 0.13636 \text{ mol/L}$
Write the balanced equation for the acid-dissociation of $C_5H_5NH^+$(aq). Using the stoichiometric relationships in this balanced equation, set up an ICE table to calculate the $[H_3O^+]$.	See the ICE table below.

	$C_5H_5NH^+$(aq) +	$H_2O(\ell)$	$\leftrightharpoons$	C_5H_5N(aq) +	H_3O^+(aq)
	$[C_5H_5NH^+]$ (mol/L)			$[C_5H_5N]$ (mol/L)	$[H_3O^+]$ (mol/L)
I	0.13636			0	0
C	$-x$			$+x$	$+x$
E	$0.13636 - x$			x	x

Write the expression for the equilibrium contant, K_a.	$K_a = \dfrac{[C_5H_5N][H_3O^+]}{[C_5H_5NH^+]}$
Test to determine whether the approximation can be applied.	$1000 \times K_a = 1000 \times 1.7 \times 10^{-9}$ $= 1.7 \times 10^{-6}$ $1.7 \times 10^{-6} \ll 0.13636$ The approximation is valid. $0.13636 - x \approx 0.13636$

Substitute the equilibrium values into the equilibrium expression and solve for x. Use x to find the concentration of the hydronium ion.	$$K_a = \frac{[C_5H_5N][H_3O^+]}{[C_5H_5NH^+]}$$ $$1.7 \times 10^{-9} = \frac{(x)(x)}{0.13636}$$ $$x^2 = (0.13636)(1.7 \times 10^{-9})$$ $$x^2 = 2.3182 \times 10^{-10}$$ $$x = 1.5226 \times 10^{-5}$$ $$[H_3O^+] = x$$ $$= 1.5226 \times 10^{-5} \text{ mol/L}$$
Calculate pH using pH = $-\log$ [H_3O^-].	$$\text{pH} = -\log [H_3O^+]$$ $$= -\log (1.5226 \times 10^{-5})$$ $$= 4.82$$

Check Your Solution
The number of significant digits to the right of the decimal is the same as the number of significant digits in the given K_b. The answer seems reasonable.

Calculating K_{sp} from Solubility Data
(Student textbook page 549)

121. It has been determined that the molar solubility of silver chloride, AgCl(aq), is 1.3×10^{-5} mol/L. What is the value of the solubility-product constant for this solid?

What Is Required?
You need to calculate the K_{sp} for silver chloride, AgCl(s).

What Is Given?
You know the solubility of silver chloride is 1.3×10^{-5} mol/L.

Plan Your Strategy	Act on Your Strategy
Write a balanced equation for the solubility equilibrium.	$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
Write the expression for the K_{sp}.	$K_{sp} = [Ag^+][Cl^-]$
Determine the concentration of each ion. 1 mol of AgCl ionizes to give 1mol of Ag^+ and 1 mol of Cl^- ion.	$[AgCl] = [Ag^+] = [Cl^-] = 1.3 \times 10^{-5}$ mol/L
Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Ag^+][Cl^-]$ $= (1.3 \times 10^{-5})(1.3 \times 10^{-5})$ $= 1.7 \times 10^{-10}$

Check Your Solution
The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value is expected for a compound of low solubility.

122. Magnesium fluoride, MgF_2(aq) has a molar solubility of 2.7×10^{-3} mol/L. Use this information to determine the K_{sp} value for the solid.

What Is Required?

You need to calculate the K_{sp} for magnesium fluoride, MgF_2.

What Is Given?

You know the solubility of magnesium fluoride is 2.7×10^{-3} mol/L.

Plan Your Strategy	Act on Your Strategy
Write a balanced equation for the solubility equilibrium.	$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$
Write the expression for the K_{sp}.	$K_{sp} = [Mg^{2+}][F^-]^2$
Determine the concentration of each ion. 1 mol of MgF_2(s) ionizes to give 1 mol of Mg^{2+} and 2 mol F^- ion.	$[Mg^{2+}] = [MgF_2] = 2.7 \times 10^{-3}$ mol/L $[F^-] = 2[MgF_2]$ $\qquad = 2 \times 2.7 \times 10^{-3}$ mol/L $\qquad = 5.4 \times 10^{-6}$ mol/L
Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Mg^{2+}][F^-]^2$ $\qquad = (2.7 \times 10^{-5})(5.4 \times 10^{-5})^2$ $\qquad = 7.9 \times 10^{-8}$

Check Your Solution

The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value is expected for a compound of low solubility.

123. Silver sulfide, $Ag_2S(aq)$ has a K_{sp} value that is equal to 5.6×10^{-49}. What is the molar solubility of the solid?

What Is Required?
You need to calculate the molar solubility of silver sulfide, $Ag_2S(s)$.

What Is Given?
You know the K_{sp} for silver sulfide is 5.6×10^{-49}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium.	$Ag_2S(s) \rightleftharpoons 2Ag^+(aq) + S^{2-}(aq)$
Write the solubility product equation. Let x represent the molar solubility of $Ag_2S(s)$.	$K_{sp} = [Ag^+]^2[S^{2-}]$
Summarize the solubility of the $Ag^+(aq)$ and $S^{2-}(aq)$ ions at equilibrium in an ICE table.	See the ICE table below.

	$Ag_2S(s)$	$\rightleftharpoons$	$2Ag^+(aq)$	$+$	$S^{2-}(aq)$
			$[Ag^+]$ (mol/L)		$[S^{2-}]$ (mol/L)
I					
C	$-x$		$+2x$		$+x$
E			$2x$		x

Write the expression for the K_{sp} and solve for x.	$K_{sp} = [Ag^+]^2[S^{2-}]$ $5.6 \times 10^{-49} = (2x)^2(x)$ $= 4x^3$ $x = 5.2 \times 10^{-17}$ mol/L

Check Your Solution
The solubility is low as expected for a compound with the given K_{sp}. The answer correctly shows 2 significant digits.

124. The K_{sp} value for lead(II) bromide, $PbBr_2(aq)$, is 6.6×10^{-6}. What is the solubility of the solid?

What Is Required?
You need to calculate the solubility of lead(II) bromide(s) in g/L.

What Is Given?
You know the K_{sp} for lead(II) bromide is 6.6×10^{-6}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium.	$PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Br^-(aq)$
Write the solubility product equation. Let x represent the molar solubility of $PbBr_2(s)$.	$K_{sp} = [Pb^{2+}][Br^-]^2$
Summarize the solubility of the $Pb^{2+}(aq)$ and $Br^-(aq)$ ions at equilibrium in an ICE table.	See the ICE table below.

	$PbBr_2(s)$	$\rightleftharpoons$	$Pb^{2+}(aq)$	$+$	$2Br^-(aq)$
			$[Pb^{2+}]$ (mol/L)		$[Br^-]$ (mol/L)
I					
C	$-x$		$+x$		$+2x$
E			x		$2x$

Write the expression for the K_{sp} expression and solve for x.	$K_{sp} = [Pb^{2+}][Br^-]^2$
Determine the molar mass, M, of $PbBr_2$ and express the solubility in g/L ($m = nM$).	$6.6 \times 10^{-6} = x(2x)^2$ $\qquad\qquad = 4x^3$ $\qquad x = 1.18 \times 10^{-2}\,mol/L$ $\qquad\quad = 1.18 \times 10^{-2}\,\cancel{mol}/L \times 367.00\,g/\cancel{mol}$ $\qquad\quad = 4.3\,g/L$

Check Your Solution
The solubility is low as expected for a compound with the given K_{sp}. The answer correctly shows 2 significant digits.

125. Which solid can have more mass ionize into 1.00 L of solution: silver chloride, AgCl(s), or copper(I) chloride, CuCl(s)?

What Is Required?
You need to compare the solubility of silver chloride and copper(I) chloride in g/L.

What Is Given?
From Appendix B, the K_{sp} for silver chloride is 1.77×10^{-10} and for copper(I) chloride is 1.72×10^{-9}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium of silver chloride.	$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
Write the solubility product equation. Let x represent the molar solubility of AgCl(s).	$K_{sp} = [Ag^+][Cl^-]$
Write the expression for the K_{sp}. Let x represent the concentrations of Ag^+ and Cl^-. Solve for x. Determine the molar mass, M, of AgCl and express the solubility in g/L ($m = nM$).	$K_{sp} = \left[Ag^+\right]\left[Cl^-\right]$ $1.77 \times 10^{-10} = (x)(x)$ $x^2 = 1.77 \times 10^{-10}$ $x = 1.33 \times 10^{-5}\,mol/L$ $= 1.33 \times 10^{-5}\,mol/L \times 143.32\,g/mol$ $= 1.91 \times 10^{-3}\,g/L$
Repeat these steps for CuCl(s) and determine which will have the greater mass dissolved in 1 L of solution.	$CuCl(s) \rightleftharpoons Cu^+(aq) + Cl^-(aq)$ $K_{sp} = \left[Cu^+\right]\left[Cl^-\right]$ $1.72 \times 10^{-9} = (x)(x)$ $x = 4.15 \times 10^{-5}\,mol/L$ $= 4.15 \times 10^{-5}\,mol/L \times 99.00\,g/mol$ $= 4.11 \times 10^{-3}\,g/L$ More CuCl(s) will dissolve per litre of solution.

Check Your Solution
The solubilities are expressed correctly to 3 significant digits and the answer is reasonable.

126. The solubility of nickel(II) phosphate, $Ni_3(PO_4)_2(aq)$, is 7.8×10^{-5} g/L. Determine the solubility-product constant for this solid.

What Is Required?
You need to calculate the K_{sp} for nickel(II) phosphate, $Ni_3(PO_4)_2(aq)$.

What Is Given?
You know the solubility of $Ni_3(PO_4)_2(aq)$ is 7.8×10^{-5} g/L.

Plan Your Strategy	Act on Your Strategy
Use the molar mass, M, of $Ni_3(PO_4)_2$ and the formula $n = \dfrac{m}{M}$ to express the solubility in mol/L.	$n = \dfrac{m}{M}$ $= \dfrac{7.8 \times 10^{-5} \cancel{g}/L}{366.01 \cancel{g}/mol}$ $= 2.13 \times 10^{-7}$ mol/L
Write a balanced chemical equation for the solubility equilibrium.	$Ni_3(PO_4)_2(s) \rightleftharpoons 3Ni^{2+}(aq) + 2PO_4^{3-}(aq)$
Determine the concentration of each ion. 1 mol of $Ni_3(PO_4)_2$ ionizes to give 3 mol of $Ni^{2+}(aq)$ ions and 2 mol $PO_4^{3-}(aq)$ ions.	$[Ni^{2+}] = 3[Ni_3(PO_4)_2]$ $\quad = 3 \times 2.13 \times 10^{-7}$ mol/L $\quad = 6.39 \times 10^{-7}$ mol/L $[PO_4^{3-}] = 2[Ni_3(PO_4)_2]$ $\quad = 2 \times 2.13 \times 10^{-7}$ mol/L $\quad = 4.26 \times 10^{-7}$ mol/L
Write the expression for the K_{sp}.	$K_{sp} = [Ni^{2+}]^3[PO_4^{3-}]^2$
Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Ni^{2+}]^3[PO_4^{3-}]^2$ $\quad = (6.39 \times 10^{-7})^3(4.26 \times 10^{-7})^2$ $\quad = 4.74 \times 10^{-32}$

Check Your Solution
The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value for K_{sp} is expected for a compound of low solubility.

127. The solubility of strontium fluoride, $SrF_2(aq)$, is 12.2 mg/100 mL. What is K_{sp} for this solid?

What Is Required?
You need to calculate the K_{sp} for strontium fluoride, $SrF_2(s)$.

What Is Given?
You know the solubility of strontium fluoride, $SrF_2(aq)$, is 12.2 mg/100 mL.

Plan Your Strategy	Act on Your Strategy
Use the molar mass, M, of SrF_2 and the formula $n = \dfrac{m}{M}$ to express the solubility in mol/L.	$\dfrac{12.2 \ \cancel{mg} \times \dfrac{1 \ g}{1000 \ \cancel{mg}}}{100 \ \cancel{mL} \times \dfrac{1 \ L}{1000 \ \cancel{mL}}} = 0.122 \ g/L$ $\text{molar solubility} = \dfrac{0.122 \ \cancel{g}/L}{125.62 \ \cancel{g}/mol}$ $= 9.71 \times 10^{-4} \ mol/L$
Write a balanced equation for the solubility equilibrium.	$SrF_2(s) \rightleftharpoons Sr^{2+}(aq) + 2F^-(aq)$
Determine the concentration of each ion. One mol of SrF_2 ionizes to give one mol of $Sr^{2+}(aq)$ ion and two mol $F^-(aq)$ ions.	$[Sr^{2+}] = [SrF_2]$ $= 9.71 \times 10^{-4} \ mol/L$ $[F^-] = 2[SrF_2]$ $= 2 \times 2.71 \times 10^{-4} \ mol/L$ $= 1.94 \times 10^{-3} \ mol/L$
Write the expression for the K_{sp}. Substitute the ion concentrations into the K_{sp} expression and solve.	$K_{sp} = [Sr^{2+}][F^-]^2$ $= (9.71 \times 10^{-4})(1.94 \times 10^{-3})^2$ $= 3.65 \times 10^{-9}$

Check Your Solution
The calculated value of K_{sp} has the correct number of significant digits that agrees with the given data. The small value is expected for a compound of low solubility.

128. Will a precipitate form if 1.00 mL of a 0.100 mol/L silver nitrate solution, $AgNO_3(aq)$, is added to 1.00 L of a 1.00×10^{-5} mol/L solution of sodium chloride, $NaCl(aq)$? Show your calculations.

What Is Required?
You need to determine whether a precipitate will form.

What Is Given?
You know that 1.00 mL of a 0.100 mol/L silver nitrate solution, $AgNO_3(aq)$, is added to 1.00 L of a 1.00×10^{-5} mol/L solution of sodium chloride, $NaCl(aq)$. The only possible precipitate that can form is silver chloride, $AgCl(s)$.
From Appendix B, you know the K_{sp} for silver chloride is 1.77×10^{-10}.

Plan Your Strategy	Act on Your Strategy
Write the balanced equation for the solubility equilibrium.	$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
Use the formula $n = cV$ to determine the amount in moles, n, of $Ag^+(aq)$ and $Cl^-(aq)$ in the original solution.	$n_{Ag^+} = cV$ $= 0.100 \text{ mol/\cancel{L}} \times 0.001 \text{ \cancel{L}}$ $= 1.00 \times 10^{-4} \text{ mol}$ $n_{Cl^-} = cV$ $= 1.00 \times 10^{-5} \text{ mol/\cancel{L}} \times 1.00 \text{ \cancel{L}}$ $= 1.00 \times 10^{-5} \text{ mol}$
Determine the total volume, V, after mixing the two solutions.	$V = 1.00 \text{ L} + 1.00 \text{ mL}$ $= 1.00 \text{ L} + 0.001 \text{ L}$ $= 1.001 \text{ L}$
Use the formula $c = \dfrac{n}{V}$ to calculate the molar concentrations, c, of $Ag^+(aq)$ and $Cl^-(aq)$.	$\left[Ag^+ \right] = \dfrac{n}{V}$ $= \dfrac{1.00 \times 10^{-4} \text{ mol}}{1.001 \text{ L}}$ $= 9.99 \times 10^{-5} \text{ mol/L}$ $\left[Cl^- \right] = \dfrac{n}{V}$ $= \dfrac{1.00 \times 10^{-5} \text{ mol}}{1.001 \text{ L}}$ $= 9.99 \times 10^{-5} \text{ mol/L}$

Write the expression for the K_{sp} for AgCl.	$Q_{sp} = [Ag^+][Cl^-]$
Use the molar concentrations calculated after mixing to determine a trial value Q_{sp}. If $Q_{sp} > K_{sp}$, a precipitate of AgCl(s) forms.	$\begin{aligned} Q_{sp} &= [Ag^+][Cl^-] \\ &= 9.99 \times 10^{-5} \times 9.99 \times 10^{-5} \\ &= 9.98 \times 10^{-10} \end{aligned}$ K_{sp} for AgCl is 1.77×10^{-10}. $Q_{sp} > K_{sp}$ A precipitate will form.

Check Your Solution
The calculations seem reasonable and the correct number of significant digits has been used.