## Answers to Study Questions

1. B
2. B
3. C
4. A
5. D
6. A
7. C
8. B
9. C
10. A

## Answers to Practice Questions

1. B
2. B
3. B
4. D
5. C
6. C
7. C
8. C
9. B
10. D
11. B
12. A
13. D
14. D
15. A
16. D
17. B
18. A
19. A
20. A
21. B
22. A
23. A
24. C
25. D
26. D
27. C
28. A
29. A
30. B

# Chapter 11 – Equilibrium

*Chapter Summary:*
This chapter will focus on equilibrium including equilibrium constants and determining the amount of reactants and products at equilibrium.
Specific topics covered in this chapter are:

- Definition of equilibrium
- Law of Mass Action
- $K_p$ and $K_c$
- Equilibrium calculations
- Reaction quotient ($Q$) of a reaction
- Le Chatlier's Principle

Previous material that is relevant to your understanding of questions in this chapter include:

- Stoichiometry (***Chapter 4***)
- Solutions and Aqueous Reactions, Part 1 (***Chapter 5***)
- Solutions and Aqueous Reactions, Part 2 (***Chapter 9***)

*Where to find this in your textbook:*
The material in this chapter typically aligns to "Fundamentals of Equilibrium Concepts" in your textbook. The name of your chapter(s) may vary.

*Practice exam:*
There are practice exam questions aligned to the material in this chapter. Because there are a limited number of questions on the practice exam, a review of the breadth of the material in this chapter is advised in preparation for your exam.

*How this fits into the big picture:*
The material in this chapter aligns to the Big Idea of Equilibrium (8) as listed on page 12 of this study guide.

## Study Questions (SQ)

**SQ-1.** What is the equilibrium expression for this reaction?

$$Ni(CO)_4(g) \rightleftharpoons Ni(s) + 4CO(g)$$

**(A)** $K_c = \dfrac{[CO]^4}{[Ni(CO)_4]}$ **(B)** $K_c = \dfrac{[Ni][CO]^4}{[Ni(CO)_4]}$ **(C)** $K_c = \dfrac{[Ni(CO)_4]}{[CO]^4}$ **(D)** $K_c = \dfrac{[Ni(CO)_4]}{[Ni][CO]^4}$

***Knowledge Required:*** (1) How to determine an equilibrium expression from a balanced equation.

***Thinking it Through:*** The question asks you determine the equilibrium expression for a given reaction. The answer options are in $K_c$; therefore, you will consider the concentrations of the materials when constructing the equilibrium expression. An equilibrium constant is defined as the concentrations of products raised to a value equal to their stoichiometric coefficient in a balanced chemical equation divided by the concentrations of reactants raised to a value equal to their stoichiometric coefficients in a balanced chemical equation. Because the concentrations of solids and liquids do not change measurably, both are not included in the equilibrium expression. Therefore, when the concentration of gaseous or solution products are raised to their power you have a numerator of $[CO]^4$; likewise for reactants, when the concentration of gaseous or solution reactants are raised to their power you have a denominator of $[Ni(CO)_4]$. Combined, the numerator and denominator are Choice **(A)**.

Choice **(B)** is not correct because it includes a solid in the expression.

Choice **(C)** is not correct because it has the reactants in the numerator and the products in the denominator, which is reversed.

Choice **(D)** is not correct because it includes a solid in the expression and is the inverse of the balanced

chemical equation as written.

***Practice Questions Related to This:*** **PQ-1, PQ-2, PQ-3, PQ-4, PQ-5, PQ-6, PQ-7, PQ-8, PQ-9, PQ-10, and PQ-11**

**SQ-2.** Given the reaction and equilibrium constant:

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g) \qquad K_c = 2.3\times10^{-7}$$

What is the equilibrium constant for this reaction at the same temperature?

$$SO_3(g) \rightleftharpoons SO_2(g) + \tfrac{1}{2}O_2(g) \qquad K_c = ?$$

**(A)** $1.2\times10^{-7}$ **(B)** $4.6\times10^{-7}$ **(C)** $4.8\times10^{-4}$ **(D)** $4.3\times10^{6}$

***Knowledge Required:*** (1) The relationship between a balanced equation and an equilibrium constant.

***Thinking it Through:*** The question asks you to determine an equilibrium constant for a given balanced chemical equation given a constant for a related balanced chemical equation. You note that the chemical equation for which you are to determine an equilibrium constant is halved compared to the chemical equation for which the equilibrium constant is known. You can write the equilibrium expressions then for both equations and compare:

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g) \qquad K = \frac{[SO_2]^2[O_2]}{[SO_3]^2}$$

$$SO_3(g) \rightleftharpoons SO_2(g) + \frac{1}{2}O_2(g) \qquad K' = \frac{[SO_2][O_2]^{1/2}}{[SO_3]} = K^{1/2} = \sqrt{K}$$

When the reaction is halved, the equilibrium constant is raised to the 0.5 power (or the square root). Thus, the equilibrium constant needed is $(2.3\times10^{-7})^{0.5}$ or $4.8\times10^{-4}$, Choice **(C)**.

Choice **(A)** is not correct because it is the known equilibrium constant divided by 2.
Choice **(B)** is not correct because it is the known equilibrium constant multiplied by 2.
Choice **(D)** is not correct because it is the inverse of the known equilibrium constant.

***Practice Questions Related to This:*** **PQ-12**

**SQ-3.** Conceptual

Consider the equilibrium reactions:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad K_1$$

$$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g) \qquad K_2$$

What is the equilibrium constant, $K$, for this reaction?

$$2SO_2(g) + 2CO_2(g) \rightleftharpoons 2SO_3(g) + 2CO(g)$$

**(A)** $K = \left(\frac{K_1}{K_2}\right)$ **(B)** $K = \left(\frac{K_2}{K_1}\right)$ **(C)** $K = K_1 \times K_2$ **(D)** $K = K_1 + \frac{1}{K_2}$

***Knowledge Required:*** (1) How to combine multiple chemical reactions and their respective equilibrium constants to form a new reaction and constant.

***Thinking it Through:*** The question asks you to determine the equilibrium constant for a chemical reaction that results from the combination of two reactions with given equilibrium constants. First, you note that in order to combine the given reactions, the second reaction must be reversed; therefore, the value of the equilibrium constant for the reverse reaction is the inverse of the given $K$ or $K_2^{-1}$.

$$\begin{array}{ll} 2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) & K_1 \\ 2CO_2(g) \rightleftharpoons 2CO(g) + O_2(g) & K_2^{-1} \\ \hline 2SO_2(g) + 2CO_2(g) \rightleftharpoons 2SO_3(g) + 2CO(g) & K_1 \times K_2^{-1} \end{array}$$

The desired reaction is now the sum of the first reaction and the inverse of the second reaction; when reactions are summed, the new equilibrium constant is the product of the two *known* equilibrium constants or $K_1 \times K_2^{-1}$ for the

reaction given. Choice **(A)**.
Choice **(B)** is not correct because it represents the inverse of the desired chemical reaction.
Choice **(C)** is not correct because it does not account for the reverse of the second chemical reaction.
Choice **(D)** is not correct because it is the sum of the two constants; however, when reactions are summed, the respective equilibrium constants are multiplied.

***Practice Questions Related to This:*** **PQ-13** and **PQ-14**

**SQ-4.** Consider the reaction: $X_2(g) + 2Y(g) \rightleftharpoons 2Z(g)$. 12.00 moles of Z are placed in an evacuated 2.00-liter flask. After the reactants and products reach equilibrium, the flask contains 6.00 moles of Y. What is the equilibrium constant, $K$, for the reaction?

**(A)** 0.333 **(B)** 0.667 **(C)** 1.50 **(D)** 3.00

***Knowledge Required:*** (1) How to determine an equilibrium expression from a balanced equation. (2) How to calculate equilibrium amounts from an equilibrium expression.

***Thinking it Through:*** The question asks you to determine the equilibrium constant given a chemical reaction, the amount of the single species present before equilibrium, and the amount of one species after equilibrium is reached. First, you are told that there are 12.0 moles of Z initially, and 6.00 moles of Y at equilibrium. You then also consider that you know how the amounts will change because you begin with only product (so the change for Z will be negative and the change for $X_2$ or Y will be positive) and the amounts of these changes from the stoichiometry of the balanced equation (an increase of $\underline{2}x$ because the coefficient for Y is $\underline{2}$, etc.). This is best summarized in a table:

| | $X_2$ | + | 2Y | $\rightleftharpoons$ | 2Z |
|---|---|---|---|---|---|
| Initial | 0 | | 0 | | 12.00 |
| Change | $+x$ | | $+2x$ | | $-2x$ |
| Equilibrium | $+x$ | | 6.00 | | $12.0 - 2x$ |

Now you can see that you have an equivalency of $2x = 6.00$ or $x = 3.00$. Using this value for x, you can complete all values in the table:

| | $X_2$ | + | 2Y | $\rightleftharpoons$ | 2Z |
|---|---|---|---|---|---|
| Initial | 0 | | 0 | | 12.00 |
| Change | +3.00 | | +2(3.00) | | −2(3.00) |
| Equilibrium | 3.00 | | 6.00 | | 6.00 |

You now have all of the molar amounts of the substances at equilibrium. To obtain the molar concentrations, you divide each value by 2.00 L (given in the problem): $[X_2] = 1.50$ M $[Y] = 3.00$ M $[Z] = 3.00$ M

To determine the value of $K$, you write an equilibrium expression for this reaction using the concentration of the product raised to its coefficient divided by the concentrations of the reactants raised to their coefficients; you can then substitute the equilibrium concentrations and calculate $K$:

$$K = \frac{[Z]^2}{[X_2][Y]^2} = \frac{(3.00\text{ M})^2}{(1.50\text{ M})(3.00\text{ M})^2} = \frac{1}{1.50} = 0.667$$ Choice **(B)**

Choice **(A)** is not correct because it is determined using moles and not molar concentrations.
Choice **(C)** is not correct because it is determined using the inverse of $K$.
Choice **(D)** is not correct because it is determined using moles and the inverse of $K$.

***Practice Questions Related to This:*** **PQ-15** and **PQ-16**

**SQ-5.** Phosgene decomposes into carbon monoxide and elemental chlorine. If the initial concentration of $COCl_2(g)$ is 0.50 M, what is the equilibrium concentration of CO(g)?

$$COCl_2(g) \rightleftharpoons CO(g) + Cl_2(g) \quad K_c = 6.6\times10^{-8}$$

**(A)** $1.8\times10^{-4}$ M **(B)** $9.1\times10^{-5}$ M **(C)** $1.7\times10^{-8}$ M **(D)** $6.6\times10^{-8}$ M

***Knowledge Required:*** (1) How to write an equilibrium expression from a balanced equation. (2) How to use the equilibrium expression to find equilibrium amounts.

***Thinking it Through:*** The question asks you to determine the equilibrium concentration of a product given that only the reactant is present before it is allowed to reach equilibrium. You can assume that the reaction is done at constant volume and therefore concentrations are proportional to amount.

Begin by setting up a table of initial-change-equilibrium concentrations. Initial concentrations are provided in the question. Mole ratios are used to determine the changes in concentration. Equilibrium values are determined by the addition/subtraction of the initial and change values:

| | $COCl_2$ | $\rightleftharpoons$ | $CO$ | $Cl_2$ |
|---|---|---|---|---|
| Initial | 0.50 M | | 0 | 0 |
| Change | $-x$ | | $+x$ | $+x$ |
| Equilibrium | $0.50 - x$ | | $x$ | $x$ |

The equilibrium expression for this reaction is the concentration of products raised to their respective coefficients divided by the concentration of the starting material raised to its respective coefficient:

$$K_c = \frac{[CO][Cl_2]}{[COCl_2]}$$

You then substitute the values obtained into this expression: $6.6\times10^{-8} = \dfrac{(x)(x)}{0.50-x}$

And simplified further to: $6.6\times10^{-8} = \dfrac{x^2}{0.50-x}$

At this point, you may be inclined to continue the algebra to set up the equation to use the quadratic equation. However, you also know you may be able to simplify the process. The value of the equilibrium constant is very small. You know this means that most of the species at equilibrium are reactants with only a very small quantity of products if you began with only reactants (which you did). Therefore, you then note that any change to the initial concentration of $COCl_2$ will be very small, leaving the initial concentration (0.50 M) essentially unchanged or:

| | $COCl_2$ | $\rightleftharpoons$ | $CO$ | $Cl_2$ |
|---|---|---|---|---|
| Initial | 0.50 M | | 0 | 0 |
| Change | $-x$ | | $+x$ | $+x$ |
| Equilibrium | $0.50 - x \approx 0.50$ | | $x$ | $x$ |

Which then makes the equilibrium expression: $6.6\times10^{-8} = \dfrac{(x)(x)}{0.50-x} \approx \dfrac{(x)(x)}{0.50}$

Which can now be solved without the quadratic equation:

$$6.6\times10^{-8} = \frac{(x)(x)}{0.50} = \frac{x^2}{0.50} \qquad \sqrt{(6.6\times10^{-8})(0.50)} = x \qquad x = 1.8\times10^{-4}$$

Of course, you also know there is nothing wrong with continuing from before and setting up the equation and using the quadric equation:

$$6.6\times10^{-8}(0.50-x) = x^2 \qquad x^2 + 6.6\times10^{-8}(x) - 3.3\times10^{-8} = 0$$

The quadratic equation can then be used to determine possible values of $x$: $\pm1.8\times10^{-4}$. Given that the concentrations of the products cannot be negative, $x$ must be equal to $+1.8\times10^{-4}$ and thus the final concentration of CO is $1.8\times10^{-4}$ M, choice **(A)**, which was the same answer you obtained using the approximation.

***Note:*** *You can check your approximation by using the value of x you calculated and using this in the equilibrium expression to calculate $K_c$ – the value will be the same as given if the approximation can be used. You can also approximate in the reverse if the value of $K_c$ is very large and you only begin with product (so the change to the product concentration will be small).*

Choice **(B)** is not correct because it would be the answer if "$(2x)^2$" was the numerator of $K_c$.
Choice **(C)** is not correct because it would be the answer if "$2x$" was the numerator of $K_c$.
Choice **(D)** is not correct because the final concentration of CO is not the value of $K_c$.

***Practice Questions Related to This:*** **PQ-17, PQ-18,** and **PQ-19**

**SQ-6.** BrCl(g) is in equilibrium with $Br_2(g)$ and $Cl_2(g)$ at 25.0 °C:

$$2BrCl(g) \rightleftharpoons Br_2(g) + Cl_2(g) \quad K_p = 0.130$$

Initially, in a closed container at 25.0 °C, BrCl(g) has a partial pressure of 0.400 atm and $Br_2(g)$ and $Cl_2(g)$ each have partial pressures of 0.800 atm. What is the partial pressure of BrCl(g) once the system reaches equilibrium?

**(A)** 0.419 atm **(B)** 0.781 atm **(C)** 1.16 atm **(D)** 1.21 atm

***Knowledge Required:*** (1) How to determine $K_p$ from a balanced equation. (2) How to calculate equilibrium partial pressures from a $K_p$ expression.

***Thinking it Through:*** The question asks you to determine the partial pressure of a species given initial partial pressures of the starting material and products. Because this is a closed system, partial pressures are proportional to amount and therefore changes in partial pressures are proportional to changes in moles of a given species.

To decide which way the reaction proceeds, you calculate a value of $Q$ using the initial pressures given in the question:

$$Q_p = \frac{(P_{Br_2})_0(P_{Cl_2})_0}{(P_{BrCl})_0^2} = \frac{(0.800)(0.800)}{(0.400)^2} = 1.6 \quad (K_p = 0.130)$$

Because $Q$ is greater than $K$, the reaction will go to the left.

You then continue by setting up a table of initial-change-equilibrium (ICE) partial pressures. Initial partial pressures are provided in the question. Your decision earlier that the reaction will proceed to the left is indicated in the ICE table by the negative (–) signs in the **change line** for the *products* and the positive (+) sign for the *reactant*.

Mole ratios are used to determine the changes in partial pressures (so $+\underline{2}x$ for $\underline{2}$BrCl and $-\underline{1}x$ for $\underline{1}Br_2$, etc.). Equilibrium values are determined by the addition/subtraction of the initial and change values.

| | 2BrCl(g) | ⇌ | $Br_2(g)$ | $Cl_2(g)$ |
|---|---|---|---|---|
| Initial | 0.400 atm | | 0.800 atm | 0.800 atm |
| Change | $+2x$ | | $-x$ | $-x$ |
| Equilibrium | $0.400 + 2x$ | | $0.800 - x$ | $0.800 - x$ |

The $K_p$ expression for this reaction is the partial pressures of products raised to their respective coefficients divided by the partial pressure of the reactant raised to its respective coefficient:

$$K_p = \frac{P_{Br_2} P_{Cl_2}}{P_{BrCl}^2}$$

You then substitute the values obtained into this expression:

$$0.130 = \frac{(0.800 - x)(0.800 - x)}{(0.400 + 2x)^2}$$

The expression is then further simplified: (The perfect square is used in the simplification.)

$$0.130 = \frac{(0.800 - x)^2}{(0.400 + 2x)^2} \qquad 0.130 = \left(\frac{0.800 - x}{0.400 + 2x}\right)^2$$

$$\sqrt{0.130} = \left(\frac{0.800 - x}{0.400 + 2x}\right) \qquad 0.361(0.400 + 2x) = 0.800 - x$$

$$0.144 + 0.722x = 0.800 - x \qquad 1.722x = 0.656 \qquad x = 0.381$$

Replacing this value of "$x$" into the table, we determine the partial pressures at equilibrium;

| | 2BrCl(g) | ⇌ | $Br_2(g)$ | $Cl_2(g)$ |
|---|---|---|---|---|
| Initial | 0.400 atm | | 0.800 atm | 0.800 atm |
| Change | + 2(0.381 atm) | | – (0.381 atm) | – (0.381 atm) |
| Equilibrium | 1.16 atm | | 0.419 atm | 0.419 atm |

Thus, the equilibrium partial pressure of BrCl(g) is 1.16 atm, Choice **(C)**.

Choice **(A)** is not correct because it is the partial pressure of either product.

Choice **(B)** is not correct because the answer comes from adding one-half of $x$ to the initial partial pressure of BrCl(g) instead of a full $x$; this is likely confused with subtracting one-half of $x$ from either product partial pressure to determine their equilibrium partial pressures.

Choice **(D)** is not correct because $K_p$ did not include the "2" exponent for the partial pressure of the reactant.

***Practice Questions Related to This:*** **PQ-20** and **PQ-21**

SQ-7. Conceptual

Which will drive the equilibrium to form more Cu(s)?

$$Cu_2O(s) + CO(g) \rightleftharpoons 2Cu(s) + CO_2(g)$$

**(A)** remove CO(g)

**(B)** remove $CO_2$(g)

**(C)** add a catalyst

**(D)** increase the volume of the container

***Knowledge Required:*** (1) Le Chatelier's principle.

***Thinking it Through:*** The question asks you to determine what would influence the equilibrium such that more product is formed. Le Chatelier's principle would suggest that removing a reactant would lead to the formation of more reactants; therefore, choice **(A)** is incorrect. Adding a catalyst, choice **(C)**, would result in the rate of the reaction changing without impacting the position of equilibrium. Le Chatelier's principle would suggest that increasing the volume would favor the side of the reaction with more moles of gas; however, there is the same number of moles of gas for both the reactants and products of this reaction, and choice **(D)**, therefore, is incorrect. Finally, Le Chatelier's principle would suggest that removal of a product would lead to the formation of more product; therefore, choice **(B)** is correct, i.e., "remove $CO_2$(g)" would lead to the formation of more Cu(s).

***Practice Questions Related to This:*** **PQ-22, PQ-23, PQ-24, PQ-25, PQ-26, PQ-27, PQ-28, PQ-29,** and **PQ-30**

## Practice Questions (PQ)

Conceptual **PQ-1.** When the reversible reaction $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ has reached a state of equilibrium,

**(A)** no further reactions occurs.

**(B)** the total moles of products must equal the remaining moles of reactant.

**(C)** the addition of a catalyst will cause formation of more NO.

**(D)** the concentration of each substance in the system will be constant.

**PQ-2.** Xenon tetrafluoride, $XeF_4$, can be prepared by heating Xe and $F_2$ together:

$$Xe(g) + 2F_2(g) \rightleftharpoons XeF_4(g)$$

What is the equilibrium constant expression for this reaction?

**(A)** $K_c = \dfrac{[XeF_4]}{[Xe][F_2]}$ **(B)** $K_c = \dfrac{[XeF_4]}{2[Xe][F_2]}$ **(C)** $K_c = \dfrac{[XeF_4]}{[Xe][F_2]^2}$ **(D)** $K_c = \dfrac{[Xe][F_2]}{[XeF_4]}$

Conceptual **PQ-3.** What is a proper description of chemical equilibrium?

**(A)** The reaction has stopped.

**(B)** The concentrations of products and reactants are equal.

**(C)** The rates of the forward and reverse reactions are zero.

**(D)** The rates of the forward and reverse reactions are the same.

**PQ-4.** What is the $K_c$ for this reaction? $P_4(s) + 6Cl_2(g) \rightleftharpoons 4PCl_3(g)$

**(A)** $K_c = \dfrac{1}{[Cl_2]^6}$ **(B)** $K_c = \dfrac{[PCl_3]^4}{[P_4][Cl_2]^6}$ **(C)** $K_c = \dfrac{[PCl_3]^4}{[Cl_2]^6}$ **(D)** $K_c = \dfrac{[P_4][Cl_2]^6}{[PCl_3]^4}$

**PQ-5.** What is the $K_p$ for this reaction? $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$

**(A)** $K_p = \dfrac{P_{SO_2}^2 P_{O_2}}{P_{SO_3}^2}$ **(B)** $K_p = \dfrac{P_{SO_2} P_{O_2}}{P_{SO_3}}$ **(C)** $K_p = \dfrac{\left(2P_{SO_2}\right)^2 P_{O_2}}{\left(2P_{SO_3}\right)^2}$ **(D)** $K_p = \dfrac{2P_{SO_2}^2 P_{O_2}}{2P_{SO_3}^2}$

**PQ-6.** What is the equilibrium expression for this reaction? $2C(s) + O_2(g) \rightleftharpoons 2CO(g)$

**(A)** $K_c = \dfrac{[CO]}{[C][O_2]}$ **(B)** $K_c = \dfrac{[CO]^2}{[C]^2[O_2]}$ **(C)** $K_c = \dfrac{[2CO]}{[2C][O_2]}$ **(D)** $K_c = \dfrac{[CO]^2}{[O_2]}$

Conceptual **PQ-7.** The value of an equilibrium constant can be used to predict each of these *except* the

**(A)** extent of a reaction.

**(B)** direction of a reaction.

**(C)** time required to reach equilibrium.

**(D)** quantity of reactant(s) remaining at equilibrium.

**PQ-8.** Chemical equilibrium is the result of

**(A)** a stoppage of further reaction.

**(B)** the unavailability of one of the reactants.

**(C)** opposing reactions attaining equal speeds.

**(D)** formation of products equal in mass to the mass of the reactants.

**PQ-9.** The photosynthetic conversion of $CO_2$ and $O_2$ can be represented by the reaction shown. What is the equilibrium expression for this reaction? $6CO_2(g) + 6H_2O(l) \rightleftharpoons C_6H_{12}O_6(s) + 6O_2(g)$

**(A)** $K_c = \dfrac{[CO_2]^6[C_6H_{12}O_6]}{[O_2]^6[H_2O]^6}$ **(B)** $K_c = \dfrac{[CO_2]^6}{[O_2]^6}$

**(C)** $K_c = \dfrac{[O_2]^6[H_2O]^6}{[CO_2]^6[C_6H_{12}O_6]}$ **(D)** $K_c = \dfrac{[O_2]^6}{[CO_2]^6}$

**PQ-10.** Which statement best describes general equilibrium?

**(A)** Equilibrium is reached when the reaction stops.

**(B)** There is only one set of equilibrium concentrations that equals the $K_c$ value.

**(C)** At equilibrium, the rate of the forward reaction is the same as the rate of the reverse reaction.

**(D)** At equilibrium, the total concentration of products equals the total concentration of reactants.

Conceptual **PQ-11.** For this chemical reaction at 25 °C, the concentration of the product is _____ the concentration of reactants at equilibrium. $C_2H_4(g) + H_2(g) \rightleftharpoons C_2H_6(g)$ $K_c = 1.3\times10^{21}$

**(A)** less than

**(B)** equal to

**(C)** greater than

**(D)** is not able to be compared to

**PQ-12.** In which case does $K_c = K_p$?

**(A)** $2A(g) + B(s) \rightleftharpoons 2C(s) + D(g)$

**(B)** $2A(g) + B(s) \rightleftharpoons C(s) + 2D(g)$

**(C)** $3A(g) + B(s) \rightleftharpoons 2C(s) + 2D(g)$

**(D)** $K_c$ and $K_p$ are equivalent in more than one of the above

Conceptual **PQ-13.** Given the equilibrium constants for these reactions:

$$4Cu(s) + O_2(g) \rightleftharpoons 2Cu_2O(s) \qquad K_{c_1}$$

$$2CuO(s) \rightleftharpoons Cu_2O(s) + \tfrac{1}{2}O_2(g) \qquad K_{c_2}$$

What is the value of $K_c$ for this reaction?

$$2Cu(s) + O_2(g) \rightleftharpoons 2CuO(s)$$

**(A)** $\dfrac{\sqrt{K_{c_1}}}{K_{c_2}}$ **(B)** $\dfrac{K_{c_2}}{\sqrt{K_{c_1}}}$ **(C)** $\sqrt{K_{c_1}} \times K_{c_2}$ **(D)** $K_{c_1} \times K_{c_2}$

**PQ-14.** Given the equilibrium constants for these reactions:

$$2CH_4(g) \rightleftharpoons C_2H_6(g) + H_2(g) \qquad K_c = 9.5\times10^{-13}$$

$$CH_4(g) + H_2O(g) \rightleftharpoons CH_3OH(g) + H_2(g) \qquad K_c = 2.8\times10^{-21}$$

What is the value of $K_c$ for this reaction?

$$2CH_3OH(g) + H_2(g) \rightleftharpoons C_2H_6(g) + 2H_2O(g)$$

**(A)** $9.5\times10^{-13}$ **(B)** $2.9\times10^{-9}$ **(C)** $3.4\times10^{8}$ **(D)** $1.2\times10^{29}$

**PQ-15.** Carbon monoxide gas reacts with hydrogen gas at elevated temperatures to form methanol.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

When 0.40 mol of CO and 0.30 mol of $H_2$ are allowed to reach equilibrium in a 1.0 L container, 0.060 mol of $CH_3OH$ are formed. What is the value of $K_c$?

**(A)** 0.50 **(B)** 0.98 **(C)** 1.7 **(D)** 5.4

**PQ-16.** At equilibrium in a 10 L vessel, there are $7.60\times10^{-2}$ moles of $SO_2$, $8.60\times10^{-2}$ moles of $O_2$, and $8.20\times10^{-2}$ moles of $SO_3$. What is the equilibrium constant $K_c$ under these conditions?

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

**(A)** 12.5 **(B)** 13.5 **(C)** 125 **(D)** 135

**PQ-17.** A mixture of 2.0 mol of CO(g) and 2.0 mol of $H_2O(g)$ was allowed to come to equilibrium in a 10.0-L flask at a high temperature. If $K_c = 4.0$, what is the molar concentration of $H_2(g)$ in the equilibrium mixture?

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

**(A)** 0.67 M **(B)** 0.40 M **(C)** 0.20 M **(D)** 0.13 M

**PQ-18.** If a 1.0 L flask is filled with 0.22 mol of $N_2$ and 0.22 mol of $O_2$ at 2000°C, what is [NO] after the reaction establishes equilibrium? ($K_c = 0.10$ at 2000°C)

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

**(A)** 0.034 M **(B)** 0.060 M **(C)** 0.079 M **(D)** 0.12 M

**PQ-19.** For the reaction, $A(g) \rightleftharpoons B(g) + C(g)$,
5 moles of A are allowed to come to equilibrium in a closed rigid container. At equilibrium, how much of A and B are present if 2 moles of C are formed?

**(A)** 0 moles of A and 3 moles of B **(B)** 1 mole of A and 2 moles of B
**(C)** 2 moles of A and 2 moles of B **(D)** 3 moles of A and 2 moles of B

Conceptual **PQ-20.** The equilibrium constant, $K_c = 3.8\times10^{-5}$ for the reaction,

$$I_2(g) \rightleftharpoons 2I(g)$$

What is the state of the system if $[I_2(g)] = 1.0$ M and $[I(g)] = 1.0\times10^{-3}$ M?

**(A)** The system is at equilibrium.
**(B)** The system is shifting towards products.
**(C)** The system is shifting towards reactants.
**(D)** There is insufficient information to describe the status of the system.

Conceptual **PQ-21.** For the reaction: $\frac{1}{2}F_2(g) \rightleftharpoons F(g)$, a reaction mixture initially contains equal amounts of $F_2(g)$ and F(g) in their standard states. If $K_p = 7.55\times10^{-2}$ at this temperature, which statement is true?

**(A)** $Q < K$, and the reaction proceeds towards the reactants.
**(B)** $Q < K$, and the reaction proceeds towards the products.
**(C)** $Q = K$, and the reaction is at equilibrium.
**(D)** $Q > K$, and the reaction proceeds towards the reactants.

Conceptual **PQ-22.** Consider this reaction at equilibrium.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -198 \text{ kJ}$$

Which change would cause an increase in the $SO_3$ / $SO_2$ mole ratio?

**(A)** adding a catalyst **(B)** removing $O_2(g)$
**(C)** decreasing the temperature **(D)** decreasing the pressure

| | $\Delta G_f^\circ$ (kJ·mol$^{-1}$) |
|---|---|
| $H_2O(l)$ | −237 |
| $H_2O(g)$ | −229 |

**Conceptual PQ-23.** At 298 K, the equilibrium constant for this reaction

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightleftharpoons H_2O(l)$$

**(A)** has a value of 1.0 at equilibrium.

**(B)** is larger than the $K_{eq}$ for $H_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons H_2O(g)$.

**(C)** cannot be computed since data on $O_2$ and $H_2$ are not provided.

**(D)** will have the same value as the $K_{eq}$ for $H_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons H_2O(g)$

**Conceptual PQ-24.** Which factors will affect both the position of equilibrium and the value of the equilibrium constant for this reaction? $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92$ kJ

**(A)** increasing the volume of the container

**(B)** adding more nitrogen gas

**(C)** removed ammonia gas

**(D)** lowering the temperature

**Conceptual PQ-25.** Consider this reaction carried out at constant volume.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \quad \Delta H = -198 \text{ kJ}$$

The concentration of $O_2(g)$ at equilibrium increases if

**(A)** $SO_2$ is added to the system.

**(B)** $SO_3$ is added to the system.

**(C)** the temperature of the system is lowered.

**(D)** an inert gas is added to the system.

**Conceptual PQ-26.** Oxygen and ozone are allowed to come to equilibrium in the *exothermic* reaction

$$2O_3(g) \rightleftharpoons 3O_2(g)$$

Which change will *increase* the numerical value of the equilibrium constant, $K$?

**(A)** adding more $O_2(g)$

**(B)** adding a suitable catalyst

**(C)** decreasing the temperature

**(D)** increasing the volume of the container

**Conceptual PQ-27.** This reaction reaches equilibrium in a closed container. What happens if the volume of the container is decreased at constant temperature?

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 2H_2(g)$$

**(A)** No change occurs.

**(B)** The equilibrium constant increases.

**(C)** The reaction shifts towards products.

**(D)** The reaction shifts towards reactants.

**Conceptual PQ-28.** For this equilibrium, what changes will favor formation of products?

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92 \text{ kJ}$$

**(A)** decrease temperature, increase pressure

**(B)** decrease temperature, decrease pressure

**(C)** increase temperature, increase pressure

**(D)** increase temperature, decrease pressure

**Conceptual PQ-29.** The table to the right gives the relative concentration of reactants and products at equilibrium for the generic reaction at two temperatures. The equilibrium constant is

| | X + Z ⇌ XZ | | |
|---|---|---|---|
| 300 K | 0.7 | 0.3 | 0.8 |
| 500 K | 1.0 | 0.6 | 0.5 |

**(A)** negative at 500 K.

**(B)** larger at 500 K than at 300 K.

**(C)** smaller at 500 K than at 300 K.

**(D)** the same at 500 K and at 300 K.

Conceptual PQ-30. A catalyst is added to a system at equilibrium. Which statement is TRUE?

(A) The temperature will decrease.

(B) The equilibrium constant will increase.

(C) The concentration of products will decrease.

(D) If the system is disturbed, it will return to equilibrium faster.


## Answers to Study Questions

1. A
2. C
3. A
4. B
5. A
6. C
7. B

## Answers to Practice Questions

1. D
2. C
3. D
4. C
5. A
6. D
7. C
8. C
9. D
10. C
11. C
12. B
13. A
14. D
15. D
16. D
17. D
18. B
19. D
20. B
21. D
22. C
23. B
24. D
25. B
26. C
27. D
28. A
29. C
30. D

# Chapter 12 – Acids and Bases

*Chapter Summary:*

This chapter will focus on acids and bases including acid-base models, determining an acid dissociation constant, and calculating pH and pOH for a given acid dissociation.

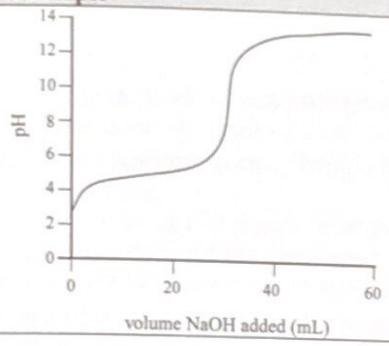
Specific topics covered in this chapter are:

- Acid-base models (Arrhenius, Brønsted-Lowry, Lewis)
- pH and pOH
- Conjugate acid-base pairs
- Titrations and neutralizations
- Buffers

Previous material that is relevant to your understanding of questions in this chapter include:

- Solutions and Aqueous Reactions, Part 1 (***Chapter 5***)
- Solutions and Aqueous Reactions, Part 2 (***Chapter 9***)
- Equilibrium (***Chapter 11***)

*Common representations used in questions related to this material:*

| Name | Example | Used in questions related to |
|---|---|---|
| Titration curves |  | titrations and buffers |

*Where to find this in your textbook:*

The material in this chapter typically aligns to "Acid-Base Equilibria" in your textbook. The name of your chapter(s) may vary.

*Practice exam:*

There are practice exam questions aligned to the material in this chapter. Because there are a limited number of questions on the practice exam, a review of the breadth of the material in this chapter is advised in preparation for your exam.

*How this fits into the big picture:*

The material in this chapter aligns to the Big Idea of Equilibrium (8) as listed on page 12 of this study guide.

## Study Questions (SQ)

**SQ-1.** Which set represents a conjugate acid/base pair?

**(A)** $HClO/HClO_2$ **(B)** $H_2PO_4^-/HPO_4^{2-}$ **(C)** $H_3O^+/OH^-$ **(D)** $NCl_3/NCl_2^-$

***Knowledge Required:*** (1) Definition of conjugate acid/base pairs from the Brønsted-Lowry acid-base model.

***Thinking it Through:*** You are asked in the question to evaluate each answer option to determine which set is a conjugate acid/base pair. Recall from the Brønsted-Lowry acid-base model that conjugate acid/base pairs differ in

one $H^+$. For choice **(A)**, the two species differ by an oxygen atom; these are a set of analogous acids. For choice **(B)**, the two species differ by an $H^+$; therefore, choice **(B)** is correct. For choice **(C)**, the two species differ by *two* $H^+$ ions; while these species are related by $H^+$, the difference in $H^+$ is one too many. For Choice **(D)**, the two species differ by a chlorine atom.

You can also conceptualize this question using the reaction of each species in water to find the conjugate base (and this is good practice):

Choice **(A)**: $\underset{\text{acid}}{HClO(aq)} + H_2O(l) \rightleftharpoons \underset{\text{conjugate base (}\textit{not } HClO_2 \textit{ as given in the problem}\text{)}}{ClO^-(aq)} + H_3O^+(aq)$

Choice **(B)**: $\underset{\text{acid}}{H_2PO_4^-(aq)} + H_2O(l) \rightleftharpoons \underset{\text{conjugate base}}{HPO_4^{2-}(aq)} + H_3O^+(aq)$

Choice **(C)**: $\underset{\text{acid}}{H_3O^+(aq)} + H_2O(l) \rightleftharpoons \underset{\text{conjugate base (}\textit{not } OH^- \textit{ as given in the problem}\text{)}}{H_2O(l)} + H_3O^+(aq)$

Choice **(D)**: $NCl_3$ would not behave as a Brønsted-Lowry acid

***Practice Questions Related to This:*** **PQ-1, PQ-2, PQ-3, PQ-4, PQ-5, PQ-6, PQ-7,** and **PQ-8**

**SQ-2.** What is the pH of a 0.820 M aqueous $NH_3$ solution? $K_b\,(NH_3) = 1.8\times10^{-5}$

**(A)** 2.42 **(B)** 9.25 **(C)** 11.58 **(D)** 13.91

***Knowledge Required:*** (1) How to write a balanced equation for an aqueous acid-base reaction. (2) How to write the base dissociation constant expression ($K_b$) for a balanced chemical equation. (3) How to use the $K_b$ expression to determine equilibrium concentrations. (4) How to convert between pH and pOH.

***Thinking it Through:*** You are asked to determine the pH of a given aqueous solution for a known base and a known $K_b$. This problem has several steps: First, you write a balanced equation for the acid-base reaction; recall that you are told that the reaction is in water, and thus water will act as an acid with the given base:

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

Next, you use the balanced acid-base equation to determine the $K_b$ expression: concentrations of products raised to power of their stoichiometric coefficients divided by the concentrations of reactants raised to power of their stoichiometric coefficients. Note: pure liquids and solids are not included in the $K_b$ expression (omitting water).

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Next, you construct an initial-change-equilibrium (ICE) table. Initial concentrations are provided in the question. Stoichiometric coefficients from the balanced chemical equation are used to determine the changes in concentrations. Equilibrium values are determined by the addition/subtraction of the initial and change values.

| | $NH_3(aq)$ | $\rightleftharpoons$ | $NH_4^+(aq)$ | + | $OH^-(aq)$ |
|---|---|---|---|---|---|
| Initial | 0.820 | | 0 | | 0 |
| Change | $-x$ | | $+x$ | | $+x$ |
| Equilibrium | $0.820 - x$ | | $x$ | | $x$ |

Because acid and base ionization constants are only about 95% accurate, you can compare the ratio of the equilibrium constant to the initial concentration of the acid or base. If the ratio is less than $2.5\times10^{-3}$, you can approximate or: $\dfrac{K_a}{[\text{acid}]_0} < 2.5\times10^{-3}$ $\dfrac{K_b}{[\text{base}]_0} < 2.5\times10^{-3}$.

For this question: $\dfrac{K_b(NH_3)}{[NH_3]_0} = \dfrac{1.8\times10^{-5}}{0.820\text{ M}} = 2.2\times10^{-5}$ which is less than $2.5\times10^{-3}$ (so you can approximate)

Therefore, you can assume that $x$ will be sufficiently small compared to the initial concentration of $NH_3$; therefore, the change to $[NH_3]_0$ is small and the approximation is $[NH_3]_0 = [NH_3]_{\text{equilibrium}}$ or $0.820 - x \approx 0.820$.

You now substitute the equilibrium concentrations into the $K_b$ expression:

$$1.8\times10^{-5}=\frac{x\cdot x}{0.820}$$

The expression is then further simplified and solved:

$$1.5\times10^{-5}=x^2 \qquad x=\pm3.84\times10^{-3}$$

Because $x$ is the final concentration of each of the products and because a concentration cannot be negative, then $x$ is equal to $+3.84\times10^{-3}$ M which is $[NH_4^+]_{equilibrium}$ and $[OH^-]_{equilibrium}$.

Next, you determine the pOH from the concentration of $[OH^-]$:

$$pOH=-\log\left([OH^-]\right)=-\log\left(3.84\times10^{-3}\right)=2.42$$

Finally, you convert from pOH to pH knowing that the sum of pH and pOH is equal to 14 for an aqueous solution.

$pOH + pH = 14$ $pH = 14 - pOH$ $pH = 14 - 2.42 = 11.58$ Which is choice **(C)**

Choice **(A)** is not correct because it is the pOH of the solution.

Choice **(B)** is not correct because it was determined using the $K_b$ value as the concentration of $OH^-$ when determining pOH.

Choice **(D)** is not correct because it was determined using the initial concentration of $NH_3$ as the concentration of $OH^-$ when determining pOH.

***Practice Questions Related to This:*** **PQ-9, PQ-10, PQ-11, PQ-12,** and **PQ-13**

**SQ-3.** Which acid is the weakest in aqueous solution?

Conceptual

**(A)** acetic acid ($K_a = 1.8\times10^{-5}$) **(B)** formic acid ($K_a = 1.8\times10^{-4}$)

**(C)** hydrocyanic acid ($K_a = 6.2\times10^{-10}$) **(D)** nitrous acid ($K_a = 4.5\times10^{-4}$)

***Knowledge Required:*** (1) How to write an acid dissociation equilibrium expression. (2) Relationship between acid dissociation constants ($K_a$) values and acid strength.

***Thinking it Through:*** You are asked to evaluate the given acids and determine which acid is the weakest. You recall that $K_a$ is expressed as the concentration of product(s) raised to the power of their stoichiometric coefficients (from the balanced chemical equation) divided by the concentration of reactant(s) raised to power of their stoichiometric coefficients. Thus, larger values of $K_a$ are associated with a larger product to reactant ratio. Conversely, smaller values of $K_a$ are associated with a smaller product to reactant ratio. For all of the choices, the values of $K_a$ are less than one (so for all of these, starting with only reactants, at equilibrium, [reactants] > [products]). However, you know you can still determine relative acid strength when all $K_a$ values are less than one by determining which value is the largest and how that corresponds to acid strength. Generically:

$$HA(aq)+H_2O(l)\rightleftharpoons A^-(aq)+H_3O^+(aq) \qquad K_a=\frac{\left[A^-\right]\left[H_3O^+\right]}{[HA]}$$

The larger the value of $K_a$, the higher $\left[H_3O^+\right]$, the stronger the acid

Thus, small values of $K_a$ are associated with weak acids with the smallest $K_a$ value associated with the weakest acid. You consider the answer options and note that hydrocyanic acid has the smallest $K_a$ ($6.2\times10^{-10}$) and conclude that it is the weakest of all the acids given as answer options; therefore, choice **(C)** is correct.

***Practice Questions Related to This:*** **PQ-14** and **PQ-15**

**SQ-4.** Which anion is the most basic?

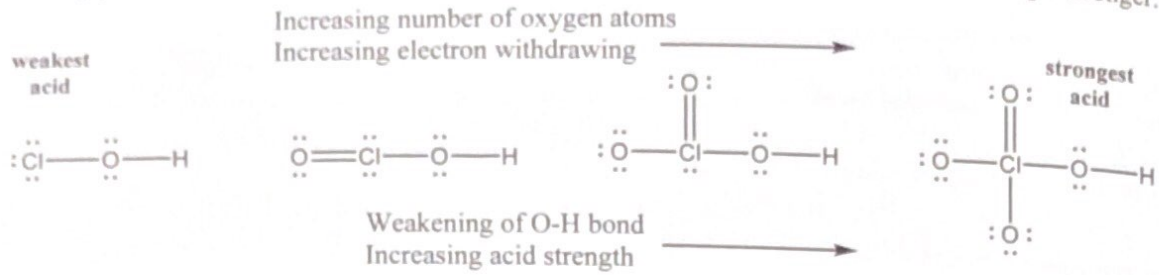
Conceptual

**(A)** $ClO^-$ **(B)** $ClO_2^-$ **(C)** $ClO_3^-$ **(D)** $ClO_4^-$

***Knowledge Required:*** (1) How to determine acidity and basicity from an analogous series of oxoacids.

***Thinking it Through:*** You are asked in the question to evaluate a series of bases to determine which is the most basic. The bases are analogous, considering the answer options from **(A)** to **(D)**, each has one more oxygen than the previous; thus, this a series of oxoacids. The general trend for a series of oxoacids is that as the number of

oxygens increases (or more oxygen atoms in the acid) the acid gets stronger. This is due to the increase in electron withdrawing power of the added oxygens, which causes the O–H bond to weaken and the acid to get stronger.




You can use the relative strength of the acid to determine the relative strengths of the conjugate bases as you know the value of $K_b$ for the conjugate base is related to the $K_a$ value of the acid through: $K_a \times K_b = K_w$.

Because $K_w$ is a constant, at a given temperature, the larger the $K_a$ value (stronger the acid) the smaller the $K_b$ value (weaker the conjugate base). Therefore, the conjugate base of the weakest acid (fewest oxygen atoms) will be the strongest base – choice **(A)**. The next strongest base would be the conjugate base of the next weakest acid – choice **(B)**. The pattern would continue. the ranking of the base strength is:

strongest base → weakest base

$$ClO^- > ClO_2^- > ClO_3^- > ClO_4^-$$

Although not part of this question, when considering the structure of acids and relative strength, you can also be provided with the same number of oxygen atoms, but differing halogens, for example comparing $HClO_4$ and $HBrO_4$. Here, the difference in acid strength is due to the higher electronegativity of chlorine compared to bromine – with $HClO_4$ as the stronger acid.

***Practice Questions Related to This:*** **PQ-16** and **PQ-17**

**SQ-5.** What is $K_b$ of $F^-$? ($K_a$ of HF is $6.8\times10^{-4}$)

**(A)** $6.8\times10^{10}$ **(B)** $1.5\times10^{3}$ **(C)** $6.8\times10^{-4}$ **(D)** $1.5\times10^{-11}$

***Knowledge Required:*** (1) Relationship between $K_a$ and $K_b$ for a conjugate acid-base pair.

***Thinking it Through:*** You are asked in the question to determine the $K_b$ for a species given a known $K_a$ for the conjugate acid of the species. To solve this problem, it is necessary to recall the relationship between $K_a$ and $K_b$ for an acid and its conjugate base dissociated in water:

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq) \qquad K_a$$
$$F^-(aq) + H_2O(l) \rightleftharpoons HF(aq) + OH^-(aq) \qquad K_b$$
$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq) \qquad K_w = K_a \times K_b$$

Therefore: $K_b = \dfrac{K_w}{K_a} = \dfrac{1\times10^{-14}}{6.8\times10^{-4}} = 1.5\times10^{-11}$ which is choice **(D)**.

Choice **(A)** is not correct because it is determined by dividing $K_a$ by $K_w$, which is the inverse of $K_b$.
Choice **(B)** is not correct because it is determined by taking the inverse of $K_a$.
Choice **(C)** is not correct because it is the value of $K_a$.

***Practice Questions Related to This:*** **PQ-18**

**SQ-6.** Which substance will dissolve in water to produce an acidic solution?

Conceptual

**(A)** $FeCl_3$ **(B)** $Na_2O$ **(C)** $NaC_2H_3O_2$ **(D)** $NH_3$

***Knowledge Required:*** (1) Dissociation of ions in aqueous solution. (2) Relative acidity/basicity of common species and ions.

***Thinking it Through:*** You are asked in the question to evaluate a series of compounds to determine which will produce an acidic solution upon dissolution in water. The method you will use to do this is (in turn for each compound):

1. Consider the ions formed when the compound dissolves in water
2. Consider the reaction of each ion with water (hydrolysis)
3. Consider the acidic/basic ions

Choice **(A)** – $FeCl_3$

| | |
|---|---|
| **Step 1** | $FeCl_3(s) \rightarrow Fe^{3+}(aq) + 3Cl^-(aq)$ |
| **Step 2** | $Fe^{3+}(aq) + 6H_2O(l) \rightleftharpoons Fe(H_2O)_6(aq) \rightleftharpoons Fe(H_2O)_5OH^-(aq) + H^+(aq)$<br>$Cl^-(aq) + H_2O(l) \rightleftharpoons$ no reaction (HCl is a strong acid) |
| **Step 3** | $Fe^{3+}$ is an acidic cation and $Cl^-$ is a neutral anion<br>*$FeCl_3$ produces an acidic solution* |

Choice **(B)** – $Na_2O$

| | |
|---|---|
| **Step 1** | $Na_2O(s) \rightarrow 2Na^+(aq) + O^{2-}(aq)$ |
| **Step 2** | $Na^+(aq) + H_2O(l) \rightleftharpoons$ no reaction (NaOH is a strong base)<br>$O^{2-}(aq) + H_2O(l) \rightleftharpoons 2OH^-$ |
| **Step 3** | $Na^+$ is neutral cation and $O^{2-}$ is a basic anion<br>*$Na_2O$ produces a basic solution* |

Choice **(C)** – $NaC_2H_3O_2$

| | |
|---|---|
| **Step 1** | $NaC_2H_3O_2(s) \rightarrow Na^+(aq) + C_2H_3O_2^-(aq)$ |
| **Step 2** | $Na^+(aq) + H_2O(l) \rightleftharpoons$ no reaction (NaOH is a strong base)<br>$C_2H_3O_2^-(aq) + H_2O(l) \rightleftharpoons HC_2H_3O_2(aq) + OH^-(aq)$ |
| **Step 3** | $Na^+$ is neutral cation and $C_2H_3O_2^-$ is a basic anion<br>*$NaC_2H_3O_2$ produces a basic solution* |

Choice **(D)** – $NH_3$ is a molecular compound and not a salt. Instead for this compound, you will consider the reaction of $NH_3$ is water directly (omitting the dissociation into ions):

$NH_3(aq) + H_2O(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ which produces a basic solution.

Choices **(B)**, **(C)**, and **(D)** all lead to the formation of hydroxide anions and basic solutions. Therefore, Choice **(A)** is correct because it leads to the formation of hydronium ion ($H_3O^+$).

***Practice Questions Related to This:*** **PQ-19**, and **PQ-20**

**SQ-7.** An acetate buffer contains equal volumes of 0.35 M $HC_2H_3O_2$ ($pK_a = 4.74$) and 0.55 M $NaC_2H_3O_2$. What is the pH of the buffer?

**(A)** 4.54 **(B)** 4.74 **(C)** 4.94 **(D)** 7.00

***Knowledge Required:*** (1) How to use the Henderson-Hasselbalch equation. (2) How a buffer works.

***Thinking it Through:*** You are asked in the question to determine the pH of a buffer solution given the $pK_a$ of the acid, concentrations of the acid and its conjugate base. The Henderson-Hasselbalch equation is used to relate pH, $pK_a$, and concentrations of conjugate acid-base pairs.

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

You are given the concentration of the acid ($HC_2H_3O_2$, 0.35 M) and the conjugate base ($NaC_2H_3O_2$, 0.55 M) in the question. In addition, you are given the $pK_a$ of the acid ($HC_2H_3O_2$, 4.74). These values are entered into the Henderson-Hasselbalch equation and pH is determined:

$$pH = 4.74 + \log\left(\frac{0.55}{0.35}\right) \qquad pH = 4.74 + \log(1.57)$$

pH = 4.74 + 0.20 = 4.94 which is choice **(C)**

Finally, you can check this answer to make sure it is reasonable. Because you have a higher concentration of conjugate base compared to the acid ($[NaC_2H_3O_2] > [HC_2H_3O_2]$), you would predict the pH would be higher than the $pK_a$ (which it is).

Choice **(A)** is not correct because it is determined using the inverse of the conjugate base to acid molar ratio in the Henderson-Hasselbalch.

Choice **(B)** is not correct because its value is the value of the $pK_a$ for the given acid.

Choice **(D)** is not correct because its value is the value of the pH for a neutral aqueous solution.

***Practice Questions Related to This:*** **PQ-21, PQ-22, PQ-23, PQ-24,** and **PQ-25**

**SQ-8.** What volume (in mL) of 0.150 M NaOH(aq) is required to neutralize 25.0 mL of 0.100 M $H_2SO_4$(aq)?

**(A)** 16.7 mL **(B)** 33.3 mL **(C)** 66.7 mL **(D)** 75.0 mL

***Knowledge Required:*** (1) How to balance the chemical equation for a Brønsted-Lowry acid-base reaction. (2) Solution Stoichiometry.

***Thinking it Through:*** You are asked in the question to determine what volume of a given solution with a known molar concentration is needed to neutralize a given volume of a solution with a known molar concentration. To begin, it is necessary to write a balanced equation for the acid-base reaction:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(l) + Na_2SO_4(aq)$$

You note that from the stoichiometric coefficients in the balanced equation, two moles of NaOH(aq) are necessary to neutralize one mole of $H_2SO_4$.

Next, you recall the relationship between concentration (M), moles ($n_{solute}$), and volume of solution (L):

$$\text{mole}(n_{solute}) = \text{molar concentration} \times \text{volume}$$

Therefore, based on the information in the question, you determine the number of moles of $H_2SO_4$ present.

$$25.0 \text{ mL}\left(\frac{0.100 \text{ mol } H_2SO_4}{\text{L}}\right)\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 2.50\times10^{-3} \text{ mol } H_2SO_4$$

To determine the volume of NaOH needed you use the stoichiometry from the balanced chemical equation:

$$2.50\times10^{-3} \text{ mol } H_2SO_4\left(\frac{2 \text{ mol NaOH}}{1 \text{ mol } H_2SO_4}\right)\left(\frac{1 \text{ L}}{0.150 \text{ mol NaOH}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 33.3 \text{ mL}$$ Choice **(B)**

Choice **(A)** is not correct because it was determined using the inverse mole ratio (1:2) instead of (2:1).

Choice **(C)** is not correct because it was determined using a 1:1 mole ratio or omitted the mole ratio.

Choice **(D)** is not correct because it was determined using the concentration of NaOH with the volume of $H_2SO_4$ and the concentration of $H_2SO_4$ when calculating the volume of NaOH.

***Practice Questions Related to This:*** **PQ-26, PQ-27, PQ-28, PQ-29,** and **PQ-30**

## Practice Questions (PQ)

Conceptual **PQ-1.** Which substance is acting as a Lewis acid? $BF_3(g) + F^-(g) \rightleftharpoons BF_4^-(g)$

**(A)** $BF_3$ **(B)** $F^-$

**(C)** $BF_4^-$ **(D)** This is not a Lewis acid-base reaction.

Conceptual **PQ-2.** What is a true statement about Lewis acids and bases?

**(A)** A Lewis acid must have a nonbonding (lone) pair of electrons.

**(B)** A Lewis base must have a nonbonding (lone) pair of electrons.

**(C)** In a Lewis acid-base reaction, a pair of electrons is donated from the acid to the base.

**(D)** In a Lewis acid-base reaction, one species always goes from having a charge to being electronically neutral.

Conceptual **PQ-3.** What are the Brønsted-Lowry bases in this reaction?

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

**(A)** $NH_3$ and $OH^-$ **(B)** $H_2O$ and $NH_4^+$ **(C)** $NH_3$ and $H_3O$ **(D)** $NH_4^+$ and $OH^-$

**PQ-4.** What is the hydroxide ion concentration in an aqueous solution with a pH of 7.0 at 25 °C?

**(A)** 0.0 M **(B)** $1.0\times10^{-14}$ M **(C)** $1.0\times10^{-7}$ M **(D)** 7.0 M

**PQ-5.** What is the pH of a 0.0050 M solution of $Ba(OH)_2(aq)$ at 25 °C?

**(A)** 2.00 **(B)** 2.30 **(C)** 11.70 **(D)** 12.00

Conceptual **PQ-6.** The formation of a complex ion such as $Cu(NH_3)_4^{2+}(aq)$ can best be categorized as a(n) ___ reaction.

$$Cu^{2+}(aq) + 4NH_3(aq) \rightarrow Cu(NH_3)_4^{2+}(aq)$$

**(A)** Arrhenius acid-base **(B)** Brønsted-Lowry acid-base

**(C)** Lewis acid-base **(D)** oxidation-reduction

**PQ-7.** A 10.0 mL portion of 0.010 M HCl is added to 100.0 mL of water. What is the pH of the resulting solution?

**(A)** between 3.02 and 3.10 **(B)** between 2.90 and 3.01

**(C)** between 2.02 and 2.10 **(D)** between 1.90 and 2.01

**PQ-8.** The pOH of pure water at 40 °C is 6.8. What is the hydronium concentration, $[H_3O^+]$, in pure water at this temperature?

**(A)** $1.0\times10^{-14}$ M **(B)** $6.3\times10^{-8}$ M **(C)** $1.0\times10^{-7}$ M **(D)** $1.6\times10^{-7}$ M

**PQ-9.** What is the pH of a 0.053 M solution of KOH at 25 °C?

**(A)** 0.89 **(B)** 1.28 **(C)** 12.72 **(D)** 13.11

**PQ-10.** What is the equilibrium concentration of nitrous acid, $HNO_2$ ($K_a = 4.5\times10^{-4}$), in a solution that has a pH of 1.65?

**(A)** 0.0032 M **(B)** 0.022 M **(C)** 0.49 M **(D)** 1.1 M

**PQ-11.** The pain reliever codeine is a weak base with a $K_b$ equal to $1.6\times10^{-6}$. What is the pH of a 0.05 M aqueous codeine solution?

**(A)** 7.1 **(B)** 10.5 **(C)** 11.1 **(D)** 12.7

Conceptual **PQ-12.** Besides water molecules, what species is/are present at the greatest concentration when $NH_3(g)$ is bubbled into water? ($K_b$ for $NH_3(aq)$ is $1.8\times10^{-5}$)

**(A)** $NH_3(aq)$ **(B)** $NH_4OH(aq)$

**(C)** $NH_4^+(aq)$ and $OH^-(aq)$ **(D)** $NH_2^-(aq)$ and $H_3O^+(aq)$

Conceptual **PQ-13.** Which 0.10 M solution will have the largest concentration of hydroxide ion?

**(A)** $NH_3$ ($K_b$ of $NH_3 = 1.8\times10^{-5}$) **(B)** NaCN ($K_a$ of $HCN = 4.9\times10^{-10}$)

**(C)** $NaClO_2$ ($K_a$ of $HClO_2 = 1.7\times10^{-4}$) **(D)** $NaHCO_3$ ($K_b$ of $HCO_3^- = 2.3\times10^{-8}$)

Conceptual **PQ-14.** Which is the strongest acid in aqueous solution?

**(A)** acetic acid ($K_a = 1.8\times10^{-5}$) **(B)** benzoic acid ($K_a = 6.3\times10^{-5}$)

**(C)** formic acid ($K_a = 1.7\times10^{-4}$) **(D)** hydrofluoric acid ($K_a = 7.1\times10^{-4}$)

Conceptual **PQ-15.** Which aqueous acid has the largest $K_a$ value?

**(A)** HBrO **(B)** $HBrO_2$ **(C)** HClO **(D)** $HClO_2$

Conceptual **PQ-16.** In non-aqueous solution, the trend in acid strength is observed to be: $K_a(HI) > K_a(HBr) > K_a(HCl)$ Which periodic trend best explains this observed pattern?

**(A)** Atomic radius: as the bond between H and the halogen becomes shorter, acid strength decreases

**(B)** Electron affinity: as the halogen becomes more attracted to electrons, acid strength increases

**(C)** Electronegativity: as the bond between the two atoms becomes less polar, acid strength decreases

**(D)** Ionization energy: as it becomes harder to remove an electron from the halogen, acid strength increases

Conceptual **PQ-17.** The reaction is observed to have a $K_{eq} > 1$.

$$HSO_3^-(aq) + HPO_4^{2-}(aq) \rightleftharpoons SO_3^{2-}(aq) + H_2PO_4^-(aq)$$

What is the strongest acid present in this equilibrium?

**(A)** $H_2PO_4^-$ **(B)** $HPO_4^{2-}$ **(C)** $HSO_3^-$ **(D)** $SO_3^{2-}$

**PQ-18.** What is the pH of a 0.400 M sodium formate ($NaCHO_2$) solution? $K_a(HCHO_2) = 1.8\times10^{-4}$

**(A)** 2.07 **(B)** 5.33 **(C)** 8.67 **(D)** 11.93

Conceptual **PQ-19.** Which salt will form a basic aqueous solution?

**(A)** NaF **(B)** KBr **(C)** LiCl **(D)** $NH_4NO_3$

**PQ-20.** The pH of a 0.050 M aqueous solution of ammonium chloride ($NH_4Cl$) falls within what range?

**(A)** 0 to 2 **(B)** 2 to 7 **(C)** 7 to 12 **(D)** 12 to 14

**PQ-21.** What is the pH of a buffer solution containing equal volumes of 0.11 M $NaCH_3COO$ and 0.090 M. $CH_3COOH$? $K_a(CH_3COOH) = 1.8\times10^{-5}$

**(A)** 2.42 **(B)** 4.83 **(C)** 11.58 **(D)** 13.91

Conceptual **PQ-22.** Which pair of compounds would be the best choice to make a buffer solution with a pH around 5?

**(A)** HF and NaF $K_a = 6.9\times10^{-4}$ **(B)** $C_2H_5CO_2H$ and $C_2H_5CO_2Na$ $K_a = 1.3\times10^{-5}$

**(C)** HClO and NaClO $K_a = 2.8\times10^{-8}$ **(D)** $NH_4Cl$ and $NH_3$ $K_a = 5.6\times10^{-10}$

Conceptual **PQ-23.** Which pair could be used to make a buffer solution?

**(A)** acetic acid and sodium chloride **(B)** acetic acid and sodium hydroxide

**(C)** hydrochloric acid and potassium chloride **(D)** hydrochloric acid and sodium hydroxide

**PQ-24.** What is the pH of a buffer solution made by adding 0.010 mole of solid NaF to 50. mL of 0.40 M HF? Assume no change in volume. $K_a$ (HF) = $6.9\times10^{-4}$

**(A)** 1.6 **(B)** 1.9 **(C)** 2.9 **(D)** 3.2

Conceptual **PQ-25.** An aqueous buffer solution contains only HCN ($pK_a = 9.31$) and KCN and has a pH of 8.50. What can be concluded about the relative concentrations of HCN and KCN in the buffer?

**(A)** [HCN] > [KCN]

**(B)** [HCN] < [KCN]

**(C)** [HCN] = [KCN]

**(D)** nothing can be concluded about the relative concentrations

Conceptual **PQ-26.** Methyl orange is an indicator that changes color from red to yellow-orange over the pH range from 2.9 to 4.5. Methyl orange would be the most appropriate indicator for which type of acid -base titration?

**(A)** A weak acid titrated with a strong base

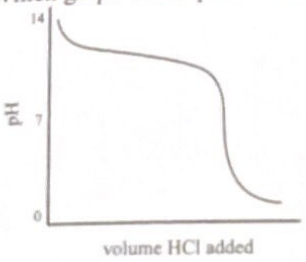
**(B)** A weak base titrated with a strong acid

**(C)** A strong acid titrated with a strong base

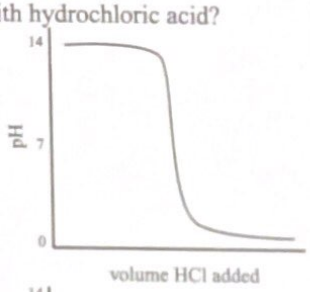
**(D)** Methyl orange would not be an appropriate indicator for any acid-base titration.

Conceptual **PQ-27.** Which graph best represents the titration of ammonia with hydrochloric acid?

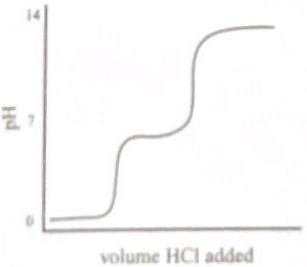
**(A)**



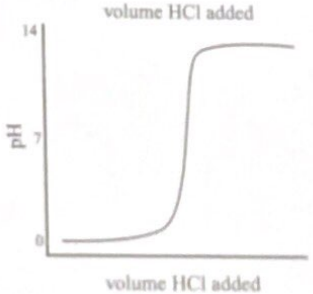

**(B)**




**(C)**




**(D)**



Conceptual

**PQ-28.** In the laboratory, 50 mL of 0.1 M HCl is mixed with 50 mL of 0.1 M NaOH and the solution is stirred gently. At equilibrium, what ionic species, if any, will be present in large amounts in the reaction mixture?

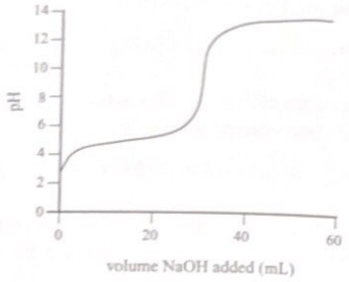
**(A)** $Na^+$ and $Cl^-$ only

**(B)** $H_3O^+$ and $OH^-$ only

**(C)** $H_3O^+$, $OH^-$, $Na^+$, and $Cl^-$

**(D)** No ions will be present

**PQ-29.** What is the approximate $pK_a$ of the weak acid in the titration curve?




**(A)** 2.0 **(B)** 4.8 **(C)** 8.8 **(D)** 11.0

**PQ-30.** What is the pH of a solution that results from mixing 25.0 mL of 0.200 M HA with 12.5 mL of 0.400 M NaOH? ($K_a = 1.0\times10^{-5}$)

**(A)** 2.94 **(B)** 4.94 **(C)** 9.06 **(D)** 11.06

## Answers to Study Questions

1. B
2. C
3. C
4. A
5. D
6. A
7. C
8. B

## Answers to Practice Questions

1. A
2. B
3. A
4. C
5. D
6. C
7. A
8. D
9. C
10. D
11. B
12. A
13. B
14. D
15. D
16. A
17. C
18. C
19. A
20. B
21. B
22. B
23. B
24. C
25. A
26. B
27. A
28. A
29. B
30. C