Chapter 8 Review Problems

INSTRUCTIONS:

You ***do not*** need to write the **question**, ONLY WRITE THE PROBLEM NUMBER and ANSWERS/SOLUTIONS.

- For problems that involve calculations, you must ***show your work*** to get full credit.
- For multiple choice questions, you can simply write the letter (a, b, c, or d) of the correct response.
- Use the *navigation buttons* at the bottom of the pages to get hints, check your answers, move to the next problem, or go back to previous pages.

Chapter Review Problems are **due** at the ***end of class period*** on the dates shown in the CHEM 108 Schedule.

- Late submissions will not be accepted unless the student can prove to the instructor that something outside of their control prevented them from turning in the problem set on the due date (see the course syllabus for more details).

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8.1) When a reaction is at equilibrium __________.

a) it produces the same amount of products as reactants.

b) it occurs very quickly using up all the reactants.

c) it has no products.

d) the rate of the forward reaction is equal to the rate of the reverse reaction.

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8.1) When a reaction is at equilibrium __________.

a) it produces the same amount of products as reactants.

HINT:

b) ~~it occurs very quickly using up all the reactants.~~

c) ~~it has no products.~~

d) the rate of the forward reaction is equal to the rate of the reverse reaction.

For more help: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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8.1) When a reaction is at equilibrium __________.

a) it produces the same amount of products as reactants.

b) it occurs very quickly using up all the reactants.

c) it has no products.

[✓ d) the rate of the forward reaction is equal to the rate of the reverse reaction.]

EXPLANATION: When a chemical reaction reaches equilibrium, the rate of the forward reaction **is equal to** the rate of the reverse reaction. The amount of products and reactants that are present at equilibrium are usually not equal and ***depend on the particular reaction***.

For more details: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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8.2) Chemical equilibrium is defined as the state in which the rate of the forward reaction is equal the rate of the reverse reaction and concentrations of the reactants and products ____________________.

a) decrease

b) are equal

c) do not change

d) increase

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8.2) Chemical equilibrium is defined as the state in which the rate of the forward reaction is equal the rate of the reverse reaction and concentrations of the reactants and products ____________________.

HINT:

a) ~~decrease~~

b) are equal

c) do not change

d) ~~increase~~

For more help: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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a) decrease

b) are equal

[✓ c) do not change]

d) increase

EXPLANATION: When a chemical reaction reaches equilibrium, the forward and reverse reactions are still occurring; however, they are occurring at the same rate and therefore the amounts (concentration) of products and reactants are not changing.

For more details: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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8.3) The chemical equation for the reaction of nitrogen with hydrogen to produce ammonia is shown below.

$$\mathbf{N_2}\,(g) + 3\,\mathbf{H_2}\,(g) \rightleftarrows 2\,\mathbf{NH_3}\,(g)$$

What substances are present in the reaction mixture when equilibrium has been obtained?

a) $\mathbf{N_2}$ (only)

b) $\mathbf{H_2}$ (only)

c) $\mathbf{N_2}$ and $\mathbf{H_2}$ (only)

d) $\mathbf{NH_3}$ (only)

e) $\mathbf{H_2}$, $\mathbf{N_2}$, and $\mathbf{NH_3}$

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e) $\mathbf{H_2}$, $\mathbf{N_2}$, and $\mathbf{NH_3}$

HINT: When a chemical reaction reaches equilibrium, the forward and reverse reactions are occurring at the same rate.

For more help: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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8.3) The chemical equation for the reaction of nitrogen with hydrogen to produce ammonia is shown below.

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d) $\mathbf{NH_3}$ (only)

[✓ e) $\mathbf{H_2}$, $\mathbf{N_2}$, and $\mathbf{NH_3}$]

EXPLANATION: When a chemical reaction reaches equilibrium, the forward and reverse reactions are still occurring; however, they are occurring at the same rate and therefore ***all products and reactants*** are present.

For more details: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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8.4) For any chemical reaction ***at equilibrium***:

$$a\mathbf{A} + b\mathbf{B} \rightleftarrows c\mathbf{C} + d\,\mathbf{D}$$

where *a*, *b*, *c*, and *d* are the *stoichiometric coefficients* for substances A, B, C, and D respectively, the concentrations of reactants and products must satisfy the ***law of mass action***:

$$K_{eq} = \frac{[\mathrm{C}]^c\ [\mathrm{D}]^d}{[\mathrm{A}]^a\ [\mathrm{B}]^b}$$

products

reactants

The ***law of mass action*** is also referred to as the ***equilibrium expression.***

The square brackets, **[ ]**, indicate concentration in **molarity**, for example, "[A]" means "**molarity of substance A**."

Write the ***equilibrium expression*** for the following reaction: $3\ \mathbf{H_2}\,(g) + \mathbf{N_2}\,(g) \rightleftarrows 2\ \mathbf{NH_3}\,(g)$

NOTE: Although it may seem unusual to use the *molarity* concentration of gases, it is not inconsistent with the definition of *molarity*. For gases, this is equal to the number of moles of a particular gas divided by the volume (L) of the container.

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$$K_{eq} = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$$

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NOTE: Although it may seem unusual to use the *molarity* concentration of gases, it is not inconsistent with the definition of *molarity*. For gases, this is equal to the number of moles of a particular gas divided by the volume (L) of the container.

HINT:

The ***equilibrium expression*** is written by multiplying the concentration of the ***products*** (raised to their stoichiometric coefficient powers) in the ***numerator***, and multiplying the concentration of the ***reactants*** (raised to their stoichiometric coefficient powers) in the ***denominator***.

For more help: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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where *a*, *b*, *c*, and *d* are the *stoichiometric coefficients* for substances A, B, C, and D respectively, the concentrations of reactants and products must satisfy the ***law of mass action***:

$$K_{eq} = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b} \begin{array}{l} \leftarrow \textit{products} \\ \leftarrow \textit{reactants} \end{array}$$

The ***law of mass action*** is also referred to as the ***equilibrium expression.***

The square brackets, **[ ]**, indicate concentration in **molarity**, for example, "[A]" means "**molarity of substance A**."

Write the ***equilibrium expression*** for the following reaction: $3\ \mathbf{H_2}\,(g) + \mathbf{N_2}\,(g) \rightleftarrows 2\ \mathbf{NH_3}\,(g)$

NOTE: Although it may seem unusual to use the *molarity* concentration of gases, it is not inconsistent with the definition of *molarity*. For gases, this is equal to the number of moles of a particular gas divided by the volume (L) of the container.

$$K_{eq} = \frac{[NH_3]^2}{[H_2]^3\ [N_2]}$$

EXPLANATION:

The *equilibrium expression* is written by multiplying the concentration of the ***products*** (raised to their stoichiometric coefficient powers) in the ***numerator***, and multiplying the concentration of the ***reactants*** (raised to their stoichiometric coefficient powers) in the ***denominator***.

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8.5) Write the ***equilibrium expression*** for the following reaction:

$$HNO_3\ (\mathit{aq}) + H_2O\ (\mathit{l}) \rightleftarrows H_3O^+\ (\mathit{aq}) + NO_3^-\ (\mathit{aq})$$

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8.5) Write the ***equilibrium expression*** for the following reaction:

$$HNO_3\ (aq) + H_2O\ (l) \rightleftarrows H_3O^+\ (aq) + NO_3^-\ (aq)$$

HINT:

Are the concentrations for liquid substances included in the equilibrium expression?

For more help: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.

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8.5) Write the ***equilibrium expression*** for the following reaction:

$$HNO_3\ (aq) + H_2O\ (l) \rightleftarrows H_3O^+\ (aq) + NO_3^-\ (aq)$$

$$K_{eq} = \frac{[H_3O^+][NO_3^-]}{[HNO_3]}$$

EXPLANATION:

The *equilibrium expression* is written by multiplying the concentration of the ***products*** (raised to their stoichiometric coefficient powers) in the ***numerator***, and multiplying the concentration of the ***reactants*** (raised to their stoichiometric coefficient powers) in the ***denominator***.

$$K_{eq} = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b} \begin{matrix} \leftarrow \textit{products} \\ \leftarrow \textit{reactants} \end{matrix}$$

When **solids** (*s*) or **liquids** (*l*) are present as reactants and/or products, ***they are omitted*** from the equilibrium expression. The only substances that appear in the equilibrium expression are gases (*g*), aqueous (*aq*) solutes, or solutes dissolved in non-aqueous solutions. It is for this reason that $[H_2O]$ does not appear in the equilibrium expression.

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8.6) For the reaction shown below, predict whether the **reactants** or the **products** are ***predominant*** at equilibrium.

$$HCN\ (aq) + H_2O\ (l) \rightleftarrows CN^-\ (aq) + H_3O^+\ (aq) \qquad K_{eq} = 4.9 \times 10^{-10}\ M$$

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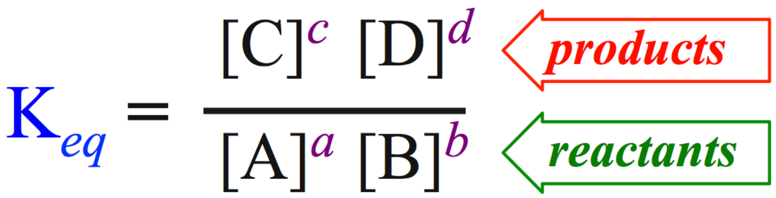
8.6) For the reaction shown below, predict whether the **reactants** or the **products** are ***predominant*** at equilibrium.

$$HCN\ (aq) + H_2O\ (l) \rightleftarrows CN^-\ (aq) + H_3O^+\ (aq) \qquad K_{eq} = 4.9 \times 10^{-10}\ M$$

HINT:

The value of the equilibrium constant allows us to know the relative amounts of products vs. reactants that are present at equilibrium for a particular reaction. The equilibrium expression for K_{eq} is a fraction consisting of the products in the numerator and the reactants in the denominator.



$$K_{eq} = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$$

If we consider the reaction of HCN and water in this problem, we see that the value of K_{eq} is **much less than 1** (4.9×10^{-10} M).

For more help: see chapter 8 part 2 video or chapter 8 section 2 in the textbook.

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8.6) For the reaction shown below, predict whether the **reactants** or the **products** are ***predominant*** at equilibrium.

ANSWER: The **reactants** are ***predominant*** at equilibrium.

$$HCN\ (aq) + H_2O\ (l) \rightleftarrows CN^-\ (aq) + H_3O^+\ (aq) \qquad K_{eq} = 4.9 \times 10^{-10}\ M$$

EXPLANATION:

The value of the equilibrium constant allows us to know the relative amounts of products vs. reactants that are present at equilibrium for a particular reaction. The equilibrium expression for K_{eq} is a fraction consisting of the products in the numerator and the reactants in the denominator.

$$K_{eq} = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$$

products

reactants

Value of K_{eq}	Predominant Species
$K_{eq} >> 1$	Products
$K_{eq} << 1$	Reactants

If K_{eq} is much **greater than 1**, then there are many more product species than reactant species present at equilibrium.

- In this case, we say that the **products** are ***predominant*** at equilibrium.

Conversely, if K_{eq} is much **less than 1**, then there are many more reactant species than product species present at equilibrium.

- In this case, we say that the **reactants** are ***predominant*** at equilibrium.

If we consider the reaction of HCN and water in this problem, we see that the value of K_{eq} is **much less than 1** (4.9×10^{-10} M). This tells us that at equilibrium, the **reactants** are *predominant*; there are many more HCN molecules present than cyanide ions (CN^-) or hydronium (H_3O^+) ions.

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For more details: see chapter 8 part 2 video or chapter 8 section 2 in the textbook.

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8.7) In the reaction below:

$$PCl_3\,(g) \;+\; Cl_2\,(g) \;\rightleftarrows\; PCl_5\,(g)$$

Increasing the concentration of Cl_2, according to Le Chatelier's principle, will ______ in order to re-establish equilibrium.

a) decrease the concentration of PCl_5

b) increase the concentration of PCl_5

c) have no effect

d) increase the concentration of PCl_3

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a) decrease the concentration of PCl_5

b) increase the concentration of PCl_5

c) have no effect

d) increase the concentration of PCl_3

HINT:

Change Made to a Reaction that was at Equilibrium:	**Response:**
Increase the concentration of a ***reactant.***	Rate of the **forward** reaction becomes **greater** than the rate of the reverse reaction until equilibrium is reestablished.
Increase the concentration of a ***product.***	Rate of the **reverse** reaction becomes **greater** than the rate of the forward reaction until equilibrium is reestablished.
Decrease the concentration of a ***reactant.***	Rate of the **forward** reaction becomes **less** than the rate of the reverse reaction until equilibrium is reestablished.
Decrease the concentration of a ***product.***	Rate of the **reverse** reaction becomes **less** than the rate of the forward reaction until equilibrium is reestablished.

How will these responses ***change*** the concentrations of products and reactants?

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Increasing the concentration of Cl_2, according to Le Chatelier's principle, will _____ in order to re-establish equilibrium.

a) decrease the concentration of PCl_5

[✓ b) increase the concentration of PCl_5]

c) have no effect

d) increase the concentration of PCl_3

Change Made to a Reaction that was at Equilibrium:	**Response:**
Increase the concentration of a ***reactant.***	Rate of the **forward** reaction becomes **greater** than the rate of the reverse reaction until equilibrium is reestablished.
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Decrease the concentration of a ***reactant.***	Rate of the **forward** reaction becomes **less** than the rate of the reverse reaction until equilibrium is reestablished.
Decrease the concentration of a ***product.***	Rate of the **reverse** reaction becomes **less** than the rate of the forward reaction until equilibrium is reestablished.

EXPLANATION:

If the concentration of reactant Cl_2 is increased, this causes an increase in the rate of the *forward reaction* because there is now a greater probability of Cl_2 colliding with PCl_3 and then reacting.

Upon the addition of substance Cl_2, PCl_3 and Cl_2 are converted to PCl_5 at a faster forward rate than the reverse rate, causing ***an increase the concentration of* PCl_5**.

This will continue to occur until enough PCl_5 is produced so that the reverse rate is once again equal to the forward rate and ***equilibrium is reestablished***.

For more details: see chapter 8 part 2 video or chapter 8 section 2 in the textbook.

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8.8)

(*i*) Complete the following equation for the **ionization of water** reaction.

$$\mathbf{2}\ H_2O\ (l) \rightleftarrows \boxed{?} + \boxed{?}$$

(*ii*) Write the ***equilibrium expression*** for this reaction.

(*iii*) What is the **value** of the ***equilibrium constant*** ($\mathbf{K_w}$) for this reaction.

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(*i*) Complete the following equation for the **ionization of water** reaction.

$$\mathbf{2}\ H_2O\ (l) \rightleftarrows \boxed{\ ?\ } + \boxed{\ ?\ }$$

(*ii*) Write the ***equilibrium expression*** for this reaction.

HINT:

$$K_{eq} = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$$

products

reactants

H_2O (*l*) does not appear in the equilibrium expression because ***it is a liquid***.

This equilibrium expression is so commonly used, that the symbol "K_w" is used for equilibrium constant (instead of K_{eq}).

(*iii*) What is the **value** of the ***equilibrium constant*** (**K_w**) for this reaction.

HINT: The value of the equilibrium constant (K_w) has been measured experimentally for this reaction.

For more help: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.

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8.8)

(*i*) Complete the following equation for the **ionization of water** reaction.

$$\mathbf{2}\ \mathrm{H_2O}\ (l) \rightleftarrows \mathbf{OH^-}\ (aq) + \mathbf{H_3O^+}\ (aq)$$

(*ii*) Write the ***equilibrium expression*** for this reaction. $\mathbf{K_w = [OH^-][H_3O^+]}$

H_2O (*l*) does not appear in the equilibrium expression because ***it is a liquid***.

This equilibrium expression is so commonly used, that the symbol "K_w" is used for equilibrium constant (instead of K_{eq}).

(*iii*) What is the **value** of the ***equilibrium constant*** (**K_w**) for this reaction. $\mathbf{K_w} = [OH^-][H_3O^+] = \mathbf{1.0 \times 10^{-14}\ M^2}$

The value of the equilibrium constant has been measured experimentally for this reaction.

The unit for K_w is M^2 because we are multiplying two molarity concentrations ($M \bullet M = M^2$).

For more details: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.

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8.9) What is the concentration of $[H_3O^+]$ in an aqueous solution when $[OH^-] = 1.8 \times 10^{-5}$ M?

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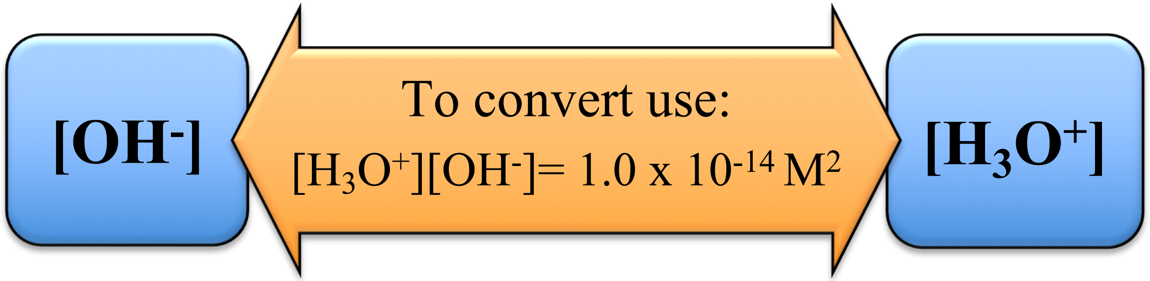
8.9) What is the concentration of $[H_3O^+]$ in an aqueous solution when $[OH^-] = 1.8 \times 10^{-5}$ M? **ANSWER: 5.6 x 10^{-10} M**

We know the equilibrium constant for *the ionization of water*: $\mathbf{K_w} = [OH^-][H_3O^+] = \mathbf{1.0 \times 10^{-14}\ M^2}$

- That means whenever we know $[OH^-]$, we can calculate the concentration of $[H_3O^+]$.
- Likewise, whenever we know $[H_3O^+]$ we can calculate the concentration of $[OH^-]$.



$[OH^-]$

To convert use:

$[H_3O^+][OH^-]= 1.0 \times 10^{-14}\ M^2$

$[H_3O^+]$

For more help: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.

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8.9) What is the concentration of $[H_3O^+]$ in an aqueous solution when $[OH^-] = 1.8 \times 10^{-5}$ M? **ANSWER: 5.6×10^{-10} M**

CLICK HERE to see the **complete solution** for this problem

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We know the equilibrium constant for *the ionization of water*: $\mathbf{K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}\ M^2}$

- That means whenever we know $[OH^-]$, we can calculate the concentration of $[H_3O^+]$.

$$K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}\ M^2$$

$$\boxed{[H_3O^+] = \frac{1.0 \times 10^{-14}\ M^2}{[OH^-]}}$$

- Likewise, whenever we know $[H_3O^+]$ we can calculate the concentration of $[OH^-]$.

$$K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}\ M^2$$

$$\boxed{[OH^-] = \frac{1.0 \times 10^{-14}\ M^2}{[H_3O^+]}}$$

In this problem, we know $[OH^-] = 1.8 \times 10^{-5}$ M, we can calculate the concentration of $[H_3O^+]$:

$$[H_3O^+] = \frac{1.0 \times 10^{-14}\ M^2}{[OH^-]} = \frac{1.0 \times 10^{-14}\ M^2}{1.8 \times 10^{-5}\ M} = \mathbf{5.6 \times 10^{-10}\ M}$$

For more details: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.

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8.10) What is the concentration of $[OH^-]$ in an aqueous solution when $[H_3O^+] = 9.3 \times 10^{-11}$ M?

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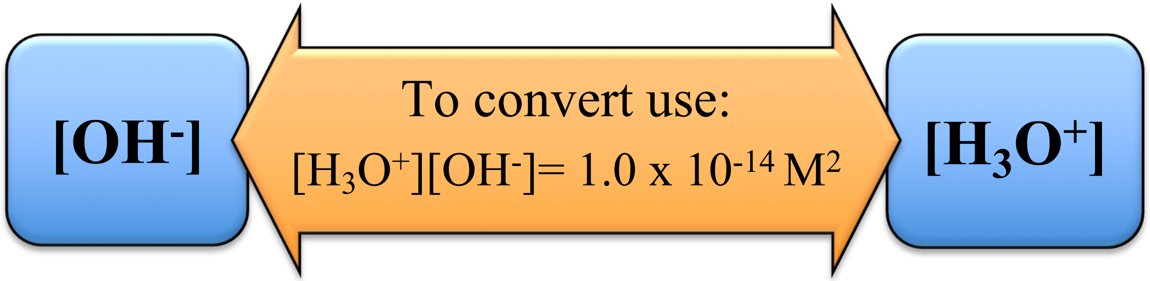
8.10) What is the concentration of $[OH^-]$ in an aqueous solution when $[H_3O^+] = 9.3 \times 10^{-11}$ M? **ANSWER: 1.1×10^{-4} M**

We know the equilibrium constant for *the ionization of water*: $\mathbf{K_w} = [OH^-][H_3O^+] = \mathbf{1.0 \times 10^{-14}\ M^2}$

- That means whenever we know $[OH^-]$, we can calculate the concentration of $[H_3O^+]$.
- Likewise, whenever we know $[H_3O^+]$ we can calculate the concentration of $[OH^-]$.



$[OH^-]$

To convert use:

$[H_3O^+][OH^-] = 1.0 \times 10^{-14}\ M^2$

$[H_3O^+]$

For more help: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.

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8.10) What is the concentration of $[OH^-]$ in an aqueous solution when $[H_3O^+] = 9.3 \times 10^{-11}$ M? **ANSWER: 1.1×10^{-4} M**

CLICK HERE to see the **complete solution** for this problem

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8.10) What is the concentration of $[OH^-]$ in an aqueous solution when $[H_3O^+] = 9.3 \times 10^{-11}$ M? **ANSWER: 1.1×10^{-4} M**

We know the equilibrium constant for *the ionization of water*: $\mathbf{K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}\ M^2}$

- Whenever we know $[H_3O^+]$ we can calculate the concentration of $[OH^-]$.

$$K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14}\ M^2$$

$$\boxed{[OH^-] = \frac{1.0 \times 10^{-14}\ M^2}{[H_3O^+]}}$$

In this problem, we know $[H_3O^+] = 9.3 \times 10^{-11}$ M, we can calculate the concentration of $[OH^-]$:

$$[H_3O^+] = \frac{1.0 \times 10^{-14}\ M^2}{[H_3O^+]} = \frac{1.0 \times 10^{-14}\ M^2}{9.3 \times 10^{-11}\ M} =$$

For more details: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.

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8.11) A compound can be classified as an "**acid**" or a "**base**" depending on its ability to *gain or lose a hydrogen ion* (H^+) in a chemical reaction.

Determine which ***reactant*** is acting as the **acid** and which reactant is acting the **base** in each of the following reactions.

a) $HPO_4^{-}\ (aq) + HNO_3\ (aq) \rightleftarrows H_2PO_4\ (aq) + NO_3^{-}\ (aq)$

b) $H_2CO_3\ (aq) + H_2O\ (l) \rightleftarrows HCO_3^{-}\ (aq) + H_3O^{+}\ (aq)$

c) $F^{-}\ (aq) + H_2O\ (l) \rightleftarrows HF\ (aq) + OH^{-}\ (aq)$

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Click here for a **hint**

Click here to **check your answer**

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Determine which ***reactant*** is acting as the **acid** and which reactant is acting the **base** in each of the following reactions.

HINT:

a) $HPO_4^- (aq) + HNO_3 (aq) \rightleftarrows H_2PO_4 (aq) + NO_3^- (aq)$

base **acid**

HNO_3 donated an H^+, so it acted as the **acid**. HPO_4^- accepted an H^+, so it acted as the **base**.

b) $H_2CO_3 (aq) + H_2O (l) \rightleftarrows HCO_3^- (aq) + H_3O^+ (aq)$

c) $F^- (aq) + H_2O (l) \rightleftarrows HF (aq) + OH^- (aq)$

The ***reactant*** that acts as an **acid** ***donates*** an H^+ in a chemical reaction.

The ***reactant*** that acts as an a **base** ***accepts*** an H^+ in a chemical reaction.

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.11) A compound can be classified as an "**acid**" or a "**base**" depending on its ability to *gain or lose a hydrogen ion* (H^+) in a chemical reaction.

Determine which ***reactant*** is acting as the **acid** and which reactant is acting the **base** in each of the following reactions.

a) $HPO_4^- \, (aq) + HNO_3 \, (aq) \rightleftarrows H_2PO_4 \, (aq) + NO_3^- \, (aq)$

base **acid**

HNO_3 donated an H^+, so it acted as the **acid**. HPO_4^- accepted an H^+, so it acted as the **base**.

b) $H_2CO_3 \, (aq) + H_2O \, (l) \rightleftarrows HCO_3^- \, (aq) + H_3O^+ \, (aq)$

acid **base**

H_3CO_3 donated an H^+, so it acted as the **acid**. H_2O accepted an H^+, so it acted as the **base**.

c) $F^- \, (aq) + H_2O \, (l) \rightleftarrows HF \, (aq) + OH^- \, (aq)$

base **acid**

H_2O donated an H^+, so it acted as the **acid**. F^- accepted an H^+, so it acted as the **base**.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.12) Compounds that can act as acids or as bases are called _______________ compounds.

a) binary

b) amphoteric

c) acid-base

d) ionizing

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8.12) Compounds that can act as acids or as bases are called _______________ compounds.

HINT:

a) binary

b) amphoteric

c) ~~acid-base~~

d) ~~ionizing~~

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.12) Compounds that can act as acids or as bases are called _______________ compounds.

a) binary

[✓ b) amphoteric]

c) acid-base

d) ionizing

An example of an **amphoteric compound** is the bicarbonate ion (HCO_3^-).

Bicarbonate acts as an ***acid*** in this reaction:

$$HCO_3^- + CN^- \rightleftarrows CO_3^{2-} + HCN$$

Bicarbonate acts as a ***base*** in this reaction:

$$HCO_3^- + HCl \rightleftarrows H_2CO_3 + Cl^-$$

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.13) Pairs of chemical species, such as HCl and Cl^- *or* H_3PO_3 and $H_2PO_3^-$, which differ only in the presence or absence of an H^+ are called _______________ __________.

a) acid-base pairs

b) complementary pairs

c) acid twins

d) conjugate pairs

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8.13) Pairs of chemical species, such as HCl and Cl^- *or* H_3PO_3 and $H_2PO_3^-$, which differ only in the presence or absence of an H^+ are called ________________ __________.

HINT:

a) acid-base pairs

b) complementary pairs

c) ~~acid twins~~

d) conjugate pairs

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.13) Pairs of chemical species, such as HCl and Cl^- *or* H_3PO_3 and $H_2PO_3^-$, which differ only in the presence or absence of an H^+ are called ________________ __________.

a) acid-base pairs

b) complementary pairs

c) acid twins

[✓ d) conjugate pairs]

EXPLANATION:

For a conjugate pair, the species that contains the ***extra*** H^+ is called the "**acid form**," and the species with ***one fewer*** H^+ is called the "**base form**."

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.14) For each conjugate pair, the species that contains the ***extra*** H^+ is called the “**acid form**,” and the species with ***one fewer*** H^+ is called the “**base form**.”

Identify the ***acid form*** and the ***base form*** in each of the conjugate pairs:

a) HBr and Br^-

b) H_3O^+ and H_2O

c) HCO_3^- and H_2CO_3

d) $H_2PO_3^-$ and H_3PO_3

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Click here for a **hint**

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8.14) For each conjugate pair, the species that contains the ***extra*** H^+ is called the "**acid form**," and the species with ***one fewer*** H^+ is called the "**base form**."

Identify the ***acid form*** and the ***base form*** in each of the conjugate pairs:

HINT:

a) HBr and Br^-

acid form **base form**

b) H_3O^+ and H_2O

c) HCO_3^- and H_2CO_3

d) $H_2PO_3^-$ and H_3PO_3

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.14) For each conjugate pair, the species that contains the ***extra*** H^+ is called the "**acid form**," and the species with ***one fewer*** H^+ is called the "**base form**."

Identify the ***acid form*** and the ***base form*** in each of the conjugate pairs:

a)	HBr **acid form**	and	Br^- **base form**
b)	H_3O^+ **acid form**	and	H_2O **base form**
c)	HCO_3^- **base form**	and	H_2CO_3 **acid form**
d)	$H_2PO_3^-$ **base form**	and	H_3PO_3 **acid form**

EXPLANATION:

Pairs of chemical species which differ only in the presence or absence of an H^+ are called conjugate pairs. The species that contains the ***extra*** H^+ is the "**acid form**," and the species with ***one fewer*** H^+ is the "**base form**."

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.15) What is the **acid form** of **NO_3^-**?

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8.15) What is the **acid form** of **NO_3^-**?

HINT:

The **acid form** of a conjugate pair contains ***one more*** **H^+** than the "**base form**."

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.15) What is the **acid form** of **NO_3^-**?

Answer: **HNO_3**

EXPLANATION:

The **acid form** of a conjugate pair contains ***one more*** **H^+** than the **base form**.

The **acid form** is obtained by ***adding*** an **H^+** to the **base form**.

- Note that when an **H^+** is ***added*** to a species, its charge **increases** by one charge unit. **NO_3^-** has a **1-** charge, however when an **H^+** is added, it is converted to **HNO_3**; the charge increases by one charge unit.

Conversely, the **base form** is obtained by ***removing*** an **H^+** from the **acid form**.

- Note that when an **H^+** is ***removed*** from a species, its charge **decreases** by one charge unit.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.16) What is the **base form** of $\mathbf{H_2S}$?

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8.16) What is the **base form** of **H_2S**?

HINT:

The **base form** of a conjugate pair contains ***one fewer*** **H^+** than the **acid form**.

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.16) What is the **base form** of **H_2S**?

Answer: **HS^-**

EXPLANATION:

The **base form** of a conjugate pair contains ***one fewer*** **H^+** than the **acid form**.

The **base form** is obtained by ***removing*** an **H^+** from the **acid form**.

- Note that when an **H^+** is ***removed*** from a species, its charge **decreases** by one charge unit. **H_2S** has a **1+** charge, however when an **H^+** is removed, it is converted to **HS^-**; the charge decreases by one charge unit.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.17) What is the **acid form** of $\mathbf{NH_3}$?

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8.17) What is the **acid form** of $\mathbf{NH_3}$?

HINT:
The **acid form** of a conjugate pair contains ***one more*** $\mathbf{H^+}$ than the "**base form**."

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.17) What is the **acid form** of **NH_3**?

Answer: **NH_4^+**

EXPLANATION:

The **acid form** of a conjugate pair contains ***one more*** **H^+** than the **base form**.

The **acid form** is obtained by ***adding*** an **H^+** to the **base form**.

- Note that when an **H^+** is ***added*** to a species, its charge **increases** by one charge unit. **NH_3** has a **zero** charge, however when an **H^+** is added, it is converted to **NH_4^+**; the charge increases by one charge unit.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.18) What is the **base form** of $\mathbf{H_3O^+}$?

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8.18) What is the **base form** of **H_3O^+**?

HINT:

The **base form** of a conjugate pair contains ***one fewer*** **H^+** than the **acid form**.

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.18) What is the **base form** of **H_3O^+**?

Answer: **H_2O**

EXPLANATION:

The **base form** of a conjugate pair contains ***one fewer*** **H^+** than the **acid form**.

The **base form** is obtained by ***removing*** an **H^+** from the **acid form**.

- Note that when an **H^+** is ***removed*** from a species, its charge **decreases** by one charge unit. **H_3O^+** has a **1+** charge, however when an **H^+** is removed, it is converted to **H_2O**; the charge decreases by one charge unit.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.

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8.19) Determine whether the following statements are **true** or **false**.

a) **pH** is most commonly defined as the "**negative logarithm of the *hydronium ion* concentration**."

b) The ***greater*** the concentration of H_3O^+ ions in solution, the ***greater*** the pH value.

c) Numbers to the right of the decimal point are not significant in pH values.

d) We do not use units in pH values.

e) If we know the concentration of hydroxide ions $[OH^-]$ in a solution, we can determine the pH value.

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Go to next question

8.19) Determine whether the following statements are **true** or **false**.

a) **pH** is most commonly defined as the "**negative logarithm of the *hydronium ion* concentration**."

b) The ***greater*** the concentration of H_3O^+ ions in solution, the ***greater*** the pH value.

HINT: Consider the effect of the negative sign in our definition of pH: $pH = -log[H_3O^+]$.

c) Numbers to the right of the decimal point are not significant in pH values.

d) We do not use units in pH values.

e) If we know the concentration of hydroxide ions $[OH^-]$ in a solution, we can determine the pH value.

HINT: $[OH^-][H_3O^+] = 1.0 \text{ x } 10^{-14} M^2$

For more help: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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8.19) Determine whether the following statements are **true** or **false**.

a) **pH** is most commonly defined as the "**negative logarithm of the *hydronium ion* concentration**." **true**

b) The ***greater*** the concentration of H_3O^+ ions in solution, the ***greater*** the pH value. **false**

- Because of the negative sign in our definition of pH (pH = -log[H_3O^+]), the ***greater*** the hydronium ion concentration, the ***lower*** the pH value.

c) Numbers to the right of the decimal point are not significant in pH values. **false**

- Numbers to the **left** of the decimal point are not significant in pH values.
- **Another way to say this is**, "only numbers ***to the right*** of the decimal place **are** significant in pH values.

d) We do not use units in pH values. **true**

e) If we know the concentration of hydroxide ions [OH^-] in a solution, we can determine the pH value. **true**

- Because [OH^-][H_3O^+] = 1.0 x 10^{-14} M^2 , whenever we know [OH^-], we can calculate [H_3O^+]. Once [H_3O^+] is known, the pH can be calculated using: pH = -log[H_3O^+].

For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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8.20)

i) What is the pH of an aqueous solution with $[H_3O^+] = 0.001$ M?

ii) What is the pH of an aqueous solution with $[H_3O^+] = 4.6 \times 10^{-5}$ M?

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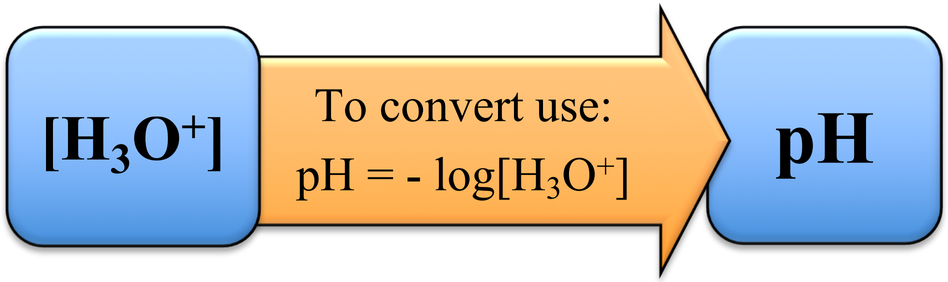
8.20)

i) What is the pH of an aqueous solution with $[H_3O^+] = 0.001$ M?

HINT:



$[H_3O^+]$

To convert use:

$pH = - \log[H_3O^+]$

pH

ii) What is the pH of an aqueous solution with $[H_3O^+] = 4.6 \times 10^{-5}$ M?

For more help: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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Click here to **check your answer**

Go to next question

8.20)

i) What is the pH of an aqueous solution with $[H_3O^+] = \boxed{0.001}$ M? **ANSWER: pH = 3.<u>0</u>**

Note: There is **one significant figure** in the concentration, so there will be **one significant figure** in the pH.

- Recall that only numbers ***to the right*** of the decimal place **are** significant in pH values.

ii) What is the pH of an aqueous solution with $[H_3O^+] = \boxed{4.6 \text{ x } 10^{-5}}$ M? **ANSWER: pH = 4.<u>34</u>**

Note: There are **two significant figures** in the concentration, so there will be **two significant figures** in the pH.

CLICK HERE to see the **complete solution** for this problem

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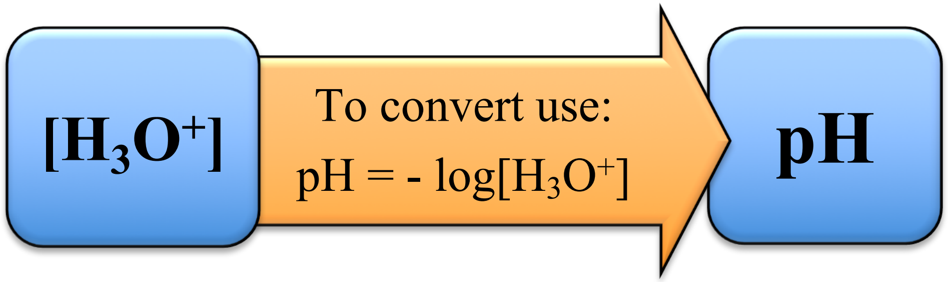
8.20)

i) What is the pH of an aqueous solution with $[H_3O^+] = \boxed{0.001}$ M? **ANSWER: pH =** $\mathbf{3.\underline{0}}$

Note: There is **one significant figure** in the concentration, so there will be **one significant figure** in the pH.

- Recall that only numbers ***to the right*** of the decimal place **are** significant in pH values.



$$pH = -\log[H_3O^+] = -\log[0.001] = -(-3.0) = \mathbf{3.0}$$

ii) What is the pH of an aqueous solution with $[H_3O^+] = \boxed{4.6 \times 10^{-5}}$ M? **ANSWER: pH =** $\mathbf{4.\underline{34}}$

Note: There are **two significant figures** in the concentration, so there will be **two significant figures** in the pH.

$$pH = -\log[H_3O^+] = -\log[4.6 \times 10^{-5}] = -(-4.34) = \mathbf{4.34}$$

For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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8.21)

i) What is the pH of an aqueous solution with $[OH^-] = 0.005$ M?

ii) What is the pH of an aqueous solution with $[OH^-] = 7.2 \times 10^{-8}$ M?

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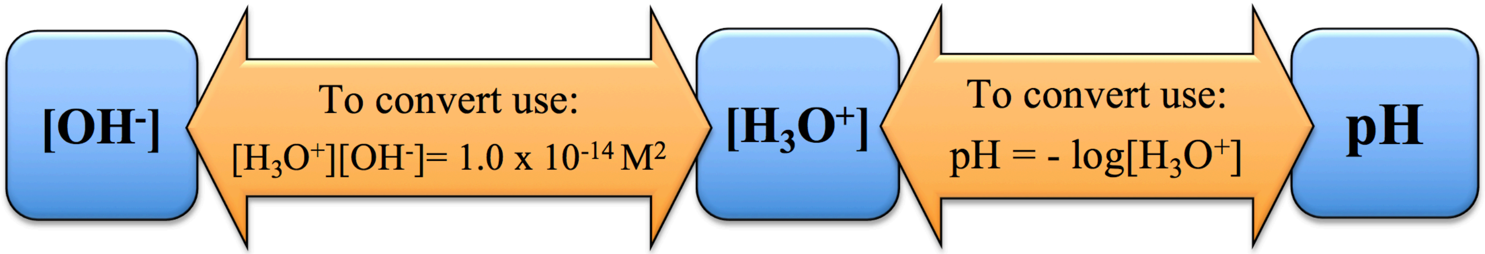
8.21)

i) What is the pH of an aqueous solution with $[OH^-] = 0.005$ M?

HINT:



$[OH^-]$

To convert use:
$[H_3O^+][OH^-] = 1.0 \times 10^{-14}\ M^2$

$[H_3O^+]$

To convert use:
$pH = -\log[H_3O^+]$

pH

In these problems, we know $[OH^-]$, so we can calculate the concentration of $[H_3O^+]$.

Once the $[H_3O^+]$ is determined, it can be used to calculate the pH.

ii) What is the pH of an aqueous solution with $[OH^-] = 7.2 \times 10^{-8}$ M?

For more help: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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8.21)

i) What is the pH of an aqueous solution with [OH$^-$] = 0.005 M? **ANSWER: pH = 11.7**

Note: There is **one significant figure** in the concentration, so there will be **one significant figure** in the pH.

ii) What is the pH of an aqueous solution with [OH$^-$] = 7.2×10^{-8} M? **ANSWER: pH = 6.85**

Note: There are **two significant figures** in the concentration, so there will be **two significant figure** in the pH.

CLICK HERE to see the **complete solution** for this problem

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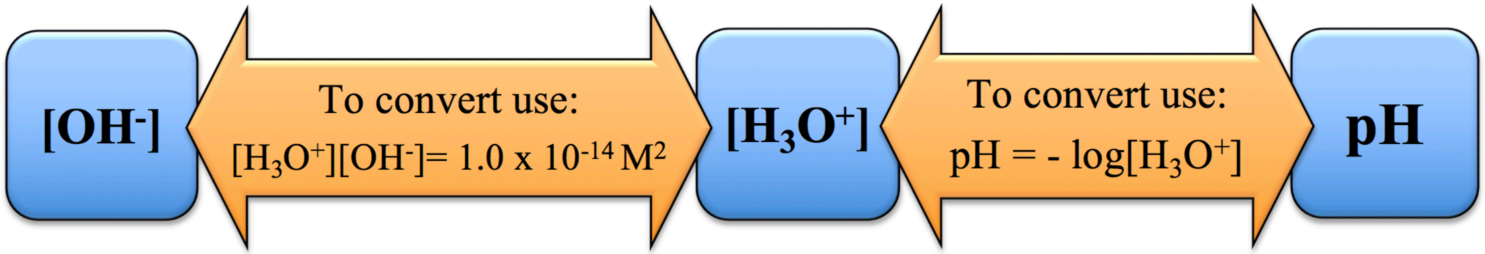
8.21)

i) What is the pH of an aqueous solution with $[OH^-] = 0.005$ M? **ANSWER: pH = 11.7**

Note: There is **one significant figure** in the concentration, so there will be **one significant figure** in the pH.



In this problem, we know $[OH^-] = 0.005$ M, we can calculate the concentration of $[H_3O^+]$:

$$[H_3O^+] = \frac{1.0 \times 10^{-14}\ M^2}{[OH^-]} = \frac{1.0 \times 10^{-14}\ M^2}{0.005\ M} = 2 \times 10^{-12}\ M$$

Once the $[H_3O^+]$ is determined, it can be used to calculate the pH:

$$pH = -\log[H_3O^+] = -\log[2 \times 10^{-12}] = -(-11.7) = \mathbf{11.7}$$

ii) What is the pH of an aqueous solution with $[OH^-] = 7.2 \times 10^{-8}$ M? **ANSWER: pH = 6.85**

Note: There are **two significant figures** in the concentration, so there will be **two significant figure** in the pH.

$$[H_3O^+] = \frac{1.0 \times 10^{-14}\ M^2}{[OH^-]} = \frac{1.0 \times 10^{-14}\ M^2}{7.2 \times 10^{-8}\ M} = 1.4 \times 10^{-7}\ M$$

$$pH = -\log[H_3O^+] = -\log[1.4 \times 10^{-7}] = -(-6.85) = \mathbf{6.85}$$

For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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8.22)

i) What is the $[H_3O^+]$ of an aqueous solution with pH = 1.25?

ii) What is the $[H_3O^+]$ of an aqueous solution with pH = 10.6?

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8.22)

i) What is the $[H_3O^+]$ of an aqueous solution with pH = 1.25?

HINT:

$[H_3O^+]$

To convert use:

$pH = -\log[H_3O^+]$

or

$10^{-pH} = [H_3O^+]$

pH

$$pH = -\log[H_3O^+]$$

Multiply both sides by (-1)

$$-pH = \log[H_3O^+]$$

Take the antilog of both sides

$$10^{(-pH)} = [H_3O^+]$$

ii) What is the $[H_3O^+]$ of an aqueous solution with pH = 10.6?

For more help: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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Click here to check **your answer**

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8.22)

i) What is the [H_3O^+] of an aqueous solution with pH = $1.\boxed{25}$? **ANSWER: $\boxed{0.056}$ M**

Note: There are **two significant figures** in the pH, so there will be **two significant figures** in the concentration.

- Recall that only numbers ***to the right*** of the decimal place **are** significant in pH values.
- Did you include the unit in your answer? The concertation units here are molar (M) *or* moles/L.

ii) What is the [H_3O^+] of an aqueous solution with pH = $10.\boxed{6}$? **ANSWER: $\boxed{3 \times 10^{-11}}$ M**

Note: There is **one significant figure** in the pH, so there will be **one significant figure** in the concentration.

CLICK HERE to see the **complete solution** for this problem

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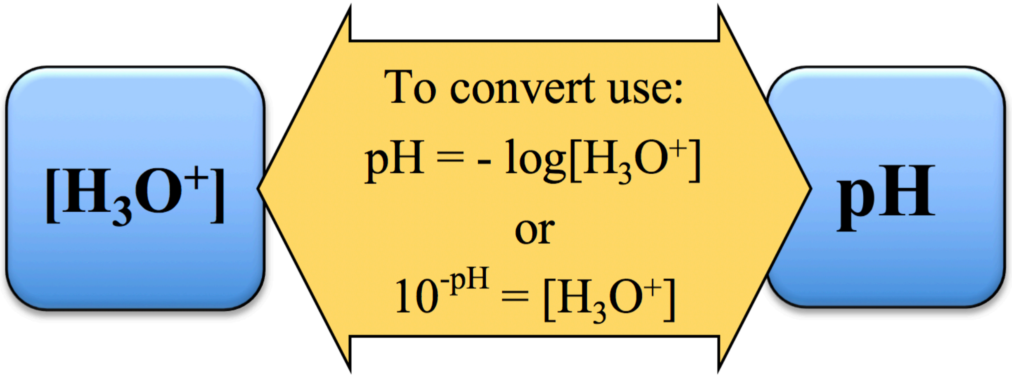
8.22)

i) What is the $[H_3O^+]$ of an aqueous solution with pH = 1.$\boxed{25}$? **ANSWER: $\boxed{0.056}$ M**

Note: There are **two significant figures** in the pH, so there will be **two significant figures** in the concentration.

- Recall that only numbers ***to the right*** of the decimal place **are** significant in pH values.



$$pH = -\log[H_3O^+]$$

Multiply both sides by (-1)

$$1.25 = -\log[H_3O^+]$$

$$-1.25 = \log[H_3O^+]$$

Take the antilog of both sides

$$10^{(-1.25)} = [H_3O^+]$$

$$[H_3O^+] = \mathbf{0.056\ M}$$

ii) What is the $[H_3O^+]$ of an aqueous solution with pH = 10.$\boxed{6}$? **ANSWER: $\boxed{3 \times 10^{-11}}$ M**

Note: There is **one significant figure** in the pH, so there will be **one significant figure** in the concentration.

$$pH = -\log[H_3O^+]$$

Multiply both sides by (-1)

$$10.6 = -\log[H_3O^+]$$

$$-10.6 = \log[H_3O^+]$$

Take the antilog of both sides

$$10^{(-10.6)} = [H_3O^+]$$

$$[H_3O^+] = \mathbf{3 \times 10^{-11}\ M}$$

For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.

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8.23) Solutions are characterized as acidic, basic, or neutral by the relative amounts of H_3O^+ and OH^- that are present.

i) Solutions that contain more H_3O^+ than OH^- are called ____________ solutions.

a) acidic
b) neutral
c) basic
d) buffer

ii) Solutions that contain more OH^- than H_3O^+ are called ____________ solutions.

a) acidic
b) neutral
c) basic
d) buffer

iii) Solutions that contain equal concentrations of H_3O^+ and OH^- are called ____________ solutions.

a) balanced
b) neutral
c) pH
d) buffer

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8.23) Solutions are characterized as acidic, basic, or neutral by the relative amounts of H_3O^+ and OH^- that are present.

i) Solutions that contain more H_3O^+ than OH^- are called ____________ solutions.

HINT:
a) acidic
b) neutral
c) basic
d) ~~buffer~~

ii) Solutions that contain more OH^- than H_3O^+ are called ____________ solutions.

HINT:
a) acidic
b) neutral
c) basic
d) ~~buffer~~

iii) Solutions that contain equal concentrations of H_3O^+ and OH^- are called ____________ solutions.

HINT:
a) ~~balanced~~
b) neutral
c) pH
d) ~~buffer~~

For more help: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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8.23) Solutions are characterized as acidic, basic, or neutral by the relative amounts of H_3O^+ and OH^- that are present.

i) Solutions that contain more H_3O^+ than OH^- are called ___________ solutions.

[✓ a) acidic]
b) neutral
c) basic
d) buffer A ***buffer solution*** is a solution that resists changes in pH when a small amount of acid or base is added.

ii) Solutions that contain more OH^- than H_3O^+ are called ___________ solutions.

a) acidic
b) neutral
[✓ c) basic]
d) buffer

iii) Solutions that contain equal concentrations of H_3O^+ and OH^- are called ___________ solutions.

a) balanced
[✓ b) neutral]
c) pH
d) buffer

Solution Characterization	pH	$[H_3O^+]$	$[OH^-]$
Acidic	less than 7.00	greater than 1.0×10^{-7} M	less than 1.0×10^{-7} M
Neutral	7.00	1.0×10^{-7} M	1.0×10^{-7} M
Basic	greater than 7.00	less than 1.0×10^{-7} M	greater than 1.0×10^{-7} M

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For more details: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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8.24) Compounds are characterized as ______________________________ depending on whether they donate or accept H^+ in a particular acid-base reaction.

a) neutral or or non neutral

b) acidic or basic

c) acids or bases

d) buffers

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Click here for a **hint**

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8.24) Compounds are characterized as ______________________ depending on whether they donate or accept H^+ in a particular acid-base reaction.

a) neutral or or non neutral

b) acidic or basic

HINT:

c) acids or bases

d) ~~buffers~~

For more help: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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8.24) Compounds are characterized as __________________________ depending on whether they donate or accept H^+ in a particular acid-base reaction.

a) neutral or or non neutral

Solutions (not compounds) that contain equal concentrations of H_3O^+ and OH^- are are characterized as **neutral**.

b) acidic or basic

Solutions (not compounds) are characterized as **acidic**, **basic**, or neutral by the relative amounts of H_3O^+ and OH^- that are present.

[✓ c) acids or bases]

d) buffers

A ***buffer solution*** is a solution that resists changes in pH when a small amount of acid or base is added.

For more details: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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8.25) For each of the following, write whether the solution condition describes an **acidic**, **basic**, or **neutral** solution.

a) $pH = 3.9$

b) $[H_3O^+] = 1 \times 10^{-5}$ M

c) $[OH^-] = 1 \times 10^{-7}$ M

d) $[OH^-] > [H_3O^+]$

e) $[H_3O^+] > [OH^-]$

f) $pH = 9.7$

g) $pH = 2.0$

h) $[OH^-] = 6.8 \times 10^{-8}$ M

i) $pH = 12.6$

j) $[H_3O^+] = [OH^-]$

k) $pH = 7.00$

l) $[H_3O^+] = 1.0 \times 10^{-7}$ M

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Click here for a **hint**

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8.25) For each of the following, write whether the solution condition describes an **acidic**, **basic**, or **neutral** solution.

a) pH = 3.9

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c) $[OH^-] = 1 \times 10^{-7}$ M

d) $[OH^-] > [H_3O^+]$

e) $[H_3O^+] > [OH^-]$

f) pH = 9.7

g) pH = 2.0

h) $[OH^-] = 6.8 \times 10^{-8}$ M

i) pH = 12.6

j) $[H_3O^+] = [OH^-]$

k) pH = 7.00

l) $[H_3O^+] = 1.0 \times 10^{-7}$ M

HINT:

Solution Characterization	pH	$[H_3O^+]$	$[OH^-]$
Acidic	less than 7.00	greater than 1.0×10^{-7} M	less than 1.0×10^{-7} M
Neutral	7.00	1.0×10^{-7} M	1.0×10^{-7} M
Basic	greater than 7.00	less than 1.0×10^{-7} M	greater than 1.0×10^{-7} M

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For more help: see chapter 8 part 6 video
or chapter 8 section 6 in the textbook.

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8.25) For each of the following, write whether the solution condition describes an **acidic**, **basic**, or **neutral** solution.

a) pH = 3.9 **acidic**

b) $[H_3O^+] = 1 \times 10^{-5}$ M **acidic**

c) $[OH^-] = 1 \times 10^{-7}$ M **neutral**

d) $[OH^-] > [H_3O^+]$ **basic**

e) $[H_3O^+] > [OH^-]$ **acidic**

f) pH = 9.7 **basic**

g) pH = 2.0 **acidic**

h) $[OH^-] = 6.8 \times 10^{-8}$ M **acidic**

i) pH = 12.6 **basic**

j) $[H_3O^+] = [OH^-]$ **neutral**

k) pH = 7.00 **neutral**

l) $[H_3O^+] = 1.0 \times 10^{-7}$ M **neutral**

EXPLANATION:

Solution Characterization	pH	$[H_3O^+]$	$[OH^-]$
Acidic	less than 7.00	greater than 1.0×10^{-7} M	less than 1.0×10^{-7} M
Neutral	7.00	1.0×10^{-7} M	1.0×10^{-7} M
Basic	greater than 7.00	less than 1.0×10^{-7} M	greater than 1.0×10^{-7} M

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For more details: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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8.26) Use the table to determine which is a **stronger acid**, *boric acid* or *acetic acid*.

Various Acids and Their Acidity Constants

Acid Name	Acid Formula	K_a
Perchloric acid	$HClO_4$	1×10^9 M (estimated)
Hydrochloric acid	HCl	1×10^7 M (estimated)
Chloric acid	$HClO_3$	1×10^3 M (estimated)
Phosphoric acid	H_3PO_4	7.5×10^{-3} M
Hydrofluoric acid	HF	6.6×10^{-4} M
Acetic acid	CH_3CO_2H	1.8×10^{-5} M
Carbonic acid	H_2CO_3	4.4×10^{-7} M
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8} M
Boric acid	H_3BO_3	5.7×10^{-10} M
Ammonium ion	NH_4^+	5.6×10^{-10} M
Hydrocyanic acid	HCN	4.9×10^{-10} M
Bicarbonate ion	HCO_3^-	5.6×10^{-11} M
Methylammonium ion	$CH_3NH_3^+$	2.4×10^{-11} M
Hydrogen phosphate ion	HPO_4^-	4.2×10^{-13} M

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Click here for a **hint**

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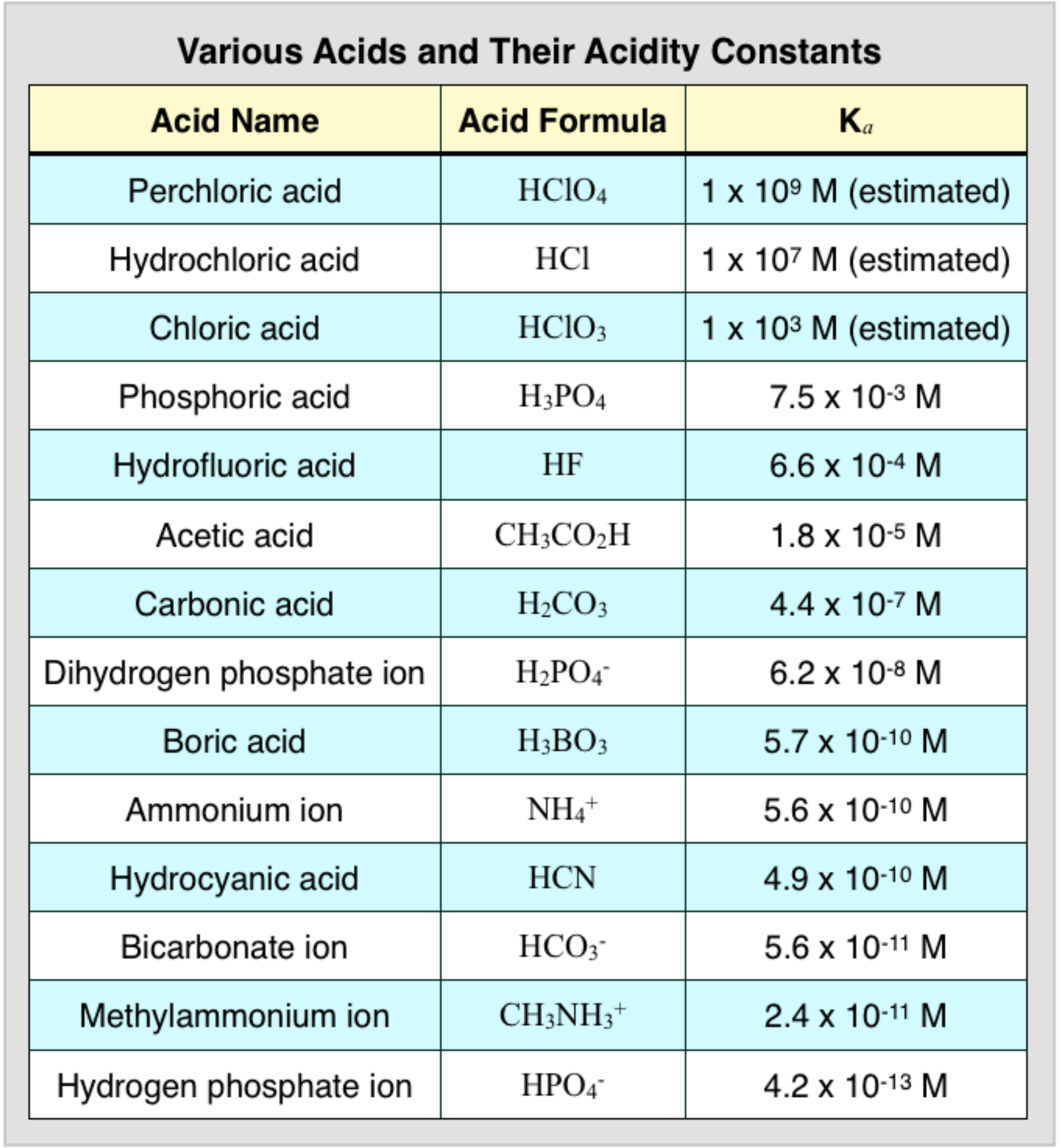
8.26) Use the table to determine which is a **stronger acid**, *boric acid* or *acetic acid*.

HINT:
The greater the K_a, the stronger the acid.

For more help: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.



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Hydrofluoric acid	HF	6.6×10^{-4} M
Acetic acid	CH_3CO_2H	1.8×10^{-5} M
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Ammonium ion	NH_4^+	5.6×10^{-10} M
Hydrocyanic acid	HCN	4.9×10^{-10} M
Bicarbonate ion	HCO_3^-	5.6×10^{-11} M
Methylammonium ion	$CH_3NH_3^+$	2.4×10^{-11} M
Hydrogen phosphate ion	HPO_4^-	4.2×10^{-13} M

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8.26) Use the table to determine which is a **stronger acid**, *boric acid* or *acetic acid*.

ANSWER: ***acetic acid*** is the stronger acid

EXPLANATION:

The greater the K_a, the stronger the acid.

Acetic acid has a greater K_a than *boric acid*.

$$1.8 \times 10^{-5} > 5.7 \times 10^{-10}$$

When acids are placed in pure water, the stronger the acid, the greater the concentration of H_3O^+.

For more details: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

Various Acids and Their Acidity Constants

Acid Name	Acid Formula	K_a
Perchloric acid	$HClO_4$	1×10^{9} M (estimated)
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Hydrocyanic acid	HCN	4.9×10^{-10} M
Bicarbonate ion	HCO_3^-	5.6×10^{-11} M
Methylammonium ion	$CH_3NH_3^+$	2.4×10^{-11} M
Hydrogen phosphate ion	HPO_4^-	4.2×10^{-13} M

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8.27) An acidic solution will react with a hydroxide-containing base compound to produce a water and an ionic compound in a reaction called **neutralization**. An example of a ***neutralization reaction*** is the reaction of hydrochloric acid (HCl) and sodium hydroxide:

$$\mathrm{HBr}\ (aq) + \mathrm{NaOH}\ (aq) \rightleftarrows \mathrm{H_2O}\ (l) + \mathrm{NaBr}\ (aq)$$

In *neutralization reactions*, the H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid (Br^- in this example) combines with the cation of the base (Na^+ in this example) to make an ionic compound called a **salt** (NaBr in this example). Although sodium chloride is commonly called “salt,” the chemical definition states that a salt is *an ionic compound formed in a neutralization reaction*.

Predict the products of the following *neutralization reactions*:

$$\mathrm{HNO_3}\ (aq) + \mathrm{NaOH}\ (aq) \rightleftarrows \underline{\hspace{3cm}} + \underline{\hspace{3cm}}$$

$$\mathrm{HNO_3}\ (aq) + \mathrm{KOH}\ (aq) \rightleftarrows \underline{\hspace{3cm}} + \underline{\hspace{3cm}}$$

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8.27) An acidic solution will react with a hydroxide-containing base compound to produce a water and an ionic compound in a reaction called **neutralization**. An example of a ***neutralization reaction*** is the reaction of hydrochloric acid (HCl) and sodium hydroxide:

$$HBr\ (aq) + NaOH\ (aq) \rightleftarrows H_2O\ (l) + NaBr\ (aq)$$

In *neutralization reactions*, the H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid (Br^- in this example) combines with the cation of the base (Na^+ in this example) to make an ionic compound called a **salt** (NaBr in this example). Although sodium chloride is commonly called "salt," the chemical definition states that a salt is *an ionic compound formed in a neutralization reaction*.

Predict the products of the following *neutralization reactions*:

HINT: $HNO_3\ (aq) + NaOH\ (aq) \rightleftarrows$ **$H_2O\ (l)$** ______ + ______________

The H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid combines with the cation of the base (Na^+ in this problem) to make an ionic compound called a **salt**.

HINT: $HNO_3\ (aq) + KOH\ (aq) \rightleftarrows$ **$H_2O\ (l)$** ______ + ______________

The H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid combines with the cation of the base (K^+ in this problem) to make an ionic compound called a **salt**.

For more help: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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8.27) An acidic solution will react with a hydroxide-containing base compound to produce a water and an ionic compound in a reaction called **neutralization**. An example of a ***neutralization reaction*** is the reaction of hydrochloric acid (HCl) and sodium hydroxide:

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In *neutralization reactions*, the H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid (Br^- in this example) combines with the cation of the base (Na^+ in this example) to make an ionic compound called a **salt** (NaBr in this example). Although sodium chloride is commonly called "salt," the chemical definition states that a salt is *an ionic compound formed in a neutralization reaction*.

Predict the products of the following *neutralization reactions*:

$HNO_3\ (aq)$ + $NaOH\ (aq)$ $\rightleftarrows$ **$H_2O\ (l)$** + **$NaNO_3\ (aq)$**

The H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid (**NO_3^-** in this problem) combines with the cation of the base (Na^+ in this problem) to make an ionic compound called a **salt** ($NaNO_3$).

$HNO_3\ (aq)$ + $KOH\ (aq)$ $\rightleftarrows$ **$H_2O\ (l)$** + **$KNO_3\ (aq)$**

The H^+ from the ***acid*** bonds to the OH^- to produce H_2O. The base form of the acid (**NO_3^-** in this problem) combines with the cation of the base (K^+ in this problem) to make an ionic compound called a **salt** (KNO_3).

For more details: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.

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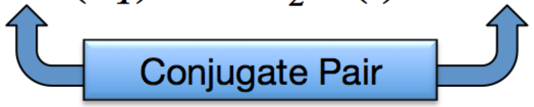
8.28) The **general form** of a chemical equation for an acid reacting with water to produce its base form and hydronium can be written as:

$$HA\,(aq) + H_2O\,(l) \rightleftarrows A^-\,(aq) + H_3O^+\,(aq)$$



HA represents the ***acid form***, and A^- represents the ***base form*** of any **conjugate pair.**

- When the **pH** of a solution is ***less*** than the **pK_a** of an acid, then the concentration of the ***acid form***, [HA], is ***greater than*** the concentration of the ***base form***, [A^-].
  - In this case, we say that ***the acid form is predominant***.
- When the **pH** of a solution is ***greater*** than the **pK_a** of an acid, then the concentration of the ***base form***, [A], is ***greater than*** the concentration of the ***acid form***, [HA].
  - In this case, we say that ***the base form is predominant***.
- When the **pH** of a solution is ***equal to*** the **pK_a** of an acid, then the concentration of the ***acid form***, [HA], is ***equal to*** the concentration of the ***base form***, [A^-].

QUESTION: When hydrocyanic acid (HCN) is placed in water, a chemical reaction occurs and an equilibrium is established as shown below. The **pK_a** of HCN is 9.31. Predict whether the ***acid form*** (HCN) or the ***base form*** (CN^-) is predominant at pH = 7.4.

$$HCN\,(aq) + H_2O\,(l) \rightleftarrows CN^-\,(aq) + H_3O^+\,(aq)$$

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Click here for a **hint**

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Go to next question

8.28) The **general form** of a chemical equation for an acid reacting with water to produce its base form and hydronium can be written as:

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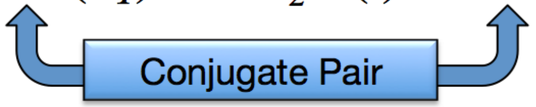

Conjugate Pair

HA represents the ***acid form***, and A^- represents the ***base form*** of any **conjugate pair.**

- When the **pH** of a solution is ***less*** than the **pK_a** of an acid, then the concentration of the ***acid form***, [HA], is ***greater than*** the concentration of the ***base form***, [A^-].
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$$HCN\ (aq) + H_2O\ (l) \rightleftarrows CN^-\ (aq) + H_3O^+\ (aq)$$

HINT: Compare the pH to the pK_a.

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For more help: see chapter 8 part 7 video
or chapter 8 section 7 in the textbook.

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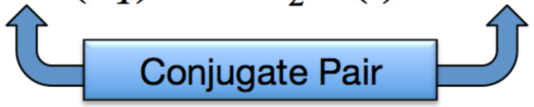
Go to next question

8.28) The **general form** of a chemical equation for an acid reacting with water to produce its base form and hydronium can be written as:

$$HA\ (aq) + H_2O\ (l) \rightleftarrows A^-\ (aq) + H_3O^+\ (aq)$$



Conjugate Pair

HA represents the ***acid form***, and A^- represents the ***base form*** of any **conjugate pair.**

- When the **pH** of a solution is ***less*** than the **pK_a** of an acid, then the concentration of the ***acid form***, [HA], is ***greater than*** the concentration of the ***base form***, [A^-].
  - In this case, we say that ***the acid form is predominant***.
- When the **pH** of a solution is ***greater*** than the **pK_a** of an acid, then the concentration of the ***base form***, [A], is ***greater than*** the concentration of the ***acid form***, [HA].
  - In this case, we say that ***the base form is predominant***.
- When the **pH** of a solution is ***equal to*** the **pK_a** of an acid, then the concentration of the ***acid form***, [HA], is ***equal to*** the concentration of the ***base form***, [A^-].

QUESTION: When hydrocyanic acid (HCN) is placed in water, a chemical reaction occurs and an equilibrium is established as shown below. The **pK_a** of HCN is 9.31. Predict whether the ***acid form*** (HCN) or the ***base form*** (CN^-) is predominant at pH = 7.4.

$$HCN\ (aq) + H_2O\ (l) \rightleftarrows CN^-\ (aq) + H_3O^+\ (aq)$$

ANSWER:
The ***acid form*** (HCN) is predominant.

EXPLANATION: The **pH** (7.4) is ***less*** than the **pK_a** (9.31), therefore the concentration of the ***acid form***, [HCN], is ***greater than*** the concentration of the ***base form***, [CN^-]. In this case, we say that ***the acid form is predominant***.

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For more details: see chapter 8 part 7 video or chapter 8 section 7 in the textbook.

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8.29) For each of the following conjugate pairs, predict whether the ***acid form*** or the ***base form*** is predominant at the given pH.

a) HF/F^- at pH = 9.7

b) $CH_3NH_3^+/CH_3NH_2$ at pH = 7.0

c) H_2CO_3/HCO_3^- at pH = 8.5

Acid Name	**Acid Formula**	**K*a***	**pK*a*** pK*a* = -log(K*a*)
Perchloric acid	$HClO_4$	1 x 10^9 M (estimated)	-9.0 (estimated)
Hydrochloric acid	HCl	1 x 10^7 M (estimated)	-7.0 (estimated)
Chloric acid	$HClO_3$	1 x 10^3 M (estimated)	-3.0 (estimated)
Phosphoric acid	H_3PO_4	7.5 x 10^{-3} M	2.12
Hydrofluoric acid	HF	6.6 x 10^{-4} M	3.18
Acetic acid	CH_3CO_2H	1.8 x 10^{-5} M	4.74
Carbonic acid	H_2CO_3	4.4 x 10^{-7} M	6.36
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2 x 10^{-8} M	7.21
Boric acid	H_3BO_3	5.7 x 10^{-10} M	9.24
Ammonium ion	NH_4^+	5.6 x 10^{-10} M	9.25
Hydrocyanic acid	HCN	4.9 x 10^{-10} M	9.31
Bicarbonate ion	HCO_3^-	5.6 x 10^{-11} M	10.25
Methylammonium ion	$CH_3NH_3^+$	2.4 x 10^{-11} M	10.62
Hydrogen phosphate ion	HPO_4^{2-}	4.2 x 10^{-13} M	12.38

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Click here for a **hint**

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8.29) For each of the following conjugate pairs, predict whether the ***acid form*** or the ***base form*** is predominant at the given pH.

a) HF/F^- at pH = 9.7

b) $CH_3NH_3^+/CH_3NH_2$ at pH = 7.0

c) H_2CO_3/HCO_3^- at pH = 8.5

HINT: Compare the pH to the pK_a

Solution Condition	Relative Amounts of Acid and Base Forms
$pH < pK_a$	$[HA] > [A^-]$
$pH > pK_a$	$[A^-] > [HA]$
$pH = pK_a$	$[HA] = [A^-]$

For more help: see chapter 8 part 7 video or chapter 8 section 7 in the textbook.

Acid Name	Acid Formula	K_a	pK_a $pK_a = -\log(K_a)$
Perchloric acid	$HClO_4$	1×10^9 M (estimated)	-9.0 (estimated)
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8.29) For each of the following conjugate pairs, predict whether the ***acid form*** or the ***base form*** is predominant at the given pH.

a) HF/F^- at pH = 9.7 ***base form*** (F^-)
pH > pK$_a$

b) $CH_3NH_3^+/CH_3NH_2$ at pH = 7.0 ***acid form*** ($CH_3NH_3^+$)
pH < pK$_a$

c) H_2CO_3/HCO_3^- at pH = 8.5 ***base form*** (HCO_3^-)
pH > pK$_a$

EXPLANATION:
Compare the pH to the pK$_a$

Solution Condition	Relative Amounts of Acid and Base Forms
$pH < pK_a$	$[HA] > [A^-]$
$pH > pK_a$	$[A^-] > [HA]$
$pH = pK_a$	$[HA] = [A^-]$

Acid Name	Acid Formula	K$_a$	pK$_a$ pK$_a$ = -log(K$_a$)
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For more details: see chapter 8 part 7 video or chapter 8 section 7 in the textbook.

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8.30) Label each of the statements below as ***true*** or ***false***.

a) A *buffer solution* is a solution that resists changes in pH when a small amount of acid or base is added.

b) A buffer is a solution that is made with fairly low concentrations of the acid and base forms of a conjugate pair.

c) The blood’s buffering system maintains the pH in the normal range, which is 6.35 - 8.45.

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Click here for a **hint**

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Go to next question

8.30) Label each of the statements below as ***true*** or ***false***.

a) A *buffer solution* is a solution that resists changes in pH when a small amount of acid or base is added.

b) A buffer is a solution that is made with fairly low concentrations of the acid and base forms of a conjugate pair.

c) The blood's buffering system maintains the pH in the normal range, which is 6.35 - 8.45.

HINT: The blood's buffering system does maintains the pH in the normal range. Do you recall the normal pH range for blood?

For more help: see chapter 8 part 8 video or chapter 8 section 7 in the textbook.

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8.30) Label each of the statements below as ***true*** or ***false***.

a) A *buffer solution* is a solution that resists changes in pH when a small amount of acid or base is added. ***true***

b) A buffer is a solution that is made with fairly low concentrations of the acid and base forms of a conjugate pair. ***false***

- A buffer is a solution that is made with fairly **HIGH** concentrations of the acid and base forms of a conjugate pair.

c) The blood's buffering system maintains the pH in the normal range, which is 6.35 - 8.45. ***false***

- The blood's buffering system maintains the pH in the normal range, which is **7.35 - 7.45**.

For more details: see chapter 8 part 8 video or chapter 8 section 7 in the textbook.

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8.31) Which of the following conjugate pairs are important *extracellular* (outside of cells) buffers?
(NOTE: There may be more than one correct choice.)

a) hydrochloric acid ($\mathbf{HCl}$)/chloride ($\mathbf{Cl^-}$)

b) carbonic acid ($\mathbf{H_2CO_3}$)/bicarbonate ($\mathbf{HCO_3^-}$)

c) chloric acid ($\mathbf{HClO_3}$)/chlorate ($\mathbf{ClO_3^-}$)

d) ammonium ($\mathbf{NH_4^+}$)/ammonia ($\mathbf{NH_3}$)

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Click here for a **hint**

Click here to **check your answer**

This is the last problem.

8.31) Which of the following conjugate pairs are important *extracellular* (outside of cells) buffers? (NOTE: There may be more than one correct choice.)

HINT:

a) ~~hydrochloric acid (**HCl**)/chloride (**Cl^-**)~~

b) carbonic acid (**H_2CO_3**)/bicarbonate (**HCO_3^-**)

c) chloric acid (**$HClO_3$**)/chlorate (**ClO_3^-**)

d) ammonium (**NH_4^+**)/ammonia (**NH_3**)

For more help: see chapter 8 part 8 video or chapter 8 section 7 in the textbook.

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Click here to **check your answer**

This is the last problem.

8.31) Which of the following conjugate pairs are important *extracellular* (outside of cells) buffers?
(NOTE: There may be more than one correct choice.)

a) hydrochloric acid (**HCl**)/chloride (**Cl^-**)

[✓ b) carbonic acid (**H_2CO_3**)/bicarbonate (**HCO_3^-**)]

c) chloric acid (**$HClO_3$**)/chlorate (**ClO_3^-**)

[✓ d) ammonium (**NH_4^+**)/ammonia (**NH_3**)]

EXPLANATION:

Important ***extracellular*** (outside of cells) buffers, in solutions such as blood or interstitial fluids, are the carbonic acid (**H_2CO_3**)/bicarbonate (**HCO_3^-**) and the ammonium (**NH_4^+**)/ammonia (**NH_3**) conjugate pairs.

- In blood, the carbonic acid (**H_2CO_3**)/bicarbonate (**HCO_3^-**) buffering pair is especially useful because the buffer conjugate pair concentrations ([**H_2CO_3**] *and* [**HCO_3^-**]) are replenished through cellular respiration and can be controlled through breathing.

An important ***intracellular*** (within cells) buffer is the dihydrogen phosphate/hydrogen phosphate conjugate pair. **Proteins** also act as *intracellular* buffers. In chapter 13 you will learn how proteins can donate or accept H^+.

For more details: see chapter 8 part 8 video or chapter 8 section 7 in the textbook.

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This is the last problem.