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).
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.
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.

HINT:

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.
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.
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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d) increase.

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

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

\[ \text{N}_2 (g) + 3 \text{H}_2 (g) \rightleftharpoons 2 \text{NH}_3 (g) \]

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

a) \(N_2\) (only)

b) \(H_2\) (only)

c) \(N_2\) and \(H_2\) (only)

d) \(NH_3\) (only)

e) \(H_2\), \(N_2\), and \(NH_3\)
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c) \( \text{N}_2 \) and \( \text{H}_2 \) (only)

d) \( \text{NH}_3 \) (only)

e) \( \text{H}_2, \text{N}_2, \) and \( \text{NH}_3 \)

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

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d) \(\text{NH}_3\) (only)

e) \(\text{H}_2\), \(\text{N}_2\), and \(\text{NH}_3\)

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

The square brackets, \([\quad]\), indicate concentration in molarity, for example, “[A]” means “molarity of substance A.”

Write the equilibrium expression for the following reaction: \(3 \text{H}_2(g) + \text{N}_2(g) \rightleftharpoons 2 \text{NH}_3(g)\)

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.
8.4) For any chemical reaction at equilibrium:

\[
aA + bB \rightleftharpoons cC + dD
\]

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:

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

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: \(3\ H_2(g) + N_2(g) \rightleftharpoons 2\ NH_3(g)\)

**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.

**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.
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:

\[
3 \text{H}_2(g) + \text{N}_2(g) \rightleftharpoons 2 \text{NH}_3(g)
\]

**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.

\[
K_{eq} = \frac{[\text{NH}_3]^2}{[\text{H}_2]^3 \cdot [\text{N}_2]}
\]

**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.

**For more details**: see chapter 8 part 1 video or chapter 8 section 2 in the textbook.
8.5) Write the *equilibrium expression* for the following reaction:

$$HNO_3 (aq) + H_2O (l) \rightleftharpoons H_3O^+ (aq) + NO_3^- (aq)$$
8.5) Write the *equilibrium expression* for the following reaction:

$$\text{HNO}_3 (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{H}_3\text{O}^+ (aq) + \text{NO}_3^- (aq)$$

**HINT:**
Are the concentrations for liquid substances included in the equilibrium expression?

*For more help:* see [chapter 8 part 1 video](http://example.com) or chapter 8 section 2 in the textbook.
8.5) Write the \textit{equilibrium expression} for the following reaction:

\[ \text{HNO}_3 \text{ (aq)} + \text{H}_2\text{O} \text{ (l)} \rightleftharpoons \text{H}_3\text{O}^+ \text{ (aq)} + \text{NO}_3^- \text{ (aq)} \]

\[ K_{eq} = \frac{[\text{H}_3\text{O}^+][\text{NO}_3^-]}{[\text{HNO}_3]} \]

\textbf{EXPLANATION:}

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

When \textit{solids} (s) or \textit{liquids} (l) are present as reactants and/or products, they are \textit{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 [\text{H}_2\text{O}] does not appear in the equilibrium expression.
For the reaction shown below, predict whether the **reactants** or the **products** are *predominant* at equilibrium.

\[
\text{HCN (aq) + H}_2\text{O (l) } \rightleftharpoons \text{CN}^- (aq) + \text{H}_3\text{O}^+ (aq) \quad K_{eq} = 4.9 \times 10^{-10} \text{ M}
\]
8.6) For the reaction shown below, predict whether the **reactants** or the **products** are *predominant* at equilibrium.

\[ \text{HCN (aq) + H}_2\text{O (l)} \rightleftharpoons \text{CN}^- (aq) + \text{H}_3\text{O}^+ (aq) \quad K_{eq} = 4.9 \times 10^{-10} \text{ M} \]

**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.

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 \( (4.9 \times 10^{-10}) \).

**For more help:** see [chapter 8 part 2 video](#) or chapter 8 section 2 in the textbook.
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.

\[
\text{HCN (aq) + H}_2\text{O (l) } \rightleftharpoons \text{CN}^- (aq) + \text{H}_3\text{O}^+ (aq) \quad K_{eq} = 4.9 \times 10^{-10} \text{ M}
\]

**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.

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

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 \times 10^{-10} \text{ 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\(_3\)O\(^+\)) ions.

**For more details:** see chapter 8 part 2 video or chapter 8 section 2 in the textbook.
8.7) In the reaction below:

\[ \text{PCl}_3 (g) + \text{Cl}_2 (g) \rightleftharpoons \text{PCl}_5 (g) \]

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\(_3\)
In the reaction below:

\[ \text{PCl}_3 (g) + \text{Cl}_2 (g) \rightleftharpoons \text{PCl}_5 (g) \]

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\(_3\)

**HINT:**

How will these responses change the concentrations of products and reactants?
In the reaction below:

\[ \text{PCl}_3 (g) + \text{Cl}_2 (g) \rightleftharpoons \text{PCl}_5 (g) \]

Increasing the concentration of Cl\(_2\), according to Le Chatelier's principle, will _____ in order to re-establish equilibrium.

<table>
<thead>
<tr>
<th>Change Made to a Reaction that was at Equilibrium:</th>
<th>Response:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Increase the concentration of a <strong>reactant</strong>.</td>
<td>Rate of the <strong>forward</strong> reaction becomes greater than the rate of the reverse reaction until equilibrium is reestablished.</td>
</tr>
<tr>
<td>Increase the concentration of a <strong>product</strong>.</td>
<td>Rate of the <strong>reverse</strong> reaction becomes greater than the rate of the forward reaction until equilibrium is reestablished.</td>
</tr>
<tr>
<td>Decrease the concentration of a <strong>reactant</strong>.</td>
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<td>Decrease the concentration of a <strong>product</strong>.</td>
<td>Rate of the <strong>reverse</strong> reaction becomes less than the rate of the forward reaction until equilibrium is reestablished.</td>
</tr>
</tbody>
</table>

**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.
8.8)

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

\[ 2 \text{ H}_2\text{O} \ (l) \ \rightleftharpoons \ ? + \ ? \]

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

(iii) What is the **value** of the **equilibrium constant** \( K_w \) for this reaction.
8.8)

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

$$2 \text{H}_2\text{O (l)} \rightleftharpoons \square + \square$$

(ii) Write the equilibrium expression for this reaction.

HINT:

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

H$_2$O (l) does not appear in the equilibrium expression because it is a liquid.

This equilibrium expression is so commonly used, that the symbol “Kw” is used for equilibrium constant (instead of $K_{eq}$).

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

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

For more help: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.
8.8)

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

\[ 2 \text{H}_2\text{O} \, (l) \quad \rightleftharpoons \quad \text{OH}^- \, (aq) \quad + \quad \text{H}_3\text{O}^+ \, (aq) \]

(ii) Write the equilibrium expression for this reaction. \( \text{K}_w = [\text{OH}^-][\text{H}_3\text{O}^+] \)

H\(_2\)O \, (l) does not appear in the equilibrium expression because it is a liquid.

This equilibrium expression is so commonly used, that the symbol “\( \text{K}_w \)” is used for equilibrium constant (instead of \( \text{K}_\text{eq} \)).

(iii) What is the value of the equilibrium constant (\( \text{K}_w \)) for this reaction. \( \text{K}_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \, \text{M}^2 \)

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

The unit for \( \text{K}_w \) is \( \text{M}^2 \) because we are multiplying two molarity concentrations (\( \text{M} \times \text{M} = \text{M}^2 \)).

For more details: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.
8.9) What is the concentration of $[\text{H}_3\text{O}^+]$ in an aqueous solution when $[\text{OH}^-] = 1.8 \times 10^{-5} \text{ M}$?
8.9) What is the concentration of \([H_3O^+]\) in an aqueous solution when \([OH^-] = 1.8 \times 10^{-5} M\)? \textbf{ANSWER: } 5.6 \times 10^{-10} M

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

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

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

For more help: see chapter 8 part 3 video or chapter 8 section 3 in the textbook.
8.9) What is the concentration of $[\text{H}_3\text{O}^+]$ in an aqueous solution when $[\text{OH}^-] = 1.8 \times 10^{-5} \text{ M}$? **ANSWER:** $5.6 \times 10^{-10} \text{ M}$
8.9) What is the concentration of $[\text{H}_3\text{O}^+]$ in an aqueous solution when $[\text{OH}^-] = 1.8 \times 10^{-5} \text{ M}$?  

**ANSWER:** $5.6 \times 10^{-10} \text{ M}$

We know the equilibrium constant for the ionization of water: $K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2$

- That means whenever we know $[\text{OH}^-]$, we can calculate the concentration of $[\text{H}_3\text{O}^+]$.

  $$K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2$$

  $$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{OH}^-]}$$

- Likewise, whenever we know $[\text{H}_3\text{O}^+]$ we can calculate the concentration of $[\text{OH}^-]$.

  $$K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2$$

  $$[\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{H}_3\text{O}^+]}$$

**In this problem**, we know $[\text{OH}^-] = 1.8 \times 10^{-5} \text{ M}$, we can calculate the concentration of $[\text{H}_3\text{O}^+]$:

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14} \text{ M}^2}{1.8 \times 10^{-5} \text{ M}} = 5.6 \times 10^{-10} \text{ M}$$

**For more details:** see chapter 8 part 3 video or chapter 8 section 3 in the textbook.
8.10) What is the concentration of $[\text{OH}^-]$ in an aqueous solution when $[\text{H}_3\text{O}^+] = 9.3 \times 10^{-11} \text{ M}$?
8.10) What is the concentration of $[\text{OH}^-]$ in an aqueous solution when $[\text{H}_3\text{O}^+] = 9.3 \times 10^{-11} \text{ M}$? \textbf{ANSWER:} $1.1 \times 10^{-4} \text{ M}$

We know the equilibrium constant for the ionization of water: $K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2$

- That means whenever we know $[\text{OH}^-]$, we can calculate the concentration of $[\text{H}_3\text{O}^+]$.
- Likewise, whenever we know $[\text{H}_3\text{O}^+]$ we can calculate the concentration of $[\text{OH}^-]$.

To convert use: $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ M}^2$

\textbf{For more help:} see chapter 8 part 3 video or chapter 8 section 3 in the textbook.
8.10) What is the concentration of [OH\(^-\)] in an aqueous solution when \([H_3O^+] = 9.3 \times 10^{-11} \text{ M}\)?

**ANSWER:** \(1.1 \times 10^{-4} \text{ M}\)
What is the concentration of $[\text{OH}^-]$ in an aqueous solution when $[\text{H}_3\text{O}^+] = 9.3 \times 10^{-11}$ M? **ANSWER: $1.1 \times 10^{-4}$ M**

We know the equilibrium constant for the ionization of water: $K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14}$ M$^2$

- Whenever we know $[\text{H}_3\text{O}^+]$ we can calculate the concentration of $[\text{OH}^-]$.  

$$K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2$$

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{H}_3\text{O}^+]}$$

_In this problem_, we know $[\text{H}_3\text{O}^+] = 9.3 \times 10^{-11}$ M, we can calculate the concentration of $[\text{OH}^-]$:  

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14} \text{ M}^2}{9.3 \times 10^{-11} \text{ M}} = 1.1 \times 10^{-4} \text{ M}$$

_For more details:_ see chapter 8 part 3 video or chapter 8 section 3 in the textbook.
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) $\text{HPO}_4^-(aq) + \text{HNO}_3(aq) \rightleftharpoons \text{H}_2\text{PO}_4(aq) + \text{NO}_3^-(aq)$

b) $\text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_3^-(aq) + \text{H}_3\text{O}^+(aq)$

c) $\text{F}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HF}(aq) + \text{OH}^-(aq)$
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 as the base in each of the following reactions.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Acid</th>
<th>Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) ( \text{HPO}_4^- (aq) + \text{HNO}_3 (aq) \rightleftharpoons \text{H}_2\text{PO}_4 (aq) + \text{NO}_3^- (aq) )</td>
<td>base</td>
<td>acid</td>
</tr>
<tr>
<td><strong>HINT:</strong></td>
<td>HNO₃ donated an H⁺, so it acted as the acid. HPO₄⁻ accepted an H⁺, so it acted as the base.</td>
<td></td>
</tr>
<tr>
<td>b) ( \text{H}_2\text{CO}_3 (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{HCO}_3^- (aq) + \text{H}_3\text{O}^+ (aq) )</td>
<td></td>
<td></td>
</tr>
<tr>
<td>c) ( \text{F}^- (aq) + \text{H}_2\text{O} (l) \rightleftharpoons \text{HF} (aq) + \text{OH}^- (aq) )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

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.
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) \( \text{HPO}_4^{2-} (aq) + \text{HNO}_3 (aq) \Leftrightarrow \text{H}_2\text{PO}_4 (aq) + \text{NO}_3^{-} (aq) \)

   \( \text{HNO}_3 \) donated an H\(^+\), so it acted as the acid. \( \text{HPO}_4^{2-} \) accepted an H\(^+\), so it acted as the base.

b) \( \text{H}_2\text{CO}_3 (aq) + \text{H}_2\text{O} (l) \Leftrightarrow \text{HCO}_3^{-} (aq) + \text{H}_3\text{O}^+ (aq) \)

   \( \text{H}_2\text{CO}_3 \) donated an H\(^+\), so it acted as the acid. \( \text{H}_2\text{O} \) accepted an H\(^+\), so it acted as the base.

c) \( \text{F}^{-} (aq) + \text{H}_2\text{O} (l) \Leftrightarrow \text{HF} (aq) + \text{OH}^{-} (aq) \)

   \( \text{H}_2\text{O} \) donated an H\(^+\), so it acted as the acid. \( \text{F}^{-} \) 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.
8.12) Compounds that can act as acids or as bases are called ______________ compounds.

a) binary
b) amphoteric
c) acid-base
d) ionizing
8.12) Compounds that can act as acids or as bases are called ______________ compounds.

a) binary
b) amphoteric
c) acid-base
d) ionizing

HINT:

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
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 ($\text{HCO}_3^-$).

Bicarbonate acts as an acid in this reaction:

$$\text{HCO}_3^- + \text{CN}^- \rightleftharpoons \text{CO}_3^{2-} + \text{HCN}$$

Bicarbonate acts as a base in this reaction:

$$\text{HCO}_3^- + \text{HCl} \rightleftharpoons \text{H}_2\text{CO}_3 + \text{Cl}^-$$

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
8.13) Pairs of chemical species, such as HCl and Cl⁻ or H₃PO₃ and H₂PO₃⁻, 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
8.13) Pairs of chemical species, such as HCl and Cl\(^-\) or H\(_3\)PO\(_3\) and H\(_2\)PO\(_3\)^-, 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.
8.13) Pairs of chemical species, such as HCl and Cl− or H₃PO₃ and H₂PO₃⁻, 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.
8.14) For each conjugate pair, the species that contains the *extra* $\text{H}^+$ is called the “*acid form*,” and the species with *one fewer* $\text{H}^+$ is called the “*base form*.”

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

a) $\text{HBr}$ and $\text{Br}^-$

b) $\text{H}_3\text{O}^+$ and $\text{H}_2\text{O}$

c) $\text{HCO}_3^-$ and $\text{H}_2\text{CO}_3$

d) $\text{H}_2\text{PO}_3^-$ and $\text{H}_3\text{PO}_3$
For each conjugate pair, the species that contains the *extra* $\text{H}^+$ is called the “*acid form,*” and the species with *one fewer* $\text{H}^+$ is called the “*base form.*”

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

<table>
<thead>
<tr>
<th></th>
<th>acid form</th>
<th>base form</th>
</tr>
</thead>
<tbody>
<tr>
<td>a)</td>
<td>HBr</td>
<td>Br$^-$</td>
</tr>
<tr>
<td>b)</td>
<td>H$_3$O$^+$</td>
<td>H$_2$O</td>
</tr>
<tr>
<td>c)</td>
<td>HCO$_3^-$</td>
<td>H$_2$CO$_3$</td>
</tr>
<tr>
<td>d)</td>
<td>H$_2$PO$_3^-$</td>
<td>H$_3$PO$_3$</td>
</tr>
</tbody>
</table>

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
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\(_3\)O\(^+\) and H\(_2\)O

c) HCO\(_3\)^- and H\(_2\)CO\(_3\)

d) H\(_2\)PO\(_3\)^- and H\(_3\)PO\(_3\)

**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.
What is the acid form of NO$_3$?
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.
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.
8.16) What is the base form of $\text{H}_2\text{S}$?
8.16) What is the base form of $\text{H}_2\text{S}$?

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

**For more help:** see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
8.16) What is the **base form** of $\text{H}_2\text{S}$?

**Answer:** $\text{HS}^-$

**EXPLANATION:**

The **base form** of a conjugate pair contains *one fewer* $\text{H}^+$ than the **acid form**.

The **base form** is obtained by *removing* an $\text{H}^+$ from the **acid form**.

- Note that when an $\text{H}^+$ is *removed* from a species, its charge **decreases** by one charge unit. $\text{H}_2\text{S}$ has a 1+ charge, however when an $\text{H}^+$ is removed, it is converted to $\text{HS}^-$; the charge decreases by one charge unit.

**For more details:** see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
What is the acid form of NH₃?
8.17) What is the acid form of $\text{NH}_3$?

HINT:
The acid form of a conjugate pair contains one more $\text{H}^+$ than the “base form.”

For more help: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
What is the acid form of NH₃?

Answer: NH₄⁺

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. NH₃ has a zero charge, however when an H⁺ is added, it is converted to NH₄⁺; the charge increases by one charge unit.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
8.18) What is the base form of $\text{H}_3\text{O}^+$?
8.18) What is the **base form** of $\text{H}_3\text{O}^+$?

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

**For more help:** see [chapter 8 part 4 video](#) or chapter 8 section 4 in the textbook.
8.18) What is the base form of $\text{H}_3\text{O}^+$?

Answer: $\text{H}_2\text{O}$

EXPLANATION:

The base form of a conjugate pair contains one fewer $\text{H}^+$ than the acid form.

The base form is obtained by removing an $\text{H}^+$ from the acid form.

- Note that when an $\text{H}^+$ is removed from a species, its charge decreases by one charge unit. $\text{H}_3\text{O}^+$ has a 1+ charge, however when an $\text{H}^+$ is removed, it is converted to $\text{H}_2\text{O}$; the charge decreases by one charge unit.

For more details: see chapter 8 part 4 video or chapter 8 section 4 in the textbook.
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$_3$O$^+$ 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.
a) pH is most commonly defined as the "negative logarithm of the hydronium ion concentration."

b) The greater the concentration of \( \text{H}_3\text{O}^+ \) ions in solution, the greater the pH value.

HINT: Consider the effect of the negative sign in our definition of pH: \( \text{pH} = -\log[\text{H}_3\text{O}^+] \).

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: \([\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2\]

For more help: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.
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$_3$O$^+$ ions in solution, the greater the pH value. false
   • Because of the negative sign in our definition of pH (pH = -log[H$_3$O$^+$]), 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$_3$O$^+$] = 1.0 x 10$^{-14}$ M$^2$, whenever we know [OH$^-$], we can calculate [H$_3$O$^+$]. Once [H$_3$O$^+$] is known, the pH can be calculated using: pH = -log[H$_3$O$^+$].

For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.
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?
8.20)

i) What is the pH of an aqueous solution with \([\text{H}_3\text{O}^+] = 0.001\ \text{M}\)?

HINT:

\[ \text{pH} = -\log[\text{H}_3\text{O}^+] \]

ii) What is the pH of an aqueous solution with \([\text{H}_3\text{O}^+] = 4.6 \times 10^{-5}\ \text{M}\)?

For more help: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.
i) What is the pH of an aqueous solution with \([\text{H}_3\text{O}^+] = 0.001 \text{ M}\)?  
**ANSWER:** \(\text{pH} = 3.0\) 

*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.

ii) What is the pH of an aqueous solution with \([\text{H}_3\text{O}^+] = 4.6 \times 10^{-5} \text{ M}\)?  
**ANSWER:** \(\text{pH} = 4.34\) 

*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**
8.20)
i) What is the pH of an aqueous solution with \([H_3O^+] = 0.001\) M?  \(\text{ANSWER: } pH = 3.0\)

**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.

\[
pH = -\log[H_3O^+] = -\log[0.001] = -(3.0) = 3.0
\]

\[
\text{To convert use: } pH = -\log[H_3O^+]
\]

\[
\text{For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.}
\]

\[
\]

ii) What is the pH of an aqueous solution with \([H_3O^+] = 4.6 \times 10^{-5}\) M?  \(\text{ANSWER: } pH = 4.34\)

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

\[
pH = -\log[H_3O^+] = -\log[4.6 \times 10^{-5}] = -(4.34) = 4.34
\]
8.21) 

i) What is the pH of an aqueous solution with \([\text{OH}^-] = 0.005 \text{ M}\)?

ii) What is the pH of an aqueous solution with \([\text{OH}^-] = 7.2 \times 10^{-8} \text{ M}\)?
8.21)  

i) What is the pH of an aqueous solution with \([\text{OH}^-] = 0.005 \text{ M}\)?

**HINT:**

\[
[\text{OH}^-] \quad \text{To convert use:} \quad [\text{H}_3\text{O}^+]\cdot[\text{OH}^-] = 1.0 \times 10^{-14} \text{ M}^2 \\
[\text{H}_3\text{O}^+] \quad \text{To convert use:} \quad \text{pH} = -\log[\text{H}_3\text{O}^+] \\
\]  

_In these problems_, we know \([\text{OH}^-]\), so we can calculate the concentration of \([\text{H}_3\text{O}^+]\). Once the \([\text{H}_3\text{O}^+]\) is determined, it can be used to calculate the pH.

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

_For more help:_ see chapter 8 part 5 video or chapter 8 section 5 in the textbook.
8.21)
i) What is the pH of an aqueous solution with $[\text{OH}^-] = 0.005 \text{ M}$? \textbf{ANSWER: $pH = 11.7$} \textbf{Note:} There is \textit{one significant figure} in the concentration, so there will be \textit{one significant figure} in the pH.

\[ \text{ii) What is the pH of an aqueous solution with $[\text{OH}^-] = 7.2 \times 10^{-8} \text{ M}$? \textbf{ANSWER: $pH = 6.85$} \textbf{Note:} There are \textit{two significant figures} in the concentration, so there will be \textit{two significant figure} in the pH.} \]
8.21) 

i) What is the pH of an aqueous solution with \([\text{OH}^-] = 0.005\) M? \(\text{ANSWER: pH} = 11.7\)

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

In this problem, we know \([\text{OH}^-] = 0.005\) M, we can calculate the concentration of \([\text{H}_3\text{O}^+]\):

\[
[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14} \text{ M}^2}{0.005 \text{ M}} = 2 \times 10^{-12} \text{ M}
\]

Once the \([\text{H}_3\text{O}^+]\) is determined, it can be used to calculate the pH:

\[
\text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(2 \times 10^{-12}) = -(11.7) = 11.7
\]

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

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

\[
[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14} \text{ M}^2}{7.2 \times 10^{-8} \text{ M}} = 1.4 \times 10^{-7} \text{ M}
\]

\[
\text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(1.4 \times 10^{-7}) = -(6.85) = 6.85
\]

For more details: see chapter 8 part 5 video or chapter 8 section 5 in the textbook.
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?
i) What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with pH = 1.25?

\[ \text{pH} = - \log[\text{H}_3\text{O}^+] \]

\[ 10^{-\text{pH}} = [\text{H}_3\text{O}^+] \]

**HINT:**

To convert use:

\[ \text{pH} = - \log[\text{H}_3\text{O}^+] \]

or

\[ 10^{-\text{pH}} = [\text{H}_3\text{O}^+] \]

Multiply both sides by (-1)

Take the antilog of both sides

\[ 10^{(\text{pH})} = [\text{H}_3\text{O}^+] \]

\[ \text{pH} = - \log[\text{H}_3\text{O}^+] \]


ii) What is the $[\text{H}_3\text{O}^+]$ 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.
i) What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with pH = 1.25?  

ANswer: \(0.056\) 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 concentration units here are molar (M) or moles/L.

ii) What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with pH = 10.6?  

ANSWER: \(3 \times 10^{-11}\) M  

Note: There is one significant figure in the pH, so there will be one significant figure in the concentration.
i) What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with pH = 1.25?

**ANSWER:** 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.

\[
pH = -\log[\text{H}_3\text{O}^+]
\]
\[
1.25 = -\log[\text{H}_3\text{O}^+]
\]
\[
-1.25 = \log[\text{H}_3\text{O}^+]
\]
\[
10^{-1.25} = [\text{H}_3\text{O}^+]
\]
\[
[\text{H}_3\text{O}^+] = 0.056 \text{ M}
\]

ii) What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with pH = 10.6?

**ANSWER:** $3 \times 10^{-11}$ M

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

\[
pH = -\log[\text{H}_3\text{O}^+]
\]
\[
10.6 = -\log[\text{H}_3\text{O}^+]
\]
\[
-10.6 = \log[\text{H}_3\text{O}^+]
\]
\[
10^{-10.6} = [\text{H}_3\text{O}^+]
\]
\[
[\text{H}_3\text{O}^+] = 3 \times 10^{-11} \text{ M}
\]

*For more details:* see chapter 8 part 5 video or chapter 8 section 5 in the textbook.
8.23) Solutions are characterized as acidic, basic, or neutral by the relative amounts of H$_3$O$^+$ and OH$^-$ that are present.

i) Solutions that contain more H$_3$O$^+$ than OH$^-$ are called ___________ solutions.
   a) acidic
   b) neutral
   c) basic
   d) buffer

ii) Solutions that contain more OH$^-$ than H$_3$O$^+$ are called ___________ solutions.
   a) acidic
   b) neutral
   c) basic
   d) buffer

iii) Solutions that contain equal concentrations of H$_3$O$^+$ and OH$^-$ are called _____________ solutions.
   a) balanced
   b) neutral
   c) pH
   d) buffer
8.23) Solutions are characterized as acidic, basic, or neutral by the relative amounts of $\text{H}_3\text{O}^+$ and $\text{OH}^-$ that are present.

i) Solutions that contain more $\text{H}_3\text{O}^+$ than $\text{OH}^-$ are called __________ solutions.

HINT: 

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

ii) Solutions that contain more $\text{OH}^-$ than $\text{H}_3\text{O}^+$ are called __________ solutions.

HINT: 

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

iii) Solutions that contain equal concentrations of $\text{H}_3\text{O}^+$ and $\text{OH}^-$ 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.
Solutions are characterized as acidic, basic, or neutral by the relative amounts of $\text{H}_3\text{O}^+$ and $\text{OH}^-$ that are present.

i) Solutions that contain more $\text{H}_3\text{O}^+$ than $\text{OH}^-$ 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 $\text{OH}^-$ than $\text{H}_3\text{O}^+$ are called __________ solutions.
   a) acidic  
   b) neutral  
   c) basic  
   d) buffer

iii) Solutions that contain equal concentrations of $\text{H}_3\text{O}^+$ and $\text{OH}^-$ are called __________ solutions.
   a) balanced  
   b) neutral  
   c) pH  
   d) buffer

<table>
<thead>
<tr>
<th>Solution Characterization</th>
<th>pH</th>
<th>$[\text{H}_3\text{O}^+]$</th>
<th>$[\text{OH}^-]$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acidic</td>
<td>less than 7.00</td>
<td>greater than 1.0 x 10^{-7} M</td>
<td>less than 1.0 x 10^{-7} M</td>
</tr>
<tr>
<td>Neutral</td>
<td>7.00</td>
<td>1.0 x 10^{-7} M</td>
<td>1.0 x 10^{-7} M</td>
</tr>
<tr>
<td>Basic</td>
<td>greater than 7.00</td>
<td>less than 1.0 x 10^{-7} M</td>
<td>greater than 1.0 x 10^{-7} M</td>
</tr>
</tbody>
</table>

For more details: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.
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
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.
Compounds are characterized as ________________________ depending on whether they donate or accept H\(^+\) in a particular acid-base reaction.

a) neutral or non-neutral

Solutions (not compounds) that contain equal concentrations of H\(_3\)O\(^+\) and OH\(^-\) are characterized as neutral.

b) acidic or basic

Solutions (not compounds) are characterized as acidic, basic, or neutral by the relative amounts of H\(_3\)O\(^+\) 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.
8.25) For each of the following, write whether the solution condition describes an **acidic**, basic, or neutral solution.

a) \( \text{pH} = 3.9 \)

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

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

d) \( [OH^-] > [H_3O^+] \)

e) \( [H_3O^+] > [OH^-] \)

f) \( \text{pH} = 9.7 \)

g) \( \text{pH} = 2.0 \)

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

i) \( \text{pH} = 12.6 \)

j) \( [H_3O^+] = [OH^-] \)

k) \( \text{pH} = 7.00 \)

l) \( [H_3O^+] = 1.0 \times 10^{-7} \text{ M} \)
8.25) For each of the following, write whether the solution condition describes an **acidic**, **basic**, or **neutral** solution.

<table>
<thead>
<tr>
<th>Condition</th>
<th>Solution Characterization</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) pH = 3.9</td>
<td>acidic</td>
</tr>
<tr>
<td>b) [H₃O⁺] = 1 x 10⁻⁵ M</td>
<td>acidic</td>
</tr>
<tr>
<td>c) [OH⁻] = 1 x 10⁻⁷ M</td>
<td>neutral</td>
</tr>
<tr>
<td>d) [OH⁻] &gt; [H₃O⁺]</td>
<td>basic</td>
</tr>
<tr>
<td>e) [H₃O⁺] &gt; [OH⁻]</td>
<td>acidic</td>
</tr>
<tr>
<td>f) pH = 9.7</td>
<td>acidic</td>
</tr>
<tr>
<td>g) pH = 2.0</td>
<td>neutral</td>
</tr>
<tr>
<td>h) [OH⁻] = 6.8 x 10⁻⁸ M</td>
<td>basic</td>
</tr>
<tr>
<td>i) pH = 12.6</td>
<td>basic</td>
</tr>
<tr>
<td>j) [H₃O⁺] = [OH⁻]</td>
<td>neutral</td>
</tr>
<tr>
<td>k) pH = 7.00</td>
<td>neutral</td>
</tr>
<tr>
<td>l) [H₃O⁺] = 1.0 x 10⁻⁷ M</td>
<td>basic</td>
</tr>
</tbody>
</table>

**HINT:**

For more help: see chapter 8 part 6 video or chapter 8 section 6 in the textbook.
For each of the following, write whether the solution condition describes an **acidic**, **basic**, or **neutral** solution.

a) pH = 3.9 **acidic**

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

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

d) \([\text{OH}^-] > [H_3O^+] \) **basic**

e) \([H_3O^+] > [\text{OH}^-] \) **acidic**

f) pH = 9.7 **basic**

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

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

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

\[ [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M} \] **neutral**

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

\[ [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M} \] **neutral**

**EXPLANATION:**

<table>
<thead>
<tr>
<th>Solution Characterization</th>
<th>pH</th>
<th>([H_3O^+])</th>
<th>([\text{OH}^-])</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acidic</td>
<td>less than 7.00</td>
<td>greater than 1.0 \times 10^{-7} \text{ M}</td>
<td>less than 1.0 \times 10^{-7} \text{ M}</td>
</tr>
<tr>
<td>Neutral</td>
<td>7.00</td>
<td>1.0 \times 10^{-7} \text{ M}</td>
<td>1.0 \times 10^{-7} \text{ M}</td>
</tr>
<tr>
<td>Basic</td>
<td>greater than 7.00</td>
<td>less than 1.0 \times 10^{-7} \text{ M}</td>
<td>greater than 1.0 \times 10^{-7} \text{ M}</td>
</tr>
</tbody>
</table>

**For more details:** see chapter 8 part 6 video or chapter 8 section 6 in the textbook.
8.26) Use the table to determine which is a **stronger acid**, boric acid or acetic acid.

<table>
<thead>
<tr>
<th>Acid Name</th>
<th>Acid Formula</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Perchloric acid</td>
<td>HClO₄</td>
<td>$1 \times 10^9$ M (estimated)</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>$1 \times 10^7$ M (estimated)</td>
</tr>
<tr>
<td>Chloric acid</td>
<td>HClO₃</td>
<td>$1 \times 10^3$ M (estimated)</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>$7.5 \times 10^{-3}$ M</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>HF</td>
<td>$6.6 \times 10^{-4}$ M</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH₃CO₂H</td>
<td>$1.8 \times 10^{-5}$ M</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>H₂CO₃</td>
<td>$4.4 \times 10^{-7}$ M</td>
</tr>
<tr>
<td>Dihydrogen phosphate ion</td>
<td>H₂PO₄⁻</td>
<td>$6.2 \times 10^{-8}$ M</td>
</tr>
<tr>
<td>Boric acid</td>
<td>H₃BO₃</td>
<td>$5.7 \times 10^{-10}$ M</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH₄⁺</td>
<td>$5.6 \times 10^{-10}$ M</td>
</tr>
<tr>
<td>Hydrocyanic acid</td>
<td>HCN</td>
<td>$4.9 \times 10^{-10}$ M</td>
</tr>
<tr>
<td>Bicarbonate ion</td>
<td>HCO₃⁻</td>
<td>$5.6 \times 10^{-11}$ M</td>
</tr>
<tr>
<td>Methylammonium ion</td>
<td>CH₃NH₃⁺</td>
<td>$2.4 \times 10^{-11}$ M</td>
</tr>
<tr>
<td>Hydrogen phosphate ion</td>
<td>HPO₄⁻</td>
<td>$4.2 \times 10^{-13}$ M</td>
</tr>
</tbody>
</table>
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.

### Various Acids and Their Acidity Constants

<table>
<thead>
<tr>
<th>Acid Name</th>
<th>Acid Formula</th>
<th>$K_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Perchloric acid</td>
<td>HClO₄</td>
<td>$1 \times 10^9$ M (estimated)</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>$1 \times 10^7$ M (estimated)</td>
</tr>
<tr>
<td>Chloric acid</td>
<td>HClO₃</td>
<td>$1 \times 10^8$ M (estimated)</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>$7.5 \times 10^{-3}$ M</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>HF</td>
<td>$6.6 \times 10^{-4}$ M</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH₃CO₂H</td>
<td>$1.8 \times 10^{-5}$ M</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>H₂CO₃</td>
<td>$4.4 \times 10^{-7}$ M</td>
</tr>
<tr>
<td>Dihydrogen phosphate ion</td>
<td>H₂PO₄⁻</td>
<td>$6.2 \times 10^{-8}$ M</td>
</tr>
<tr>
<td>Boric acid</td>
<td>H₃BO₃</td>
<td>$5.7 \times 10^{-10}$ M</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH₄⁺</td>
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<td>Hydrocyanic acid</td>
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</tr>
<tr>
<td>Hydrogen phosphate ion</td>
<td>HPO₄⁻</td>
<td>$4.2 \times 10^{-13}$ M</td>
</tr>
</tbody>
</table>
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 $H_3O^+$.

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

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

\[
\text{HBr (aq)} + \text{NaOH (aq)} \rightleftharpoons \text{H}_2\text{O (l)} + \text{NaBr (aq)}
\]

In **neutralization reactions**, the H\(^+\) from the **acid** bonds to the OH\(^-\) to produce H\(_2\)O. 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**:

\[
\text{HNO}_3 (aq) + \text{NaOH (aq)} \rightleftharpoons \text{_____________} + \text{_____________}
\]

\[
\text{HNO}_3 (aq) + \text{KOH (aq)} \rightleftharpoons \text{_____________} + \text{_____________}
\]
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:

\[
HBr (aq) + NaOH (aq) \rightleftharpoons H_2O (l) + NaBr (aq)
\]

In neutralization reactions, the H\(^+\) from the acid bonds to the OH\(^-\) to produce H\(_2\)O. 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:

\[
\text{HINT: } \text{HNO}_3 (aq) + \text{NaOH (aq)} \rightleftharpoons \_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_\_\_
\]

The H\(^+\) from the acid bonds to the OH\(^-\) to produce H\(_2\)O. 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.

\[
\text{HINT: } \text{HNO}_3 (aq) + \text{KOH (aq)} \rightleftharpoons \_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_\_\_
\]

The H\(^+\) from the acid bonds to the OH\(^-\) to produce H\(_2\)O. 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.
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:

\[
\text{HBr (aq)} + \text{NaOH (aq)} \rightleftharpoons \text{H}_2\text{O (l)} + \text{NaBr (aq)}
\]

In neutralization reactions, the H\(^+\) from the acid bonds to the OH\(^-\) to produce H\(_2\)O. 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:

\[
\text{HNO}_3 (aq) + \text{NaOH (aq)} \rightleftharpoons \_\text{H}_2\text{O (l)} + \_\text{NaNO}_3 (aq)
\]

The H\(^+\) from the acid bonds to the OH\(^-\) to produce H\(_2\)O. 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\(_3\)).

\[
\text{HNO}_3 (aq) + \text{KOH (aq)} \rightleftharpoons \_\text{H}_2\text{O (l)} + \_\text{KNO}_3 (aq)
\]

The H\(^+\) from the acid bonds to the OH\(^-\) to produce H\(_2\)O. 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.
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:

\[
\text{HA (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{A}^-\text{(aq)} + \text{H}_3\text{O}^+(aq)
\]

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.

\[
\text{HCN (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{CN}^-\text{(aq)} + \text{H}_3\text{O}^+(aq)
\]
The general form of a chemical equation for an acid reacting with water to produce its base form and hydronium can be written as:

\[ \text{HA (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{A}^- (aq) + \text{H}_3\text{O}^+ (aq) \]

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.

\[ \text{HCN (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{CN}^- (aq) + \text{H}_3\text{O}^+ (aq) \]

**HINT:** Compare the pH to the pK\(_a\).
The general form of a chemical equation for an acid reacting with water to produce its base form and hydronium can be written as:

\[
HA (aq) + H_2O (l) \rightleftharpoons A^- (aq) + H_3O^+ (aq)
\]

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.

\[
HCN (aq) + H_2O (l) \rightleftharpoons CN^- (aq) + H_3O^+ (aq)
\]

**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.

For more details: see chapter 8 part 7 video or chapter 8 section 7 in the textbook.
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) CH₃NH₃⁺/CH₃NH₂ at pH = 7.0

c) H₂CO₃/HCO₃⁻ at pH = 8.5
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) CH₃NH₃⁺/CH₃NH₂ at pH = 7.0

c) H₂CO₃/HCO₃⁻ at pH = 8.5

### HINT: Compare the pH to the pKₐ

<table>
<thead>
<tr>
<th>Solution Condition</th>
<th>Relative Amounts of Acid and Base Forms</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH &lt; pKₐ</td>
<td>[HA] &gt; [A⁻]</td>
</tr>
<tr>
<td>pH &gt; pKₐ</td>
<td>[A⁻] &gt; [HA]</td>
</tr>
<tr>
<td>pH = pKₐ</td>
<td>[HA] = [A⁻]</td>
</tr>
</tbody>
</table>

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

### Table of Acid Constants

<table>
<thead>
<tr>
<th>Acid Name</th>
<th>Acid Formula</th>
<th>Kₐ (M)</th>
<th>pKₐ</th>
</tr>
</thead>
<tbody>
<tr>
<td>Perchloric acid</td>
<td>HClO₄</td>
<td>1 x 10⁹</td>
<td>-9.0</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>1 x 10⁻⁷</td>
<td>-7.0</td>
</tr>
<tr>
<td>Chloric acid</td>
<td>HClO₃</td>
<td>1 x 10⁻³</td>
<td>-3.0</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>7.5 x 10⁻³</td>
<td>2.12</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>HF</td>
<td>6.6 x 10⁻⁴</td>
<td>3.18</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH₃CO₂H</td>
<td>1.8 x 10⁻⁵</td>
<td>4.74</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>H₂CO₃</td>
<td>4.4 x 10⁻⁷</td>
<td>6.36</td>
</tr>
<tr>
<td>Dihydrogen phosphate ion</td>
<td>H₂PO₄</td>
<td>6.2 x 10⁻⁸</td>
<td>7.21</td>
</tr>
<tr>
<td>Boric acid</td>
<td>H₃BO₃</td>
<td>5.7 x 10⁻¹⁰</td>
<td>9.24</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH₄⁺</td>
<td>5.6 x 10⁻¹⁰</td>
<td>9.25</td>
</tr>
<tr>
<td>Hydrocyanic acid</td>
<td>HCN</td>
<td>4.9 x 10⁻¹⁰</td>
<td>9.31</td>
</tr>
<tr>
<td>Bicarbonate ion</td>
<td>HCO₃⁻</td>
<td>5.6 x 10⁻¹¹</td>
<td>10.25</td>
</tr>
<tr>
<td>Methylammonium ion</td>
<td>CH₃NH₃⁺</td>
<td>2.4 x 10⁻¹¹</td>
<td>10.62</td>
</tr>
<tr>
<td>Hydrogen phosphate ion</td>
<td>HPO₄²⁻</td>
<td>4.2 x 10⁻¹³</td>
<td>12.38</td>
</tr>
</tbody>
</table>
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)  $\text{HF/F}^-$ at pH = 9.7  *base form* ($\text{F}^-$)
   \[ \text{pH} > \text{pK}_a \]

b)  $\text{CH}_3\text{NH}_3^+/\text{CH}_3\text{NH}_2$ at pH = 7.0  *acid form* ($\text{CH}_3\text{NH}_3^+$)
   \[ \text{pH} < \text{pK}_a \]

c)  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ at pH = 8.5  *base form* ($\text{HCO}_3^-$)
   \[ \text{pH} > \text{pK}_a \]

**EXPLANATION:**
Compare the pH to the pK\(_a\).

<table>
<thead>
<tr>
<th>Solution Condition</th>
<th>Relative Amounts of Acid and Base Forms</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH &lt; pK(_a)</td>
<td>[HA] &gt; [A(^-)]</td>
</tr>
<tr>
<td>pH &gt; pK(_a)</td>
<td>[A(^-)] &gt; [HA]</td>
</tr>
<tr>
<td>pH = pK(_a)</td>
<td>[HA] = [A(^-)]</td>
</tr>
</tbody>
</table>

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

• 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.  

• 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.
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 (H₂CO₃)/bicarbonate (HCO₃⁻)

c) chloric acid (HClO₃)/chlorate (ClO₃⁻)

d) ammonium (NH₄⁺)/ammonia (NH₃)
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 (H₂CO₃)/bicarbonate (HCO₃⁻)
- c) chloric acid (HClO₃)/chlorate (ClO₃⁻)
- d) ammonium (NH₄⁺)/ammonia (NH₃)

**HINT:**

For more help: see [chapter 8 part 8 video](#) or chapter 8 section 7 in the textbook.
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 (H₂CO₃)/bicarbonate (HCO₃⁻)

c) chloric acid (HClO₃)/chlorate (ClO₃⁻)

d) ammonium (NH₄⁺)/ammonia (NH₃)

EXPLANATION:
Important extracellular (outside of cells) buffers, in solutions such as blood or interstitial fluids, are the carbonic acid (H₂CO₃)/bicarbonate (HCO₃⁻) and the ammonium (NH₄⁺)/ammonia (NH₃) conjugate pairs.

- In blood, the carbonic acid (H₂CO₃)/bicarbonate (HCO₃⁻) buffering pair is especially useful because the buffer conjugate pair concentrations ([H₂CO₃] and [HCO₃⁻]) 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.