How many significant figures in the answer?

\[ \log (7.310 \times 10^{23}) = \]

1. 23.
2. 23.9
3. 23.86
4. 23.8639
How many significant figures in the answer?

\[
\log (7.310 \times 10^{23}) =
\]

1. 23.
2. 23.9
3. 23.86
4. 23.8639

4% 19% 11% 66%
For: $\text{NH}_3 (\text{aq}) + \text{H}_2\text{O (l)} \rightleftharpoons \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq})$

Fill in the chart below:

<table>
<thead>
<tr>
<th></th>
<th>$\text{NH}_3 (\text{aq})$</th>
<th>$\text{NH}_4^+ (\text{aq})$</th>
<th>$+$</th>
<th>$\text{OH}^- (\text{aq})$</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.</td>
<td>initial molarity</td>
<td>0.15</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td></td>
<td>change in molarity</td>
<td>$+x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td></td>
<td>equilibrium molarity</td>
<td>$0.15+x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td>2.</td>
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<td>$+x$</td>
</tr>
<tr>
<td></td>
<td>equilibrium molarity</td>
<td>$0.15$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
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<td>initial molarity</td>
<td>0.15</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td></td>
<td>change in molarity</td>
<td>$-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
<tr>
<td></td>
<td>equilibrium molarity</td>
<td>$0.15-x$</td>
<td>$+x$</td>
<td>$+x$</td>
</tr>
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For: $\text{NH}_3 \text{ (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{NH}_4^+ \text{ (aq)} + \text{OH}^- \text{ (aq)}$

Fill in the chart below:

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<th>OH$^-$ (aq)</th>
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<tr>
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<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change in molarity</td>
<td>+x</td>
<td>+x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>equilibrium molarity</td>
<td>0.15+x</td>
<td>+x</td>
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<tr>
<td>2. initial molarity</td>
<td>0.15</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change in molarity</td>
<td>0</td>
<td>+x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>equilibrium molarity</td>
<td>0.15</td>
<td>+x</td>
<td>+x</td>
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<tr>
<td>3. initial molarity</td>
<td>0.15</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>change in molarity</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
<td></td>
</tr>
<tr>
<td>equilibrium molarity</td>
<td>0.15-x</td>
<td>+x</td>
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When asked to “check assumption,” what do you do?

1. \[
\frac{0.00164}{0.15} \times 100 \% = 1.1 \% \quad (1.1\% \text{ is less than } 5\%, \text{ assumption is okay})
\]

2. \[
0.15 - 0.00164 = 0.14836 \quad (\text{within sig figs, } 0.00164 \text{ is small compared to } 0.15)
\]

3. \[
0.00164 \times 0.15 \times 100 \% = 0.0246 \% \quad (0.0246 \% \text{ is less than } 5\%, \text{ assumption is okay})
\]

4. \[
(0.15 - 0.00164) \times 100 \% = 14.8 \% \quad (14.8 \text{ is greater than } 5\%, \text{ assumption is not okay})
\]
When asked to “check assumption,” what do you do?

1. \( \frac{0.00164}{0.15} \times 100\% = 1.1\% \) (1.1% is less than 5%, assumption is okay)

2. \( 0.15 - 0.00164 = 0.14836 \) (within sig figs, 0.00164 is small compared to 0.15)

3. \( 0.00164 \times 0.15 \times 100\% = 0.0246\% \) (0.0246% is less than 5%, assumption is okay)

4. \( (0.15 - 0.00164) \times 100\% = 14.8\% \) (14.8 is greater than 5%, assumption is not okay)
Predict whether the pH is acidic, neutral, or basic for a solution of NaCH₃COO(aq).

- $K_a$ of CH₃COOH is $1.76 \times 10^{-5}$.

1. acidic
2. neutral
3. Basic
4. Not enough information
Predict whether the pH is acidic, neutral, or basic for a solution of $\text{NaCH}_3\text{COO}(\text{aq})$.

$K_a$ of $\text{CH}_3\text{COOH}$ is $1.76 \times 10^{-5}$.

1. acidic
2. neutral
3. Basic
4. Not enough information
A strong acid and the salt of its conjugate base don’t make a good buffer. Why?

1. The conjugate base of a strong acid is a **weak** base, and a weak base can’t neutralize added acid so pH is not maintained.

2. The conjugate base of a strong acid is **ineffective** as a base, and an ineffective base can’t neutralize added acid so pH is not maintained.

3. The conjugate base of a strong acid is a **strong** base, and a strong base changes the pH.
A strong acid and the salt of its conjugate base don’t make a good buffer. Why?

1. The conjugate base of a strong acid is a weak base, and a weak base can’t neutralize added acid so pH is not maintained.

2. The conjugate base of a strong acid is ineffective as a base, and an ineffective base can’t neutralize added acid so pH is not maintained.

3. The conjugate base of a strong acid is a strong base, and a strong base changes the pH.
Which is the correct simplified expression for $K_a$ following application of the assumption that $x$ is small compared to 1.00 and 0.500?

1. $K_a = 0.500/1.00$
2. $K_a = (x)(0.500+x)/1.00$
3. $K_a = x^2/1.00$
4. $K_a = (x)(0.500)/1.00$
5. $K_a = (x)(0.500)/1.00-x$
6. $K_a = x/1.00$
Which is the correct simplified expression for $K_a$ following application of the assumption that $x$ is small compared to 1.00 and 0.500?

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4. $K_a = (x)(0.500)/1.00$ \[\checkmark\]
5. $K_a = (x)(0.500)/1.00-x$
6. $K_a = x /1.00$
Which of the following $K_a$ expressions is correct following the addition of 0.100 mol of HCl?

1. $K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$
   
   $K_a = \frac{(0.400 + x)(x)}{(1.10 - x)}$

2. $K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH][H_2O]}$
   
   $K_a = \frac{(0.400 + x)(x)}{(1.10 - x)}$

3. $K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$
   
   $K_a = \frac{x^2}{(1.10 - x)}$

4. $K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$
   
   $K_a = \frac{(0.500 + x)(x)}{(1.00 - x)}$
Which of the following $K_a$ expressions is correct following the addition of 0.100 mol of HCl?

1. $K_a = [H_3O^+][HCOO^-]/[HCOOH]$
   $K_a = (0.400 + x)(x) / (1.10 - x)$

2. $K_a = [H_3O^+][HCOO^-]/[HCOOH][H_2O]$
   $K_a = (0.400 + x)(x) / (1.10 - x)$

3. $K_a = [H_3O^+][HCOO^-]/[HCOOH]$
   $K_a = x^2 / (1.10 - x)$

4. $K_a = [H_3O^+][HCOO^-]/[HCOOH]$
   $K_a = (0.500 + x)(x) / (1.00 - x)$