This print-out should have 37 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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001 10.0 points

Calculate the equilibrium constant at 25°C for a reaction for which $\Delta G^0 = -4.85$ kcal/mol.

1. 3592.86 **correct**

2. 7185.72

3. 1796.43

4. -3592.86

5. 35928.6

Explanation:
$T = 25°\text{C} + 273 = 298$ K
$\Delta G^0 = -4850$ cal/mol
At equilibrium

$$\begin{aligned}\Delta G^0 &= -RT\,\ln K\\ -4850 &= (-1.987\ \text{cal/mol}\cdot\text{K})\\ &\quad\times(298.15\ \text{K})(\ln K)\\ K &= 3592.86\end{aligned}$$

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002 10.0 points

The standard molar Gibbs free energy of formation of NO_2 (g) at 298 K is 51.30 $\text{kJ}\cdot\text{mol}^{-1}$ and that of N_2O_4 (g) is 97.82 $\text{kJ}\cdot\text{mol}^{-1}$. What is the equilibrium constant at 25°C for the reaction

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

1. None of these

2. 0.657

3. 9.72×10^{9}

4. 7.01×10^{-9}

5. 1.02×10^{-10}

6. 0.145

7. 1.00

8. 6.88 **correct**

Explanation:
$\Delta G^0_{\text{products}} = 97.82\ \text{kJ}\cdot\text{mol}^{-1}$
$\Delta G^0_{\text{reactants}} = 51.30\ \text{kJ}\cdot\text{mol}^{-1}$

$$\begin{aligned}\Delta G^0_{\text{rxn}} &= \sum n\,\Delta G^0_{\text{products}}\\ &\quad-\sum n\,\Delta G^0_{\text{reactants}}\\ &= 97.82 - (2)(51.30)\\ &= (-4.78\ \text{kJ/mol})\left(\frac{1000\ \text{J}}{\text{kJ}}\right)\\ &= -4780\ \text{J/mol}\end{aligned}$$

$$\begin{aligned}\Delta G^0 &= -RT\,\ln K\\ K &= e^{-\Delta G^0/(RT)}\\ &= \exp\left[-\frac{-4780\ \text{J/mol}}{(8.3145\ \text{J/mol}\cdot\text{K})(298\ \text{K})}\right]\\ &= 6.88395\end{aligned}$$

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003 10.0 points

The reaction

$$A + B \rightleftharpoons C + 2\,D$$

has an equilibrium constant of 3.7×10^{-3}. Consider a reaction mixture with

$[A] = 2.0\times10^{-2}$ M $[C] = 2.4\times10^{-6}$ M
$[B] = 1.7\times10^{-4}$ M $[D] = 3.5\times10^{-3}$ M

Which of the following statements is definitely true?

1. The forward reaction can occur to a greater extent than the reverse reaction until equilibrium is established. **correct**

2. Heat will be evolved.

3. No conclusions about the system can be made without additional information.

4. The system is at equilibrium.

5. The reverse reaction can occur to a greater extent than the forward reaction until equilibrium is established.

Explanation:

$$Q = \frac{[\mathrm{C}]\,[\mathrm{D}]^2}{[\mathrm{A}]\,[\mathrm{B}]} = \frac{(2.4 \times 10^{-6}\ \mathrm{M})\,(0.0035\ \mathrm{M})^2}{(0.02\ \mathrm{M})\,(0.00017\ \mathrm{M})}$$
$$= 8.64706 \times 10^{-6}$$

Since $Q < K$ the foward reaction is favored.

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004 10.0 points

The reaction

$$\mathrm{N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)}\,,$$

has an equilibrium constant of 4.0×10^8 at 25°C. What will eventually happen if 44.0 moles of NH_3, 0.452 moles of N_2, and 0.108 moles of H_2 are put in a 10.0 liter container at 25°C?

1. More NH_3 will be formed. **correct**

2. More N_2 and H_2 will be formed.

3. Nothing; the system is at equilibrium.

Explanation:

$K = 4.0 \times 10^8$ $\qquad [\mathrm{NH_3}] = \dfrac{44.0\ \mathrm{mol}}{10\ \mathrm{L}}$

$[\mathrm{N_2}] = \dfrac{0.452\ \mathrm{mol}}{10\ \mathrm{L}}$ $\qquad [\mathrm{H_2}] = \dfrac{0.108\ \mathrm{mol}}{10\ \mathrm{L}}$

$$Q = \frac{[\mathrm{NH_3}]^2}{[\mathrm{N_2}]\,[\mathrm{H_2}]^3} = \frac{(4.40\ \mathrm{M})^2}{(0.0452\ \mathrm{M})\,(0.0108\ \mathrm{M})^3} = 3.4 \times 10^8$$

Since $Q < K$ equilibrium will shift to the right, forming more NH_3.

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005 10.0 points

$K_c = 2.6 \times 10^8$ at 825 K for the reaction

$$\mathrm{2\,H_2(g) + S_2(g) \rightleftharpoons 2\,H_2S(g)}$$

The equilibrium concentration of H_2 is 0.0020 M and that of S_2 is 0.0010 M. What is the equilibrium concentration of H_2S?

1. 10 M

2. 1.02 M **correct**

3. 0.10 M

4. 0.0010 M

Explanation:

$K_c = 2.6 \times 10^8$ $\qquad [\mathrm{H_2}]_{\mathrm{eq}} = 0.0020\ \mathrm{M}$

$[\mathrm{S_2}]_{\mathrm{eq}} = 0.0010\ \mathrm{M}$

$$\mathrm{2\,H_2(g) + S_2 \rightleftharpoons 2\,H_2S}$$

$$K_c = \frac{[\mathrm{H_2S}]^2}{[\mathrm{H_2}]^2\,[\mathrm{S_2}]}$$
$$[\mathrm{H_2S}] = \sqrt{K_c\,[\mathrm{H_2}]^2\,[\mathrm{S_2}]}$$
$$= \sqrt{(2.6 \times 10^8)\,(0.0020\ \mathrm{M})^2\,(0.0010\ \mathrm{M})}$$
$$= 1.0\ \mathrm{M}$$

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006 10.0 points

Suppose the reaction

$$\mathrm{H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)}$$

has an equilibrium constant $K_c = 49$ and the initial concentration of H_2 and I_2 is 0.5 M and HI is 0.0 M. Which of the following is the correct value for the final concentration of HI(g)?

1. 0.389 M

2. 0.219 M

3. 0.778 M **correct**

4. 0.250 M

5. 0.599 M

Explanation:

$K_c = 49$ $\qquad [\mathrm{H_2}]_{\mathrm{ini}} = 0.5\ \mathrm{M}$

$[\mathrm{I_2}]_{\mathrm{ini}} = 0.5\ \mathrm{M}$ $\qquad [\mathrm{HI}]_{\mathrm{ini}} = 0\ \mathrm{M}$

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons 2\,HI(g)$
Ini, M	0.5		0.5	0
Δ, M	$-x$		$-x$	$+2\,x$
Equil, M	$0.5 - x$		$0.5 - x$	$2\,x$

$$\begin{aligned} K_c &= \frac{[HI]^2}{[H_2]\,[I_2]} \\ 49 &= \frac{(2x)^2}{(0.5-x)^2} \\ 7 &= \frac{2x}{0.5-x} \\ 7(0.5-x) &= 2x \\ 3.5 - 7x &= 2x \\ 3.5 &= 9x \\ x &= \frac{3.5}{9} = 0.389\text{ M} \end{aligned}$$

Looking back at our equilibrium values, we see that the final concentration of HI is equal to $2x$, so $2(0.389) = 0.778$ M.

007 10.0 points

The system

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

is at equilibrium at a fixed temperature with a partial pressure of H_2 of 0.200 atm, a partial pressure of I_2 of 0.200 atm, and a partial pressure of HI of 0.100 atm. An additional 0.32 atm pressure of HI is admitted to the container, and it is allowed to come to equilibrium again. What is the new partial pressure of HI?

Correct answer: 0.164 atm.

Explanation:

$P_{I_2} = P_{H_2} = 0.2$ atm $\qquad P_{HI} = 0.1$ atm

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

$$\begin{aligned} K_p &= \frac{(P_{HI})^2}{P_{H_2} \cdot P_{I_2}} \\ &= \frac{(0.1\text{ atm})^2}{(0.2\text{ atm})\,(0.2\text{ atm})} = 0.25 \end{aligned}$$

new $P_{HI} = (0.1 + 0.32)$ atm $= 0.42$ atm

Adding the products shifts the equilibrium to the left.

	$H_2(g)$	+	$I_2(g)$	$\rightleftharpoons$	$2\,HI(g)$
ini, atm	0.200		0.200		0.42
Δ, atm	$+x$		$+x$		$-2x$
eq, atm	$0.200 + x$		$0.200 + x$		$0.42 - 2x$

$$\begin{aligned} \frac{(0.42 - 2\,x)^2}{(0.2 + x)^2} &= 0.25 \\ \frac{0.42 - 2\,x}{0.2 + x} &= \sqrt{0.25} \\ 0.42 - 2\,x &= (0.5)(0.2) + 0.5\,x \\ 2.5\,x &= 0.32 \\ x &= 0.128 \end{aligned}$$

$P_{HI} = 0.42$ atm $-$ (2) (0.128 atm) $= 0.164$ atm

008 10.0 points

Consider the reaction

$$Ni(CO)_4(g) \rightarrow Ni(s) + 4\,CO(g)\,.$$

If the initial concentration of $Ni(CO)_4(g)$ is 1.0 M, and x is the equilibrium concentration of CO(g), what is the correct equilibrium relation?

1. $K_c = \dfrac{x}{1.0 - \dfrac{x}{4}}$

2. $K_c = \dfrac{x^4}{1.0 - \dfrac{x}{4}}$ **correct**

3. $K_c = \dfrac{x^5}{1.0 - \dfrac{x}{4}}$

4. $K_c = \dfrac{4\,x}{1.0 - 4x}$

5. $K_c = \dfrac{x^4}{1.0 - 4x}$

Explanation:

009 10.0 points

At 990°C, $K_c = 1.6$ for the reaction

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

How many moles of $H_2O(g)$ are present in an equilibrium mixture resulting from the addition of 1.00 mole of H_2, 2.00 moles of CO_2, 0.75 moles of H_2O, and 1.00 mole of CO to a 5.00 liter container at 990°C?

1. 1.1 mol **correct**

2. 1.4 mol

3. 1.7 mol

4. 0.80 mol

5. 1.0 mol

6. 0.60 mol

Explanation:
$K_c = 1.6$

$$[H_2] = \frac{1.00\text{ mol}}{5\text{ L}} \qquad [CO_2] = \frac{2.00\text{ mol}}{5\text{ L}}$$

$$[H_2O] = \frac{0.75\text{ mol}}{5\text{ L}} \qquad [CO] = \frac{1.00\text{ mol}}{5\text{ L}}$$

$$Q = \frac{[H_2O][CO]}{[H_2][CO_2]} = \frac{\left(\frac{0.75\text{ mol}}{5\text{ L}}\right)\left(\frac{1\text{ mol}}{5\text{ L}}\right)}{\left(\frac{1\text{ mol}}{5\text{ L}}\right)\left(\frac{2\text{ mol}}{5\text{ L}}\right)} = 0.375 < K_c = 1.6$$

Therefore equilibrium moves to the right.

$$\begin{array}{cccc} H_2(g) + & CO_2(g) \rightleftharpoons & H_2O(g) + & CO(g) \\ 0.2 & 0.4 & 0.15 & 0.2 \\ -x & -x & x & x \\ \hline 0.2-x & 0.4-x & 0.15+x & 0.2+x \end{array}$$

$$\frac{(0.15+x)(0.2+x)}{(0.2-x)(0.4-x)} = 1.6$$

$$\frac{x^2 + 0.35x + 0.03}{x^2 - 0.6x + 0.08} = 1.6$$

$$x^2 + 0.35x + 0.03 = 1.6x^2 - 0.96x + 0.128$$

$$0.6x^2 - 1.31x + 0.098 = 0$$

$$x = \frac{1.31 \pm \sqrt{(1.31)^2 - 4(0.6)(0.098)}}{1.2} = 7.756 \times 10^{-2}\text{ M}$$

$$\text{mol } H_2O = 5.00\text{ L} \times \left(0.15 + 7.75 \times 10^{-2}\right)\frac{\text{mol}}{\text{L}} = 1.1\text{ mol } H_2O$$

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010 10.0 points

What happens to the concentration of NO(g) when the total pressure on the equilibrium reaction

$$3\,NO_2(g) + H_2O(\ell) \rightleftharpoons 2\,HNO_3(aq) + NO(g)$$

is increased (by compression)?

1. remains the same

2. Unable to determine

3. increases **correct**

4. decreases

Explanation:
Increasing the total pressure on the system by decreasing its volume will shift the equilibrium toward the side of the reaction with fewer numbers of moles of gaseous components. If the total number of moles of gas is the same on the product and reactant sides of the balanced chemical equation, then changing the pressure will have little or no effect on the equilibrium distribution of species present. The amount and concentration of NO will increase.

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011 10.0 points

Consider the reaction

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

where $\Delta H_{\text{rxn}} = -198$ kJ. The amount of $SO_2(g)$ at equilibrium increases when

1. the pressure is increased.

2. the volume is increased. **correct**

3. SO_3 is removed.

4. the temperature is decreased.

5. more oxygen is added.

Explanation:

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_2(g) + \text{heat}$$

According to Le Chatelier's principle, the amount of reactant SO_2(g) is increased when the equilibrium shifts to the left. This will happen when another reactant (O_2) is removed. For an exothermic reaction decreasing temperature removes heat and sends equilibrium to the right. Increasing pressure sends the equilibrium in the direction that has fewer numbers of moles of gas. Increasing the volume is equivalent to decreasing gas pressure.

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012 10.0 points

For an exothermic reaction, what would happen to the numerical value of K_c, if we increase the temperature at constant pressure?

1. K_c would increase.

2. K_c would decrease. **correct**

3. K_c would not change.

4. K_c would either increase or decrease, depending on the number of moles of gas involved.

5. K_c would either increase or decrease, depending on the concentrations.

Explanation:

An exothermic reaction gives off heat as products are formed, so at a higher temperature the reverse endothermic reaction is favoured, decreasing K_c, which reflects the ratio of the products to reactants.

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013 10.0 points

Suppose the reaction mixture

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

is at equilibrium at a given temperature and pressure. The pressure is then increased at constant temperature by compressing the reaction mixture, and the mixture is then allowed to reestablish equilibrium. At the new equilibrium,

1. there is more ammonia present than there was originally. **correct**

2. there is less ammonia present than there was originally.

3. there is the same amount of ammonia present as there was originally.

4. the nitrogen is used up completely.

Explanation:

LeChatelier's Principle states that if a change occurs in a system at equilibruim, the system responds to relieve the stress and reach a new equilibrium. Here, the number of moles of gaseous reactants is greater than the number of moles of products. Increasing the pressure of the above system will result in the reaction proceeding to reduce that pressure increase. The system will shift to the right (the side that has fewer moles of gas), so the pressure will be reduced; thus more ammonia will be produced.

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014 10.0 points

Consider the system

$$2\,N_2O_5(g) \rightleftharpoons 2\,N_2O_4(g) + O_2(g) + \text{heat}$$

at equilibrium at 25°C. If the temperature were raised would the equilibrium be shifted to produce more N_2O_5 or more N_2O_4?

1. more N_2O_5 **correct**

2. There would be no effect.

3. more N_2O_4

Explanation:

This is an exothermic reaction and so increasing temperature provides more heat to the system, shifting equilibrium to the left and producing more N_2O_5.

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015 10.0 points

Given the reaction

$$2\,NH_3(g) \rightleftharpoons N_2(g) + 3\,H_2(g)$$

at equilibrium, if the pressure is doubled (think of the volume of the container halving), in which direction will the reaction shift?

1. right

2. left **correct**

3. no change

Explanation:

Increasing pressure shifts the equilibrium in the direction that produces fewer moles of gas.

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016 10.0 points

The system

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

is at equilibrium at some temperature. At equilibrium a 4.00 L vessel contains 1.00 mole CO_2, 1.00 mole H_2, 2.40 moles H_2O, and 2.40 moles CO. How many moles of CO_2 must be added to the system to bring the equilibrium CO concentration to 0.677 mol/L?

Correct answer: 1.148 moles.

Explanation:

$V = 4.0$ L $\qquad n_{CO_2} = 1.0$ mol

$n_{H_2} = 1.0$ mol $\qquad n_{H_2O} = 2.40$ mol

$n_{CO} = 2.40$ mol

At equilibrium, the concentrations are

$$[CO_2] = [H_2] = \frac{1\text{ mol}}{4\text{ L}} = 0.25\text{ M}$$

$$[H_2O] = [CO] = \frac{2.4\text{ mol}}{4\text{ L}} = 0.6\text{ M}$$

First calculate the value of K:

$$K = \frac{[CO]\,[H_2O]}{[CO_2]\,[H_2]} = \frac{(0.6\text{ M})\,(0.6\text{ M})}{(0.25\text{ M})\,(0.25\text{ M})} = 5.76$$

Set up your equilibrium table knowing that the initial CO_2 concentration is now higher than 0.25 M. The reaction will go to the right and you can get the following...

CO_2	+	H_2	$\rightleftharpoons$	H_2O	+	CO
?		0.25		0.6		0.6
$-x$		$-x$		$+x$		$+x$
$? - x$		$0.25 - x$		$0.6 + x$		$0.6 + x$

The final concentration of CO is 0.677, so

$$0.6 + x = 0.677$$
$$x = 0.677 - 0.6 = 0.077$$

and $[H_2] = 0.25 - x = 0.173$ M.

We also now know all of the final concentrations except CO_2, which is calculated from the equilibrium expression

$$5.76 = \frac{(0.677\text{ M})\,(0.677\text{ M})}{[CO_2]\,(0.173\text{ M})}$$

$$[CO_2] = \frac{(0.677\text{ M})\,(0.677\text{ M})}{(5.76\text{ M})\,(0.173\text{ M})}$$

so that $[CO_2]_{\text{fin}} = 0.4599$ mol/L, which corresponds to $0.4599 + 0.077 = 0.5369$ for the initial concentration of CO_2 BEFORE the equilibrium occurred.

That value corresponds to 2.148 moles of CO_2. Now substract the 1.00 mole of CO_2 that was already there and you get 1.148 moles of CO_2 that had to be added.

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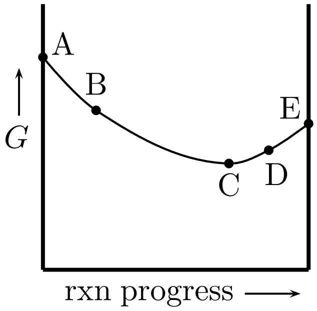
017 10.0 points

The figure represents a reaction at 298 K.



Based on the figure, which of the following statements (if any) is false?

1. None of the statements is false. **correct**

2. At point B, $Q < K$.

3. For this reaction ΔG° is negative.

4. At point C, the system is at equilibrium.

5. At point D, the reaction will move toward the reactants to get to equilibrium.

Explanation:
ΔG° is negative (point E is lower free energy than point A), $K > 1$. At point B, the reaction will move to the products ($Q < K$). At point E, the reaction will move to the reactants ($Q > K$). At point C, the reaction is at equilibrium ($Q = K$).

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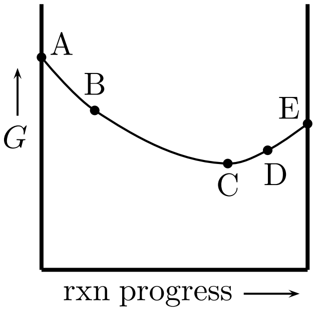
018 10.0 points

The following figure represents the progress of a given reaction at 298 K.



At point B on this figure, what is the relationship of Q to K?

1. $Q = K$

2. $Q > K$

3. $Q < K$ **correct**

4. Cannot be determined

Explanation:
Point B is on the reactants-heavy side of equilibrium, so Q is less than K. Note also that dG (slope) is negative here which means the reaction would be spontaneous in the forward direction. Spontaneous in a forward direction corresponds to Q being less than K.

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019 10.0 points

Given the hypothetical reaction

$$X(g) \rightleftharpoons Y(g)$$

predict what will happen when 1.0 mol Y is placed into an evacuated container.

1. Q will decrease until $Q = K$. **correct**

2. Q will increase until $Q = K$.

3. ΔG° will decrease until $\Delta G^\circ = 0$.

4. Nothing; the products are already formed, so no reaction occurs.

Explanation:
The container begins with products only. The reverse reaction will predominate until equilibrium is reached. The reaction quotient Q starts out larger than K and decreases until $Q = K$.

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020 10.0 points

Consider the reaction:
$C_{graphite}(s) + O_2(g) \leftrightarrow CO_2(g)$
$\Delta G^\circ = -400\ \text{kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$
Which of the following is a possible value of K for this reaction?

1. -0.56

2. 0.56

3. 10^{-70}

4. 10^{70} **correct**

Explanation:
$\Delta G^\circ = -400\ \text{kJ} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ is very large and negative, requiring that K be a very large number.

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021 10.0 points

The hydronium ion is _____.

1. OH^-

2. H_2O^+

3. HO^+

4. H_3O^+ **correct**

Explanation:
The hydronium ion H_3O^+ represents hydrated protons in aqueous solution.

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022 10.0 points

A strong acid (or base) is one which

1. dissolves metals.

2. dissolves glass.

3. reacts with a salt to form water.

4. should only be used when wearing goggles and gloves.

5. dissociates completely in aqueous solution. **correct**

Explanation:
By definition, a strong acid is completely dissociated in solution.

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023 10.0 points

HCN is classified as a weak acid in water. This means that it produces

1. a relatively small fraction of the maximum number of possible hydronium ions. **correct**

2. no hydronium ions.

3. a relatively large fraction of the maximum number of possible hydronium ions.

4. 100% of the maximum number of possible hydronium ions.

Explanation:
In water only a small percentage of a weak acid hydrolyzes, producing H_3O^+ ions.

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024 10.0 points

An acid is strong if it

1. causes metals to corrode.

2. makes acid-base indicators change color.

3. is very concentrated.

4. ionizes completely in water. **correct**

Explanation:
In water a weak acid is partially hydrolyzed and a strong acid is completely hydrolyzed. In both cases H_3O^+ ions are produced.

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025 10.0 points

Which of the following substances is NOT a weak acid?

1. HSO_3

2. HCN

3. HCl **correct**

4. H_3PO_4

5. HF

6. H_2CO_3

Explanation:
HCl is not a weak acid.

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026 10.0 points

Which of the following substances is a weak acid?

1. HNO_3

2. H_2SO_4

3. HI

4. HCl

5. $HClO_3$

6. H_2CO_3 **correct**

7. HBr

Explanation:
H_2CO_3 is a weak acid.

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027 10.0 points

In the reversible reaction

$$H_3O^+ + NH_3 \rightleftharpoons H_2O + NH_4^+,$$

what two substances act as acids in the Bronsted-Lowry sense?

1. some other pair

2. NH_3 and H_2O

3. H_3O^+ and NH_4^+ **correct**

4. H_3O^+ and H_2O

5. NH_3 and NH_4^+

Explanation:

In this reaction, H_3O^+ is donating a proton to OH^-. Therefore H_3O^+ is acting as an acid and OH^- is acting as a base. In the reverse reaction, NH_4^+ is the proton donor and thus, it is the other acid.

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028 10.0 points

In the two reactions represented by

$$HCN + H_2O \rightleftharpoons CN^- + H_3O^+,$$

the two Bronsted-Lowry acids are

1. HCN and H_3O^+. **correct**

2. H_2O and CN^-.

3. HCN and CN^-.

4. H_2O and H_3O^+.

5. There is only one Bronsted-Lowry acid shown.

Explanation:

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029 10.0 points

A water solution of sodium acetate is basic because

1. the acetate ion acts as a Bronsted-Lowrey base in a reaction with water. **correct**

2. sodium acetate is only weakly ionized.

3. the acetate ion acts as a Bronsted-Lowrey acid in a reaction with water.

4. the conjugate base of the acetate ion is a strong base.

5. the statement is false; a water solution of sodium acetate is acidic.

Explanation:

In water, sodium acetate dissociates into sodium ions (spectator ions) and acetate ions which hydrolyze to produce hydroxide ions.

$$CH_3COO^- + H_2O \leftrightarrow CH_3COOH + OH^-$$

The acetate ion accepts a proton from a water molecule.

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030 10.0 points

What is the conjugate acid of NO_3^-?

1. HNO_3 **correct**

2. H^+

3. ${NO_3}^{2-}$

4. ${NO_2}^-$

5. NH_3

6. OH^-

Explanation:

Since the question asks for the conjugate acid, we can assume NO_3^- is acting as a base. This means that it is a proton acceptor. To form the conjugate acid, it accepts a H making HNO_3.

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031 10.0 points

According to the Bronsted-Lowry concept of acids and bases, which of the following statements about a base is NOT true?

1. A base reacts with an acid to form a salt.

2. A base must contain a hydroxide group. **correct**

3. A base will share one of its electron pairs to bind H^+.

4. If a base is strong, then its conjugate acid will be relatively weaker.

Explanation:

According to the Bronsted-Lowry definition of bases, a base must be a proton acceptor. It does not need to contain a hydroxide group.

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032 10.0 points

Which statement is true for the following reaction?

$$CCl_3COOH + H_2O \leftrightarrow CCl_3CO_2^- + H_3O^+$$

1. H_3O^+ is the conjugate acid of $CCl_3CO_2^-$.

2. CCl_3COOH is the conjugate acid of $CCl_3CO_2^-$. **correct**

3. Cl_3COOH is the conjugate base of $CCl_3CO_2^-$.

4. H_2O is the conjugate base of $CCl_3CO_2^-$.

5. H_3O^+ is the conjugate base of $CCl_3CO_2^-$.

6. H_2O is the conjugate acid of $CCl_3CO_2^-$.

Explanation:

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033 10.0 points

According to Bronsted-Lowry Theory an acid is

1. amphoteric.

2. a proton acceptor.

3. a proton donor. **correct**

4. a soluble ionic hydroxide.

Explanation:

The Bronsted-Lowry theory of acids and bases defines an acid as a substance which is able to donate a proton.

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034 10.0 points

Which is NOT a conjugate base-acid pair?

1. $H_2O : H_3O^+$

2. $HSO_4^- : SO_4^{2-}$ **correct**

3. $OH^- : H_2O$

4. $CN^- : HCN$

5. $F^- : HF$

Explanation:

In the Bronsted-Lowry sense, a base is a proton acceptor and an acid is a proton donor. HSO_4^- is a proton donor and an acid. Therefore this is a conjugate acid-base pair, not a conjugate base-acid pair.

---

035 10.0 points

The conjugate base of H_2SO_4 is:

1. H_3O^+

2. SO_4^{2-}

3. OH^-

4. H_2SO_3

5. HSO_4^- **correct**

6. H_2O

7. $H_3SO_4^+$

Explanation:

Remove one H^+ ion from a species to find its conjugate base.

---

036 10.0 points

A given weak acid HZ has a $K_a = 3.6 \times 10^{-6}$. What is the H_3O^+ concentration of a solution of HZ that has a concentration of 0.76 mol/L?

Correct answer: 0.00165 mol/L.

Explanation:

$K_a = 3.6 \times 10^{-6}$ $[H_3O^+] = ?$

The expression for K_a here is

$$K_a = \frac{[H^+]\,[Z^-]}{[HZ]}$$

Substituting in values for the concentrations and K_a, we have

$$3.6 \times 10^{-6} = \frac{(x)\,(x)}{(0.76 - x)}$$

As HZ has a small K_a, we can assume that x is very small compared to 0.76, so we can simplify this equation to

$$\begin{aligned} 3.6 \times 10^{-6} &= \frac{(x)^2}{(0.76)} \\ 2.736 \times 10^{-6} &= x^2 \\ x &= 0.00165 \end{aligned}$$

Remembering that we called $[H^+]$ x in the equation, $[H^+]$ is 0.00165 mol/L.

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037 10.0 points

Assume that five weak acids, identified only by numbers (1, 2, 3, 4, and 5), have the following ionization constants.

Acid	Ionization Constant K_a value
1	1.0×10^{-3}
2	3.0×10^{-5}
3	2.6×10^{-7}
4	4.0×10^{-9}
5	7.3×10^{-11}

The anion of which acid is the strongest base?

1. 3

2. 5 **correct**

3. 2

4. 1

5. 4

Explanation: