

This print-out should have 50 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

A reaction has a negative change in entropy. This reaction can only be spontaneous if...

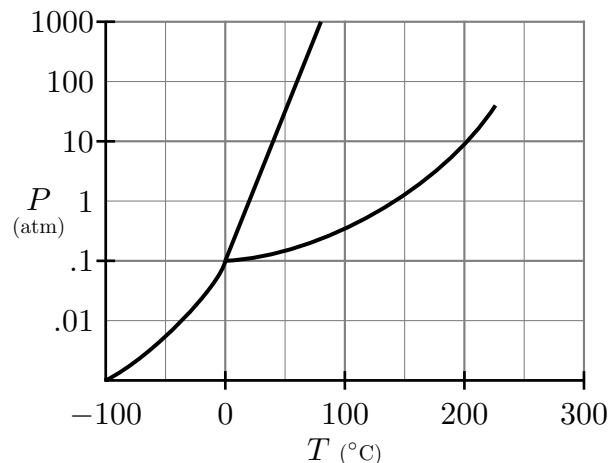
1. heat is absorbed by the system at any temperature
2. heat is released at any temperature
3. heat is absorbed at a sufficiently high temperature
4. None of these choices are correct because a reaction with a negative change in entropy can never be spontaneous
5. heat is released at a sufficiently low temperature **correct**

Explanation:

A reaction must have at least one favorable term between entropy and enthalpy in order to be spontaneous. The negative entropy term is unfavorable, meaning the reaction must be exothermic (heat is released). Moreover, the reaction will only have a negative free energy change if the temperature is small enough. Conceptually, this is because the energy must be able to flow from the system to the low temperature surroundings.

002 (part 1 of 2) 10.0 points

Use the following phase diagram for the next two questions.



What is the normal melting point for this substance? Note: the vertical axis is logarithmic in scale.

1. 150°C
2. 45°C
3. 120°C
4. 20°C **correct**
5. 75°C
6. 0°C

Explanation:

The normal melting point is the point on the line where the solid and liquid phases coincide at 1 atm. This is found just less than halfway between the 0 and 50°C gridlines.

003 (part 2 of 2) 10.0 points

A sample of this substance is held at 0.1 atm and -50°C . The sample is pressurized to 3 atm and then heated to 250°C . In total, what phase transitions occurred?

1. melting and boiling **correct**
2. sublimation only
3. sublimation and condensation
4. melting and condensation

5. melting and freezing

Explanation:

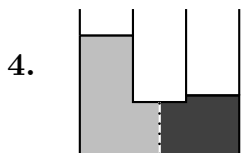
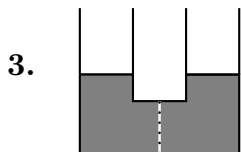
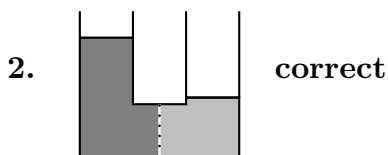
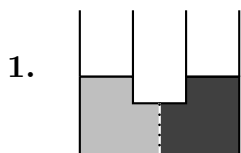
Pressurizing from 0.1 atm to 3 atm does not result in a change of phase. However, heating to 250°C results in the melting of the solid followed by the vaporization of the liquid.

004 10.0 points

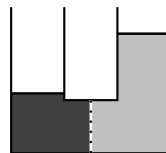
The following diagram shows a solution on the left (dark shade) and just the solvent on the right (light shade) separated by a semipermeable membrane.



Which diagram best represents the final state of this system after equilibrium is achieved?



5.

**Explanation:**

Only the solvent goes through the membrane thus increasing the volume on the solution side and decreasing volume on the solvent side. The solution side is therefore diluted slightly and the shade is lightened somewhat. The solute itself (the darker color) cannot pass through the membrane so the right side **MUST** also stay the same color of pure solvent (the lightest shade shown).

005 10.0 points

Consider the following substances: acetic acid (CH_3COOH), propane (C_3H_8), and acetone (CH_3COCH_3). The boiling points (in no particular order) are -42°C , 56°C , and 118°C . The vapor pressures (in no particular order) are 225 Torr, 15 Torr, and 6400 Torr. What is the boiling point and vapor pressure for acetic acid?

1. 118°C , 6400 Torr2. -42°C , 6400 Torr3. 56°C , 225 Torr4. -42°C , 15 Torr5. 118°C , 15 Torr **correct****Explanation:**

Of the three compounds listed, acetic acid has the strongest IMFs (hydrogen bond). This will correspond to the lowest vapor pressure and highest boiling point: 15 Torr and 118°C .

006 10.0 points

The enthalpy of vaporization of a liquid is measured to be about 28.4 kJ/mol and its normal boiling point is 128°C . At what temperature is the partial pressure of this substance 1180 torr?

1. -7.92°C
2. 181°C
3. -381°C
4. 176°C
5. 150°C **correct**
6. -281°C
7. 162°C

Explanation:

Here we use the Clausius-Clapeyron equation to solve for T_1 :

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

$$P_1 = 1180 \text{ torr}$$

$$P_2 = 760 \text{ torr}$$

$$T_2 = 124^{\circ}\text{C} + 273 = 401 \text{ K}$$

$$\Delta H_{\text{vap}} = 28.4 \text{ kJ/mol}$$

$$R = 0.008314 \text{ kJ/mol K}$$

And...

$$T_1 = 150^{\circ}\text{C}$$

007 10.0 points

A sample of 44.1 g of *para*-dichlorobenzene ($\text{C}_6\text{H}_4\text{Cl}_2$, 147.0 g/mol) is dissolved into 350 mL of hexane (C_6H_{14} , 86.18 g/mol, density 0.661 g/mL). What is the molality of this solution?

1. 1.30 *m* **correct**
2. 0.567 *m*
3. 0.101 *m*
4. 1.17 *m*
5. 0.857 *m*

Explanation:

Molality is defined as the moles of solute divided by kg solvent. Solve first for the mass

of solvent:

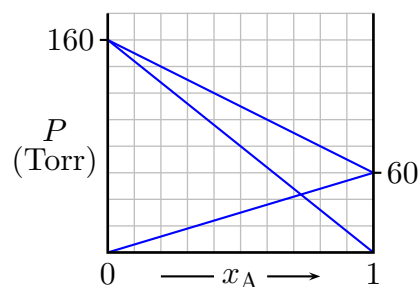
$$350 \text{ mL} \times \frac{0.661 \text{ g}}{\text{mL}} = 231.35 \text{ g}$$

mol of solute = $44.1/147 = 0.300$ mol solute
Next take the ratio of moles solute per kg solvent:

$$\frac{0.300 \text{ mol}}{0.23135 \text{ kg}} = 1.2967 = 1.30 \text{ m}$$

008 10.0 points

A and B are mildly volatile solvents. A mixture is made by combining 2 moles of A with 3 moles of B. Interpret the diagram below to determine the vapor pressure of this mixture.

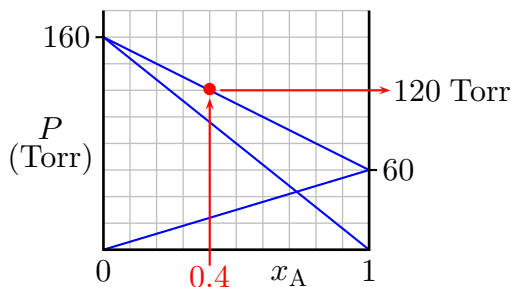


1. 80 Torr
2. 70 Torr
3. 100 Torr
4. 120 Torr **correct**
5. 140 Torr
6. 130 Torr
7. 110 Torr
8. 150 Torr
9. 90 Torr

Explanation:

Mixing 2 mol A and 3 mol B is a 2/5 mol fraction of A, or 0.4. Each vertical grid line is 0.1 mol fraction units, so counting 4 over we have the right line. Following it all the way up to the mixture vapor pressure we get 120 Torr (see diagram below). Alternatively, you can

use Raoult's Law and Dalton's law to take the sum of partial pressures, which is simply 120 torr = $0.4(60) + 0.6(160)$



009 10.0 points

Calculate the number of moles of oxygen that will dissolve in 45 L of water at 20°C if the partial pressure of oxygen is 0.21 atm. The Henry's Law constant for oxygen in water at 20°C is 0.0013 M/atm.

1. 0.0013 mol
2. 0.00027 mol
3. 0.28 mol
4. 0.012 mol **correct**
5. 0.0062 mol

Explanation:

010 10.0 points

A 19.7 g sample of an unknown salt (formula = MX_2) is dissolved in 249.4 mL water. The boiling point of water in this solution is 100.657°C. What is the molecular weight of the unknown salt?

1. 129.8 g/mol
2. 55.4 g/mol
3. 46.1 g/mol
4. 185 g/mol **correct**
5. 61.6 g/mol

Explanation:

Using the equation $\Delta T = i \cdot k_b \cdot m$, you can

solve for the molality of the unknown salt in water.

$$m = \frac{0.657}{3 \cdot 0.512}$$

$$m = 0.4277 \text{ mol kg}^{-1}$$

Then convert molality into moles by multiplying by the mass of the H_2O solvent:

$$\frac{0.4277 \text{ mol}}{\text{kg}} \left(\frac{0.2494 \text{ kg}}{1} \right) = 0.1067 \text{ mol}$$

Finally, we can use the gram quantity and number of moles to solve for the molecular weight of the salt:

$$\frac{19.7 \text{ g}}{0.1067 \text{ mol}} = 185 \text{ g/mol}$$

011 10.0 points

Isocarboxazid (MW = 231.25 g/mol) is an organic monoamine oxidase inhibitor used to treat depression disorders. 38.00 grams of isocarboxazid are added to water to make a 350 mL aqueous solution. What is the osmotic pressure exerted by this solution across a semi-permeable membrane at 37°C?

1. 12.10 atm
2. 329.6 atm
3. 11.95 atm **correct**
4. 1.43 atm
5. 144.4 atm
6. 23.90 atm

Explanation:

Osmotic pressure is calculated using the formula:

$$\Pi = MRT$$

$$R = 0.08206 \text{ L} \cdot \text{atm/mol} \cdot \text{K}$$

$$T = 310.15 \text{ K}$$

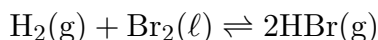
Solving first for molarity:

$$M = \frac{(38.0 \text{ g}) \left(\frac{\text{mol}}{231.25 \text{ g}} \right)}{0.350 \text{ L}} = 0.4695 \text{ M}$$

$$11.95 \text{ atm} = (0.4695)(0.08206)(310.15)$$

012 10.0 points

Write the equilibrium constant for the following reaction.



1. $K_p = \frac{P_{\text{HBr}}}{P_{\text{H}_2}}$

2. $K_p = \frac{P_{\text{HBr}}^2}{P_{\text{H}_2} P_{\text{Br}_2}}$

3. $K_p = \frac{P_{\text{H}_2}}{P_{\text{HBr}}^2}$

4. $K_p = \frac{P_{\text{HBr}}^2}{P_{\text{H}_2} [\text{Br}_2]}$

5. $K_p = \frac{P_{\text{HBr}}^2}{P_{\text{H}_2}}$

correct

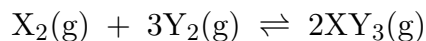
Explanation:

Take the concentration of the products raised to the power of their coefficients divided by the reactants raised to the power of their coefficients, ignoring any liquid or solids.

$$K_p = \frac{P_{\text{HBr}}^2}{P_{\text{H}_2}}$$

013 10.0 points

Consider the following generic gas phase reaction.



The value of K_p for this reaction is 107. Calculate the equilibrium partial pressure of gas Y_2 if the equilibrium partial pressures of XY_3 is 0.50 atm and X_2 is 0.15 atm?

1. 5.6 atm

2. 0.25 atm **correct**

3. 0.016 atm

4. 0.42 atm

5. 0.33 atm

6. 0.031 atm

7. 0.18 atm

Explanation:

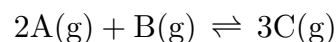
$$107 = K_p$$

$$107 = \frac{(P_{\text{XY}_3})^2}{(P_{\text{X}_2})(P_{\text{Y}_2})^3} = \frac{(0.50)^2}{(0.15)(P_{\text{Y}_2})^3}$$

$$P_{\text{Y}_2} = 0.25 \text{ atm}$$

014 10.0 points

0.834 atm A, 0.565 atm B, and 1.24 atm C are placed into a container to run the following reaction:



At equilibrium, 0.435 atm C remains. What is K_p for this reaction?

1. 0.0526 **correct**

2. 0.107

3. 0.0230

4. 8.12

5. 18.8

Explanation:

The first thing to observe in this question is that the reaction runs in reverse, because C is consumed over the course of the reaction. Running a RICE table, you will see that the change in C is equal to $-3x = 0.435 - 1.24 = -0.805$; therefore, $x = 0.2683$.

$$P_{\text{A}} = 0.834 + 2x = 1.37067$$

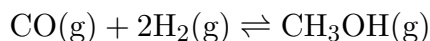
$$P_{\text{B}} = 0.565 + x = .83333$$

Now you can solve for the equilibrium constant, K_p :

$$K_p = \frac{0.435^3}{1.37067^2 \cdot .83333} = 0.05258$$

015 (part 1 of 2) 10.0 points

Consider the following reaction for the next two questions:



Beginning at equilibrium, which of the following will result in a shift toward the right of this reaction?

1. Adding CH_3OH gas
2. Adding an inert gas at constant pressure
3. Adding an inert gas at constant volume
4. Removing CO gas
5. Reducing the volume of the container **correct**

Explanation:

The only option that will cause a shift to the right of this reaction is reducing the volume of the container. By reducing the volume of the container, the pressure is increased. Increasing the overall pressure results in a shift way from the side of the reaction with the most gas moles.

016 (part 2 of 2) 10.0 points

If this reaction is exothermic, lowering the temperature will cause the reaction to...

1. shift right due to a decreased Q value
2. shift left due to a smaller K value
3. shift left due to a smaller Q value
4. shift right due to a larger K value **correct**
5. remain at equilibrium

Explanation:

Reducing the temperature of an exothermic reaction will cause the reaction to shift to the right. This is because an exothermic reaction will have a larger K value when the temperature is decreased.

017 10.0 points

Calculate the pH of a 0.018 M Ba(OH)_2 solution.

1. 8.44
2. 12.26
3. 12.56 **correct**
4. 5.26
5. 1.44
6. 1.74

Explanation:

0.018 M Ba(OH)_2 is a strong base that dissociates completely, resulting in a 0.036 M hydroxide ion concentration.

$$\text{pOH} = -\log(0.036 \text{ M})$$

$$\text{pH} = 14 - \text{pOH} = 12.56$$

018 10.0 points

What is the pH of a 0.12 M solution of hydroxylammonium bromide (NH_3OHBr)?

1. 9.56
2. 3.72
3. 3.24
4. 3.61
5. 3.33
6. 4.44
7. 6.04
8. 3.48 **correct**

Explanation:

$$M_{\text{NH}_3\text{OHBr}} = 0.12 \text{ M} \quad K_b = 1.1 \times 10^{-8}$$

It's a salt of a weak base (BHX). This means you need a K_a for the weak acid BH^+ . Use $K_a = \frac{K_w}{K_b}$ and you'll get the $K_a = 9.09091 \times$

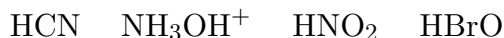
10^{-7} . You CAN use the approximation for the equilibrium which means that

$$[\text{H}^+] = \sqrt{K_a \cdot C_{\text{BH}^+}} = 0.000330289 \text{ M}$$

$$\text{pH} = -\log(0.000330289) = 3.48111$$

019 10.0 points

Rank the following acids in increasing order of acidity.



- $\text{NH}_3\text{OH}^+ < \text{HNO}_2 < \text{HBrO} < \text{HCN}$
- $\text{NH}_3\text{OH}^+ < \text{HBrO} < \text{HCN} < \text{HNO}_2$
- $\text{HCN} < \text{NH}_3\text{OH}^+ < \text{HNO}_2 < \text{HBrO}$
- $\text{HCN} < \text{HBrO} < \text{NH}_3\text{OH}^+ < \text{HNO}_2$ **correct**
- $\text{HNO}_2 < \text{NH}_3\text{OH}^+ < \text{HBrO} < \text{HCN}$
- $\text{HNO}_2 < \text{HBrO} < \text{HCN} < \text{NH}_3\text{OH}^+$

Explanation:

Look at the coversheet for the values of $\text{p}K_a$.

$$\text{HCN} = 9.21$$

$$\text{NH}_3\text{OH}^+ = 14 - 7.96 = 6.04$$

$$\text{HNO}_2 = 3.40$$

$$\text{HBrO} = 8.20$$

$$\text{HCN} < \text{HBrO} < \text{NH}_3\text{OH}^+ < \text{HNO}_2$$

020 10.0 points

A weak acid, HA, ionizes 7.65% at a 0.250 M concentration. What is the hydroxide ion concentration in this solution?

- $1.91 \times 10^{-16} \text{ M}$
- $1.31 \times 10^{-13} \text{ M}$
- $1.00 \times 10^{-14} \text{ M}$
- $5.23 \times 10^{-13} \text{ M}$ **correct**
- $1.91 \times 10^{-2} \text{ M}$

Explanation:

Percent ionization quickly gives you the dissociation of a weak electrolyte at a particular concentration:

$$[\text{H}^+] = 0.0765 \times 0.25 \text{ M}$$

$$[\text{H}^+] = 0.019125 \text{ M}$$

Use K_w to convert this into $[\text{OH}^-]$:

$$[\text{OH}^-] = \frac{1 \times 10^{-14}}{0.019125 \text{ M}}$$

$$[\text{OH}^-] = 5.23 \times 10^{-13} \text{ M}$$

021 10.0 points

What is the pH after 250 mL of 0.25 M HNO_3 is added to 350 mL of 0.50 M CH_3NH_2 (methylamine) ?

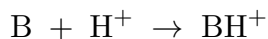
- 10.90 **correct**
- 6.21
- 5.05
- 3.36
- 10.64
- 10.38
- 3.62
- 2.29

Explanation:

This is a simple Henderson-Hasselbach calculation.

$$\text{pH} = \text{p}K_a + \log\left(\frac{C_B}{C_{\text{BH}^+}}\right)$$

The $\text{p}K_b$ for methylamine (the base, B) is 3.36 which means the $\text{p}K_a$ for the methylammonium ion (BH^+) is 10.64. You may use the ratio of moles because this is a partial neutralization (buffer) problem. The weak base methylamine, B, is converted to the conjugate acid, BH^+ via the reaction with the strong acid.



The acid is the limiting reactant and determines the amount of BH^+ made. Remember, volume times concentration equals amount (moles).

$$\text{mol } BH^+ = 0.25(0.250) = 0.0625 \text{ mol}$$

$$\text{mol } B = .50(0.350) - 0.0625 \text{ mol} = .1125 \text{ mol}$$

Now solve for pH using the H-H equation:

$$\text{pH} = 10.64 + \log \left(\frac{0.1125}{0.0625} \right)$$

$$\text{pH} = 10.64 + 0.26$$

$$\text{pH} = 10.90$$

022 10.0 points

What is the ratio of potassium acetate to acetic acid necessary to make a buffer with a pH equal to 5.12?

- 3.1
- 2.4 **correct**
- 0.38
- 9.86
- 0.42
- 2.9

Explanation:

You can solve for the ratio using the H-H equation and knowledge of log rules:

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

$$5.12 = 4.74 + \log \left(\frac{[A^-]}{[HA]} \right)$$

$$10^{5.12-4.74} = \frac{[A^-]}{[HA]}$$

$$= 2.4$$

023 10.0 points

What is the dominant species in solution at the equivalence point of a weak base-strong

acid titration?

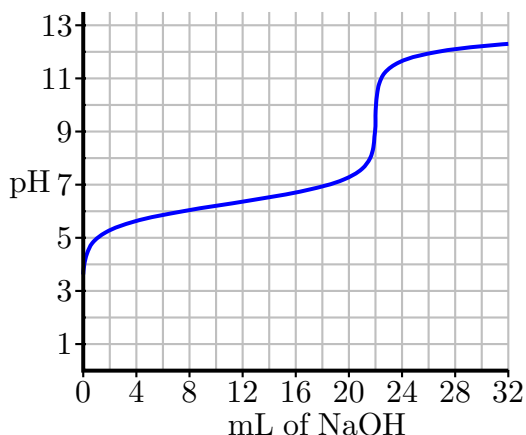
- Strong base
- Equal parts weak acid and weak base
- Weak acid **correct**
- Neutral salt
- Weak base
- Strong acid

Explanation:

The dominant species in solution at the equivalence point will be the conjugate of the analyte. In this case, the dominant species in solution will be the conjugate of the weak base, which is just a weak acid.

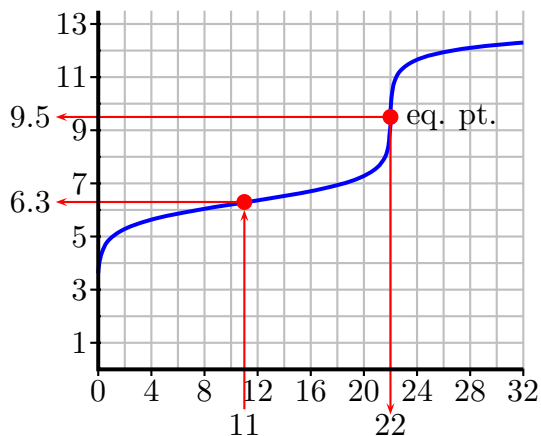
024 (part 1 of 3) 10.0 points

A sample of 30 mL of a weak acid (HA) solution was titrated with 0.075 M NaOH. The pH curve for this titration is shown.



What is the concentration of the original weak acid solution (the 30 mL) ?

- 0.032 M
- 0.022 M
- 0.055 M **correct**
- 0.075 M
- 0.048 M

Explanation:

equivalence point is 22 mL of NaOH
 $\times 0.075 \text{ M} = 1.65 \text{ mmol OH}^-$
 $\text{mmol acid} = \text{mmol base}$
 $1.65 \text{ mmol} / 30 \text{ mL} = 0.055 \text{ M HA}$

025 (part 2 of 3) 10.0 points

Which of the following is the value of K_a for the weak acid, HA ?

1. 1.3×10^{-6}
2. 5.0×10^{-7} **correct**
3. 3.2×10^{-10}
4. 1.2×10^{-7}
5. 7.6×10^{-5}

Explanation:

The half titration point is at 11 mL and the $\text{pH} = \text{pK}_a$ there which is 6.3. That would match a K_a of 5.0×10^{-7} .

026 (part 3 of 3) 10.0 points

Below is a listing of five indicators and their associated pK_a values. Which indicator would be the best one to use for this titration?

1. methyl red, 5.0
2. bromophenol blue, 4.1
3. bromocresol purple, 6.4

4. phenol red, 7.4

5. thymol blue, 9.3 **correct**

6. alizarin yellow, 10.9

Explanation:

The equivalence point is at a pH of 9.5. Only thymol blue in this listing covers that pH during the color transition.

027 10.0 points

Which of the following salts is the most soluble in pure water?

1. BaSO_4 $K_{\text{sp}} = 1.1 \times 10^{-10}$
2. CuBr $K_{\text{sp}} = 6.3 \times 10^{-9}$
3. All of these salts have the same solubility
4. CaF_2 $K_{\text{sp}} = 3.5 \times 10^{-11}$ **correct**

Explanation:

Calculate the molar solubility, x , for each substance:

$$\text{CaF}_2: K_{\text{sp}} = 3.5 \times 10^{-11} = 4x^3$$

$$x = 2.06 \times 10^{-4} \text{ (correct)}$$

$$\text{BaSO}_4: K_{\text{sp}} = 1.1 \times 10^{-10} = x^2$$

$$x = 1.05 \times 10^{-5}$$

$$\text{CuBr}: K_{\text{sp}} = 6.3 \times 10^{-9} = x^2$$

$$x = 7.94 \times 10^{-5}$$

028 10.0 points

Barium fluoride (BaF_2) is most soluble in which of the following solutions?

1. 0.15 M $\text{Ba}(\text{OH})_2$
2. 0.50 M NaF
3. 0.18 M NaF
4. 0.005 M $\text{Ba}(\text{OH})_2$ **correct**
5. The molar solubility of barium fluoride is the same in each of these solutions

Explanation:

According to the common ion effect, solubility decreases as the effective concentration of a common ion increases. The smallest common ion concentration would be the 0.005 M Ba(OH)₂.

029 10.0 points

You mix 0.02 mmol Sr(NO₃)₂ solution and 0.05 mmol NaF solution to form a 100 mL solution. What precipitate (if any) forms?

1. NaF
2. SrF₂
3. No precipitate forms **correct**
4. Sr(NO₃)₂
5. NaNO₃

Explanation:

First, come up with the double-displacement equation for this problem to show that SrF₂ is the solid product. However, this product will only be formed if $Q > K$. Therefore, you must solve for Q_{sp} at the instant of mixing:

$$Q_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2$$

$$Q_{sp} = (.0002)(0.0005)^2 = 5.11 \times 10^{-11}$$

Compare to K_{sp} , which is 4.3×10^{-9} . Because Q is less than K , the solution is undersaturated and no precipitate forms.

030 10.0 points

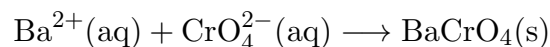
What is the mass of the barium chromate precipitate resulting from the addition of 300 mL 0.025 M Ba(OH)₂ to 200 mL 0.040 M Na₂CrO₄?

1. 3.48 g
2. 2.03 g
3. 1.90 g **correct**
4. 3.80 g

5. 2.52 g

Explanation:

Start with the relevant net ionic formula based on what was provided in the question stem (Note: you can confirm that a precipitate will form using Q_{sp}):



Here you should see that this is a limiting reagent problem. Because barium chromate has a very small K_{sp} , it is safe to assume the reaction goes approximately 100%. Solve for moles of each reagent:

$$\text{Ba}^{2+} : 0.0250 \text{ M} \times 0.300 \text{ L} = 0.00750 \text{ mol}$$

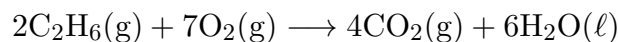
$$\text{CrO}_4^{2-} : 0.0400 \text{ M} \times 0.200 \text{ L} = 0.00800 \text{ mol}$$

This reaction is 1:1:1, so you can determine that the limiting reagent is the reactant with the fewest number of moles. This is also the number of moles of product formed. Convert to mass:

$$0.00750 \times 253.37 \text{ g/mol} = 1.90 \text{ g}$$

031 10.0 points

Consider the following reaction:



Oxygen is being consumed at a rate equal to 1.24 M/s. What is the initial rate at which carbon dioxide is forming?

1. 4.96 M/s
2. 8.68 M/s
3. 2.17 M/s
4. 1.24 M/s
5. 0.709 M/s **correct**

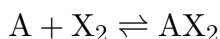
Explanation:

The stoichiometric ration between carbon dioxide and oxygen is 4:7. This means that carbon dioxide is being formed at $\frac{4}{7}$ the initial rate of oxygen.

$$1.24 \text{ M/s} \times \frac{4 \text{ CO}_2}{7 \text{ O}_2} = 0.709 \text{ M/s}$$

032 (part 1 of 2) 10.0 points

You run an experiment to determine the initial rates of the following generic reaction at various starting conditions:



	[A] M	[X ₂] M	initial rate M · s ⁻¹
Trial 1	0.60	1.56	2.00 × 10 ⁻³
Trial 2	0.60	3.12	8.00 × 10 ⁻³
Trial 3	1.20	1.56	4.00 × 10 ⁻³
Trial 4	0.90	2.40	7.10 × 10 ⁻³

What is the correct rate law for the reaction?

1. Rate = (1.40)[A]
2. Rate = (3.84 × 10⁻³)[A]⁻¹[X₂]²
3. Rate = (3.84 × 10⁻³)[A][X₂]²
4. Rate = (1.37 × 10⁻³)[A][X₂]
5. Rate = (1.37 × 10⁻³)[A][X₂]² **correct**

Explanation:

Begin with the equation:

$$\text{Rate} = k[\text{A}]^x[\text{B}]^y$$

Compare trials to determine the value of x and y . Two good comparisons are trials 1 and 2 and trials 1 and 3.

Between trials 1 and 2, the concentration of X₂ doubles and the rate quadruples. Therefore, $y = 2$, meaning the reaction is second order in X₂.

Between trials 1 and 3, [A] doubles and the rate doubles. This means $x = 1$ and the reaction is first order in A.

Lastly, you can solve for k using any trial:

$$2.00 \times 10^{-3} = (1.37 \times 10^{-3})(0.60)(1.56)^2$$

033 (part 2 of 2) 10.0 points

What are the units of the rate constant in the previous question?

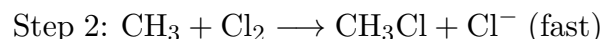
1. $\frac{1}{\text{M}^2 \cdot \text{s}}$ **correct**
2. $\frac{1}{\text{M}^3 \cdot \text{s}}$
3. $\frac{1}{\text{M}^4 \cdot \text{s}}$
4. $\frac{1}{\text{M} \cdot \text{s}}$
5. $\frac{\text{M}}{\text{s}}$

Explanation:

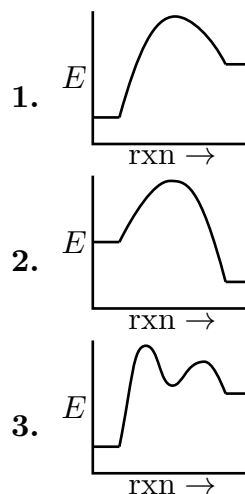
The rate constant has units that cancel out the concentrations in the rate law to give a rate in M/s. The rate constant has units of $\frac{1}{\text{M}^2 \cdot \text{s}}$

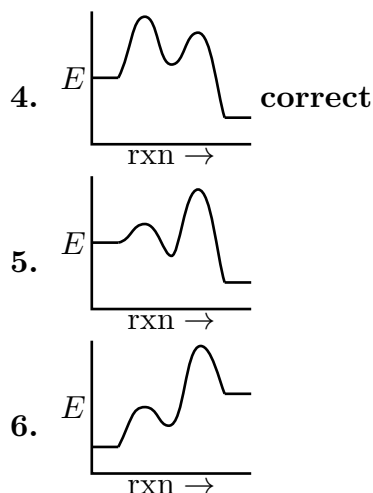
034 10.0 points

The chlorination of methane is an exothermic reaction with a two-step mechanism shown below:



Which of the following reaction coordinate diagrams best fits this data?

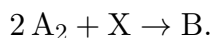


**Explanation:**

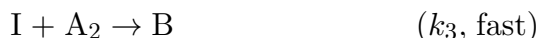
The reaction coordinate diagram should match the mechanism and thermodynamics of the reaction. The reaction profile is exothermic, has two humps, and the first hump should be the largest (rate-determining step).

035 10.0 points

Consider the following overall reaction:



Using the overall reaction, determine the rate law for the following mechanism:



1. Rate = $k' [\text{Z}][\text{Y}]$

2. Rate = $k' [\text{A}_2] [\text{X}]$ **correct**

3. Rate = $k' [\text{A}_2]^2 [\text{X}]$

4. Rate = $k' [\text{Z}] [\text{X}]$

5. Rate = $k' [\text{A}_2]^2$

6. Rate = $k' [\text{A}_2] [\text{Z}] [\text{X}]$

Explanation:

The slowest step is the rate determining step and is used to write the rate law:

$$\text{Rate} = k_2 [\text{Z}] [\text{Y}]$$

For the final answer, we need the rate law in terms of only the reactants of the overall reaction. We need to substitute in for $[\text{Z}]$ using the previous fast step in equilibrium:

$$k_1 [\text{A}_2] [\text{X}] = k_{-1} [\text{Z}][\text{Y}]$$

$$[\text{Z}] = \frac{k_1 [\text{A}_2] [\text{X}]}{k_{-1} [\text{Y}]}$$

If you fit this into your original rate law, you will see that $[\text{Y}]$ cancels out:

$$\text{Rate} = \frac{k_2 k_1 [\text{A}_2] [\text{X}] [\text{Y}]}{k_{-1} [\text{Y}]}$$

So our final answer is just:

$$\text{Rate} = k' [\text{A}_2] [\text{X}]$$

036 10.0 points

How does a catalyst affect the rate of a chemical reaction?

1. A catalyst increases the energy of the transition state such that a larger number of particles have sufficient energy to overcome the activation energy

2. A catalyst increases the rate constant by providing an alternate mechanism with a lower activation energy **correct**

3. A catalyst increases the rate constant by increasing the activation energy

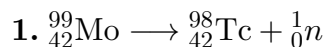
4. A catalyst decreases the rate constant by lowering the activation energy

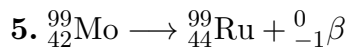
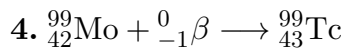
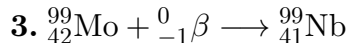
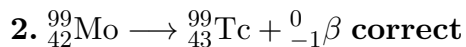
Explanation:

A catalyst increases the rate constant by providing an alternate mechanism with a lower activation energy

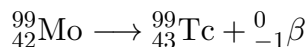
037 10.0 points

${}_{42}^{99}\text{Mo}$ undergoes radioactive decay by emitting a single beta particle. Which of the following reactions corresponds to this process?



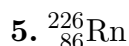
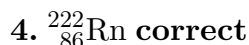
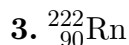
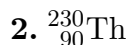
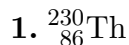
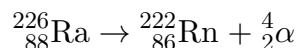
**Explanation:**

Write the balanced reaction involving the emission of a ${}_{-1}^0\beta$ particle.



038 10.0 points

Identify the missing isotope in the nuclear reaction.

**Explanation:**

039 10.0 points

${}^{123}\text{I}$ is a radioactive isotope ($t_{1/2} = 13.22$ hours) useful for clinical imaging. How long will it take for a dose to diminish to 18.7% of its original value?

1. 22.2 hours

2. 38.7 hours

3. 557 hours

4. 29.7 hours

5. 87.9 hours

6. 34.2 hours

7. 70.7 hours

8. 32.0 hours **correct**

Explanation:

This is radioactive decay, meaning we must follow first-order kinetics:

$$\ln\left(\frac{100}{18.7}\right) = kt$$

Solve for k first:

$$k = \frac{\ln(2)}{13.22 \text{ hours}}$$

$$k = 0.052431708 \text{ hours}^{-1}$$

Now complete the equation up top to solve for t :

$$\ln\left(\frac{100}{18.7}\right) / 0.052431708 \text{ hours}^{-1} = t$$

$$t = 32.0 \text{ hours}$$

040 10.0 points

When direct heat is applied to potassium chlorate, KClO_3 , it decomposes to form KCl and other byproducts. Was chlorine oxidized or reduced? How many electrons were transferred during the process?

1. oxidized, 6 electrons

2. oxidized, 4 electrons

3. reduced, 9 electrons

4. reduced, 3 electrons

5. reduced, 6 electrons **correct**

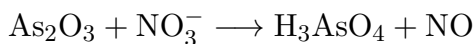
6. oxidized, 3 electrons

Explanation:

For the conversion of KClO_3 to KCl , the Cl atom is going from a +5 to a -1 oxidation state. This is a reduction reaction that involves the gain of 6 electrons.

041 10.0 points

The following reaction occurs in acidic conditions. What is the coefficient of water in the overall balanced equation? Is it a reactant or a product?



1. 7; reactant **correct**

2. 3; product

3. 2; product

4. 2; reactant

5. 4; reactant

6. 3; reactant

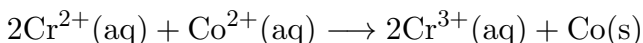
Explanation:

The correctly balanced equation is:

$$7\text{H}_2\text{O} + 4\text{H}^+ + 3\text{As}_2\text{O}_3 + 4\text{NO}_3^- \longrightarrow 6\text{H}_3\text{AsO}_4 + 4\text{NO}$$

042 (part 1 of 2) 10.0 points

What is the shorthand notation for the following electrochemical cell?



1. Pt | Cr²⁺, Cr³⁺ || Co²⁺ | Co **correct**

2. Cr²⁺, Cr³⁺ || Co²⁺ | Co

3. Co | Co²⁺ || Cr²⁺, Cr³⁺ | Pt

4. Co²⁺ | Co || Cr²⁺, Cr³⁺

5. Cr²⁺ | Cr³⁺ || Co²⁺ | Co

Explanation:

You must use an inert electrode in the chromium solution compartment because you do not have a solid metal anode. Otherwise, follow your convention of: anode—*anodic solution*—cathode—*cathodic solution*—cathode

043 (part 2 of 2) 10.0 points

What is the oxidizing agent in the previous problem?

1. Co

2. Pt

3. Cr²⁺

4. Co²⁺ **correct**

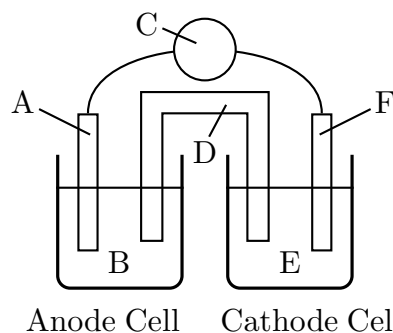
5. Cr³⁺

Explanation:

The oxidizing agent is the species being reduced. This is Co²⁺.

044 (part 1 of 2) 10.0 points

The following two questions refer to this diagram for a voltaic cell. Neither of the two electrodes are an inert electrode.



Where would you find the species that is being oxidized?

1. E

2. C

3. B

4. D

5. A **correct**

Explanation:

A is the electrode in the anode cell. Oxidation always occurs at the anode.

045 (part 2 of 2) 10.0 points

If the half-reaction for the anode involves Fe²⁺ and Fe, which of these redox pairs could be in the cell on the right?

1. Mn^{2+} and Mn
2. H^+ and H_2
3. None of these can give a voltaic cell
4. Sn^{2+} and Sn **correct**
5. Cr^{3+} and Cr

Explanation:

Because the Fe^{2+}/Fe half-reaction is the anode and this is a galvanic cell, the other half-reaction must have a standard reduction potential that is more positive (higher up on the standard reduction potential table) than that of Fe^{2+}/Fe . Only two answer choices have a more positive reduction potential: Sn^{2+}/Sn , and H^+/H_2 . But the H^+/H_2 half-reaction has no solid component to serve as an electrode and therefore requires a platinum electrode, and the problem specifies that there are no platinum electrodes. That leaves Sn^{2+}/Sn as the only option.

046 10.0 points

Using an electroplating system operating at 7.35 amps, it take 1.50 hours to plate out 5.00 grams of an unknown metal from its molten chloride salt, MCl_2 . Identify the metal M.

1. Cu
2. Fe
3. Cd
4. Zn
5. Mg **correct**

Explanation:

$$\frac{I \cdot t}{n \cdot F} = \text{moles of metal}$$

$$\frac{7.35 (1.5 \times 60 \times 60)}{2 \cdot 96485} = .206 \text{ mol metal}$$

$$\text{molar mass} = 5 \text{ g} / .206 \text{ mol} = 24.3 \text{ g}$$

This matches the molar mass of Mg.

047 10.0 points

Consider a standard voltaic cell at equilibrium. Which of the following is true?

1. $E < 0, \Delta G > 0, K < 0$
2. $E > 0, \Delta G > 0, K > 1$
3. $E < 0, \Delta G > 0, K < 1$
4. $E = 0, \Delta G = 0, K = 1$
5. $E = 0, \Delta G = 0, K > 1$ **correct**

Explanation:

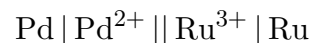
For a voltaic cell at equilibrium:

$$E = 0, \Delta G = 0, K > 1$$

It is important to realize here that E and ΔG are not standard values, so at equilibrium they are equal to zero (think of this voltaic cell as a dead battery).

048 10.0 points

Consider the following cell:



What is ΔG° for the overall cell reaction that is represented here? Balance the reaction using the lowest possible integer values.

1. -91.2 kJ
2. $+182 \text{ kJ}$ **correct**
3. $+877 \text{ kJ}$
4. -182 kJ
5. $+91.2 \text{ kJ}$
6. -877 kJ

Explanation:

$$\text{Pd}^{2+}/\text{Pd} = +0.915 \text{ V}$$

$$\text{Ru}^{3+}/\text{Ru} = +0.60 \text{ V}$$

$$E^\circ = 0.60 - 0.915 = -0.315 \text{ V}$$

$$\Delta G^\circ = -nFE^\circ$$

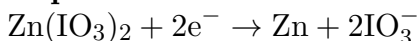
$$= -6(96485)(-0.315) = +182356 \text{ J}$$

$$= +182 \text{ kJ}$$

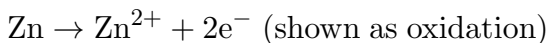
049 10.0 points

Use half-reactions from the standard reduction table to calculate the K_{sp} for $\text{Zn}(\text{IO}_3)_2$.

1. 1.7×10^{-26}
2. 6.8×10^{-32}
3. 4.8×10^{-12}
4. 3.9×10^{-6} **correct**
5. 7.3×10^{-19}

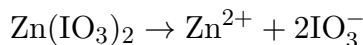
Explanation:

$$E^\circ = -0.922 \text{ V}$$



$$E^\circ = +0.762 \text{ V}$$

Adding the half-reactions together, we get the solubility reaction for $\text{Zn}(\text{IO}_3)_2$:



$$E^\circ = -0.922 + 0.762 = -0.160 \text{ V}$$

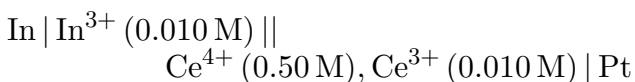
K for this reaction then will be K_{sp} for $\text{Zn}(\text{IO}_3)_2$.

$$K_{\text{sp}} = 10^{\frac{nE^\circ}{0.05916}}$$

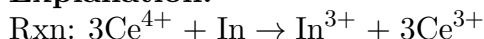
$$= 10^{\frac{(2)(-0.16)}{0.05916}} = 3.9 \times 10^{-6}$$

050 10.0 points

What is the potential for the following cell?



1. 2.37 V
2. 1.95 V
3. 1.81 V
4. 2.02 V
5. 2.09 V **correct**

6. 1.88 V**Explanation:**

$$E^\circ = 1.61 - (-0.34) = 1.95 \text{ V}$$

$$Q = \frac{[\text{Ce}^{3+}]^3[\text{In}^{3+}]}{[\text{Ce}^{4+}]^3} = \frac{(0.01)^3(0.01)}{(0.5)^3}$$

$$= 8.0 \times 10^{-8}$$

$$E = 1.95 - \frac{0.05916}{3} \log(8.0 \times 10^{-8}) = 2.09 \text{ V}$$