General Chemistry Questions

Electronic Structure and Periodic Table

1. What value or values of m_l are allowable for an orbital with $l = 2$?

a. 0
b. 2
c. -1
d. none of the above
e. all of the above

2. According to Bohr Theory, which of the following transitions in the hydrogen atom will give rise to the **least energetic** photon?
Use the equation: $E_n = (-2.18 \times 10^{-18}\ \text{J})(1/n^2)$

a. $n = 5$ to $n = 3$
b. $n = 6$ to $n = 1$
c. $n = 4$ to $n = 3$
d. $n = 5$ to $n = 4$
e. $n = 6$ to $n = 5$

3. Consider a $3d_{xz}$ orbital. Which of the following statements is **incorrect**?

a. The *xz* plane is a nodal surface.
b. The *xz* plane divides the electron probability distribution into two identical mirror-image halves.
c. The *xy* plane divides the electron probability distribution into two identical mirror-image halves.
d. The *yz* plane divides the electron probability distribution into two identical mirror-image halves.
e. The nucleus is located at a node.

4. The electronic configuration of the element whose atomic number is 26 is:

a. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^8$
b. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$
c. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$
d. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4\ 4p^2$
e. none of the above

5. Which of the following has the largest radius?

a. F
b. N
c. C
d. O
e. Ne

6. Which of the following elements has the largest ionization energy?

a. Na
b. Ne
c. F
d. K
e. Rb

7. Which of the following has the greatest electron affinity (most negative value)?

a. Cl
b. K
c. He
d. Na
e. Rb

8. Which of the following species is not isolectronic with any of the others?

a. V^{3+}
b. Ca^{2+}
c. Ar
d. Cl^{-}
e. S^{2-}

9. In Bohr's model of the hydrogen atom, the radius of an orbit

a. is proportional to n^2.
b. is smallest for the highest energy state.
c. increases when a photon of light is emitted from an excited atom.
d. can have any value that is larger than the ground-state radius.
e. none of the above

10. Which of the following atoms is **not** a one-electron system?

a. H
b. He^{+}
c. Li^{2+}
d. Be^{2+}
e. O^{7+}

11. Which of the following statements about periodic properties is incorrect?

a. Both electron affinity and ionization energy decrease down a group.
b. Atomic size increases to the right across a period.
c. Ionization energy increases to the right across a period.
d. Atomic size increases down a group.
e. Electron affinity increases to the right across a period.

Bonding

1. Which one of the following is most likely to be an ionic compound?

a. HNF_2
b. H_2CO
c. N_2H_4
d. $CaCl_2$
e. CH_3Cl

2. In which of the following processes does the enthalpy change (ΔH) directly represent the magnitude of the lattice energy of KCl(*s*)?

a. $Cl_2(g) + 2K(s) \rightarrow 2KCl(s)$
b. $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$
c. $KCl(s) \rightarrow K^+(g) + Cl^-(g)$
d. $KCl(s) \rightarrow K(s) + Cl^-(g)$
e. $KCl(s) \rightarrow K(s) + Cl(g)$

3. Order the following by increasing bond strength: N≡N, N=N, N-N

a. N≡N, N=N, N-N
b. N≡N, N-N, N=N
c. N-N, N=N, N≡N
d. N=N, N-N, N≡N
e. N=N, N≡N, N-N

4. Which of the following compounds has the greatest bond polarity?

a. PH_3
b. NH_3
c. HF
d. H_2S
e. CH_4

5. Which of the following is not planar?

a. BCl_3
b. ClF_3
c. PCl_3
d. XeF_4
e. C_2H_4

6. Use VSEPR theory to predict the ideal bond angles around the two carbon atoms in acetaldehyde, CH_3CHO. (The first carbon has single bonds to three H atoms and one C atom; the second carbon has single bonds to C and H, and a double bond to O.)

a. 109°, 109°
b. 109°, 120°
c. 120°, 109°
d. 120°, 90°
e. 105°, 105°

7. In a carbon-carbon triple bond, what is the nature of the bonding between the carbons?

a. two 2*s* orbitals overlapping
b. two 2*p* orbitals overlapping
c. two *sp* orbitals overlapping, two $2p_y$ overlapping and two $2p_z$ overlapping
d. an *sp* and sp^2 overlapping and 2*p* orbitals overlapping
e. an sp^2 and sp^2 overlapping and 2*p* orbitals overlapping

8. Which of the following molecules has sp^3 hybridization and a dipole moment?

a. SiH_4
b. BF_3
c. NH_3
d. BrF_3
e. PCl_5

9. In the molecular orbital description of bonding in benzene (C_6H_6), how many electrons occupy delocalized πMOs?

a. 2
b. 3
c. 4
d. 5
e. 6

10. In which of the following species is the octet rule violated by the central atom?

a. CH_4
b. SF_4
c. PCl_4^+
d. SO_2
e. NH_3

11. The number of electron dots in the Lewis symbol for an element equals the

a. number of outermost *p* electrons.
b. number of electrons needed to fill the outermost *p* orbital.
c. period number that contains the element.
d. number of outermost *s* and *p* electrons.
e. number of outermost *s* electrons.

Phases and Phase Equilibria

1. Calculate the pressure of 0.55 mol of NH_3 gas in a 2.00 L vessel at 25 °C, using the ideal gas law.

a. 2.5 atm
b. 6.7 atm
c. 0.6 atm
d. 7.5 atm
e. 3.4 atm

2. A steel tank contains carbon dioxide at 34 °C and is at a pressure of 13.0 atm. Determine the internal gas pressure when the tank and its contents are heated to 100 °C.

a. 10.7 atm
b. 9.4 atm
c. 38.2 atm
d. 1.9 atm
e. 15.8 atm

3. Deviations from the ideal gas law are less at:

a. high temperatures and high pressures
b. high temperatures and low pressures
c. low temperatures and high pressures
d. low temperatures and low pressures
e. high volumes and low temperatures

4. A mixture of three gases has a pressure of 1380 mmHg at at 298 K. The mixture is analyzed and is found to contain 1.27 mol CO_2, 3.04 mol CO, and 1.50 mol Ar. What is the partial pressure of Ar?

a. 238 mm Hg
b. 302 mm Hg
c. 356 mm Hg
d. 1753 mm Hg
e. 8018 mm Hg

5. Which of the following exhibits the most hydrogen bonding?

a. LiH
b. CH_4
c. NH_3
d. H_2S
e. CH_2F_2

6. Which of the following carbon compounds has the highest melting point?

a. CF_4
b. CCl_4
c. CBr_4
d. CI_4
e. CH_4

7. Water has such a high specific heat because

a. it has such a low molecular weight.
b. it is rather dense.
c. the O-H single bond has a high bond energy.
d. it has many relatively strong hydrogen bonds.
e. it dissolves both ionic and covalent compounds.

8. The triple point is

a. an end to the liquid-gas line in a phase diagram.
b. the relationship between the boiling point, melting point and vapor pressure of a substance.
c. the point on a phase diagram where solid, liquid, and gas are in equilibrium.
d. the three pieces of data needed to solve the Clausius-Clapeyron equation.
e. the (P,V,T) coordinate of a point on a phase diagram.

9. The main forces responsible for the structure of DNA are

a. ionic bonds and covalent bonds.
b. covalent bonds and ionic bonds.
c. hydrogen bonds and dipole-dipole interactions.
d. covalent bonds and hydrogen bonds.
e. covalent bonds and dipole-dipole interactions.

10. Which of the following is not likely to exhibit hydrogen bonding?

a. CH_3CH_2OH
b. CH_3NH_2
c. H_2O
d. NH_2OH
e. $(CH_3)_3N$

Stoichiometry

1. What is the mass of one mole of acetylsalicylic acid (aspirin), $C_9H_8O_4$?

a. 29 g
b. 108 g
c. 196 g
d. 180. g
e. none of the above

2. Determine the number of moles of aluminum in 2.154×10^{-1} kg of Al.

a. 5816 mol
b. 7.984 mol
c. 6.02×10^{23} mol
d. 4.801 mol
e. 8.783 mol

3. How many grams of zinc are there in 22.7 g of $ZnCl_2$?

a. 0.35 g
b. 0.17 g
c. 10.9 g
d. 1476 g
e. 0.32 g

4. A compound with a composition of 87.5 % N and 12.5 % H was recently discovered. What is the empirical formula for this compound?

a. NH_2
b. N_2H_3
c. NH
d. N_2H_2
e. N_2H

5. This equation is unbalanced: $PCl_3 + H_2O \rightarrow H_3PO_3 + HCl$ When it is correctly balanced, the coefficients are, respectively

a. 1,3,1,1
b. 1,1,1,3
c. 1,3,1,3
d. 2,3,2,3
e. none of the above

6. Given 6 mol of each reactant, which one would be limiting in the following reaction?
$4Au + 8NaCN + O_2 + 2H_2O \rightarrow 4NaAu(CN)_2 + 4NaOH$

a. Au
b. NaCN
c. O_2
d. H_2O
e. There is no limiting reactant.

7. In the direct reaction of silicon with Cl_2 the yield of $SiCl_4$ is 50. %. How many grams of silicon must be reacted with excess chlorine in order to obtain 17 g $SiCl_4$?

a. 1.4 g
b. 2.8 g
c. 5.6 g
d. 17 g
e. 28 g

8. In the reaction of Fe_3O_4 with carbon to form carbon dioxide and iron, the number of moles of carbon required to convert 23 g of Fe_3O_4 to products is

a. 0.05
b. 0.1
c. 0.2
d. 0.3
e. 0.4

9. A 20.0 mL sample of an element with a density of 3.0 g/mL contains 4×10^{23} atoms. What is the atomic weight of this element?

a. 300
b. 40
c. 60
d. 90
e. none of the above

10. How many moles of oxygen gas will react with 12.4 mol aluminum?
Equation: $4Al + 3O_2 \rightarrow 2Al_2O_3$

a. 0.24 mol
b. 0.42 mol
c. 4.8 mol
d. 9.3 mol
e. 16.8 mol

11. Balance the following redox equation occurring in aqueous solution:
$KMnO_4 + KCl + H_2SO_4 \rightarrow MnSO_4 + K_2SO_4 + H_2O + Cl_2$
What is the stoichiometric coefficient for chlorine (Cl_2) when the equation is balanced with smallest whole number coefficients?

a. 1
b. 3
c. 5
d. 8
e. 10

Thermodynamics and Thermochemistry

1. Data:
(1) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$ $\Delta H = -241.8$ kJ
(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H = -285.8$ kJ

On the basis of the above data, which of the following statements is **false**?

a. Reaction (1) is exothermic.
b. Reaction (2) is the formation reaction for $H_2O(l)$.
c. The reverse of reaction (2) is endothermic.
d. The energy content of $H_2O(g)$ is lower than $H_2O(l)$.
e. ΔH for the reaction: $H_2O(l) \rightarrow H_2O(g)$ is + 44 kJ/mol.

2. What is the amount of heat necessary to raise the temperature of 8.5 kg of water from 12.5 °C to 84 °C?

a. 3.0×10^3 kJ
b. 36 J
c. 2.5×10^3 kJ
d. 2.5×10^6 kJ
e. 25 kJ

3. Data:
$\Delta H°_f$ values: $CH_4(g)$, -74.8 kJ; $CO_2(g)$, -393.5 kJ; $H_2O(l)$, -285.8 kJ.
Using the $\Delta H°_f$ data above, calculate $\Delta H°_{rxn}$ for the reaction below.
Reaction: $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

a. -604.2 kJ
b. 890.3 kJ
c. -997.7 kJ
d. -890.3 kJ
e. none of the above

4. Data:
$2Ba(s) + O_2(g) \rightarrow 2BaO(s) \quad \Delta H° = -1107.0$ kJ
$Ba(s) + CO_2(g) + \frac{1}{2}O_2(g) \rightarrow BaCO_3(s) \quad \Delta H° = -822.5$ kJ
Given the data above, calculate $\Delta H°$ for the reaction below.
Reaction: $BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$

a. -1929.5 kJ
b. -1376.0 kJ
c. -284.5 kJ
d. 269.0 kJ
e. 537 kJ

5. Which of the following is **not** a state function?

a. ΔE
b. ΔH
c. q
d. P
e. V

6. Two solutions (the system), each of 25.0 mL volume and at 25.0 °C, are mixed in a beaker. A reaction occurs between them, causing the temperature to drop to 20.0 °C. After the products have equilibrated with the surroundings, the temperature is again 25.0 °C and the total volume is 50.0 mL. No gases are involved in the reaction. Which one of the following relationships concerning the change from initial to final states (both at 25.0 °C) is correct?

a. $\Delta E = 0$
b. $\Delta H = 0$
c. $\Delta E < 0$
d. $w = 0$
e. $q = 0$

7. Which one of the following processes is exothermic?

a. $H_2(l) \rightarrow H_2(g)$
b. $CO_2(s) \rightarrow CO_2(g)$
c. $H_2O(g) \rightarrow H_2O(l)$
d. $16CO_2(g) + 18H_2O(l) \rightarrow 2C_8H_{18}(l) + 25O_2(g)$
e. $H_2(g) \rightarrow 2H(g)$

8. Predict the signs of $\Delta H°$, $\Delta S°$, and $\Delta G°$ for the vaporization of liquid water at 150°C.

a. $\Delta H° > 0, \Delta S° > 0, \Delta G° > 0$
b. $\Delta H° < 0, \Delta S° < 0, \Delta G° < 0$
c. $\Delta H° > 0, \Delta S° < 0, \Delta G° > 0$
d. $\Delta H° > 0, \Delta S° > 0, \Delta G° < 0$
e. none of the above

9. Which of the following substances has the lowest standard molar entropy ($S°$) at 25°C?

a. $CH_3OH(l)$
b. $CO(g)$
c. $MgO(s)$
d. $H_2O(l)$
e. $CaCO_3(s)$

10. When crystalline solid barium hydroxide octahydrate and crystalline solid ammonium nitrate are mixed in a beaker at room temperature, a spontaneous reaction occurs. The temperature of the beaker contents rapidly falls to below 0º C. Use this information to decide whether the reaction is exothermic or endothermic and what the signs of ΔH and ΔS are.

a. endothermic; $\Delta H > 0; \Delta S > 0$
b. exothermic; $\Delta H < 0; \Delta S > 0$
c. endothermic; $\Delta H < 0; \Delta S < 0$
d. endothermic; $\Delta H < 0; \Delta S > 0$
e. exothermic; $\Delta H > 0; \Delta S < 0$

11. Sodium carbonate can be made by heating sodium hydrogen carbonate:
$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$
For this reaction, $\Delta H° = 128.9$ kJ and $\Delta S° = 321$ J/K. At approximately what temperature will $K = 1$?

a. 401.6 K
b. 401.6º C
c. 33.1 K
d. 33.1º C
e. none of the above

Rate Processes in Chemical Reactions—Kinetics and Equilibrium

1. For the overall hypothetical reaction $A + 5B \rightarrow 4C$, the rate of appearance of C given by $\Delta[C]/\Delta t$ is the same as

a. $\Delta[A]/\Delta t$
b. $-(5/4)(\Delta[B]/\Delta t)$
c. $-(4/5)(\Delta[B]/\Delta t)$
d. $-(1/4)(\Delta[A]/\Delta t)$
e. none of the above.

2. The initial rate of the reaction
$PCl_5 \rightarrow PCl_3 + Cl_2$
is increased a factor of four when the concentration of PCl_5 is doubled. Therefore, the rate

a. depends on the concentrations of PCl_3 and Cl_2.
b. is first order with respect to PCl_5.
c. is second order with respect to PCl_5.
d. is fourth order with respect to PCl_5.
e. is first order with respect to PCl_3.

3. Consider the reaction A → products. Which of the following plots is consistent with a zero-order reaction?

a. [A] plotted against time gives a horizontal, straight line.
b. In [A] plotted against time gives a straight line of negative slope.
c. 1/[A] plotted against time gives a straight line of positive slope.
d. [A] plotted against time gives a straight line of negative slope.
e. [A] plotted against time gives a curved line of negative slope, decreasing in magnitude as time increases

4. The rate constant of a first-order reaction is $3.68 \times 10^{-2}\ s^{-1}$ at 150°C, and the activation energy is 71 kJ/mol. What is the value of the rate constant at 170°C?

a. $9.2 \times 10^{-2}\ s^{-1}$
b. $3.7 \times 10^{-2}\ s^{-1}$
c. $2.49\ s^{-1}$
d. $4.0 \times 10^{-2}\ s^{-1}$
e. none of the above

5. The reaction
$3ClO^-(aq) \rightarrow ClO_3^-(aq)$
$+ 2Cl^-(aq)$ has been proposed to occur by the following mechanism.
$ClO^-(aq) + ClO^-(aq) \rightarrow ClO_2^-(aq) + Cl^-(aq)$ (slow)
$ClO_2^-(aq) + ClO^-(aq) \rightarrow ClO_3^-(aq) + Cl^-(aq)$ (fast)
Which rate law is consistent with this mechanism?

a. rate = $k[ClO^-]$
b. rate = $k\,[ClO^-]^3$
c. rate = $k\,[ClO_2^-][ClO^-]$
d. rate = $k\,[ClO^-]^2$
e. rate = $k\,[Cl^-][ClO^-]^2$

6. A catalyst speeds up a reaction by

a. increasing the number of high-energy molecules.
b. increasing the temperature of the molecules in the reaction.
c. increasing the number of collisions between molecules.
d. increasing the activation energy for the reaction.
e. providing a new reaction pathway for molecules.

7. Consider the following gas-phase equilibrium:
$H_2(g) + I_2(g) \leftrightarrow 2HI(g)$
At a certain temperature, the equilibrium constant K_c is 4.0. Starting with equimolar quantities of H_2 and I_2 and no HI, when equilibrium was established, 0.20 moles of HI was present. How much H_2 was used to start the reaction?

a. 0.10 mol
b. 0.23 mol
c. 0.20 mol
d. 4.0 mol
e. Need to know the volume of the reaction vessel.

8. At a certain temperature the equilibrium constant $K_p = 0.132$ for the reaction:
$PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$
At equilibrium, the partial pressures of both PCl_5 and PCl_3 are 100. mmHg. What is the total pressure of the equilibrium system, in mmHg?

a. 100. mmHg
b. 200. mmHg
c. 300. mmHg
d. 400. mmHg
e. 332 mmHg

9. Ammonium iodide dissociates reversibly to ammonia and hydrogen iodide:
$NH_4I(s) \leftrightarrow NH_3(g) + HI(g)$
At 400ºC, $K_p = 0.215$. If 150 g of ammonium iodide is placed into a 3.00-L vessel and heated to 400º C, calculate the partial pressure of ammonia when equilibrium is reached.

a. 0.22 atm
b. 0.46 atm
c. 0.11 atm
d. 0.88 atm
e. 1.2 atm

10. Consider the equilibrium reaction:
$3ClO^-(aq) \leftrightarrow ClO_3^-(aq) + 2Cl^-(aq)$
The equilibrium constant $K_c = 3.2 \times 10^3$. The following concentrations are present: $[Cl^-] = 0.50$ mol/L; $[ClO_3^-] = 0.32$ mol/L; $[ClO^-] = 0.24$ mol/L. Is the mixture at equilibrium and, if not, in which direction will reaction proceed?

a. The system is at equilibrium.
b. The system is not at equilibrium; reaction will proceed left to right.
c. The system is not at equilibrium; reaction will proceed right to left.
d. The system cannot reach equilibrium since the ClO_3^- and Cl^- concentrations are not in the stoichiometric ratio.
e. There is not enough information to tell.

11. Consider the following reaction in the gas phase:
$H_2 + I_2 \leftrightarrow 2HI$
If the pressure increased by reducing the the volume of the flask,

a. more HI will be produced.
b. more H_2 and I_2 will be produced.
c. the results will depend on what the amounts of each are.
d. the amount of HI will remain the same.
e. the equilibrium constant will change.

Solution Chemistry

1. Which of the following ions has an **incorrect** charge?

a. N^{3-}
b. Al^{3+}
c. S^{2-}
d. Cl^{-}
e. Mg^{2-}

2. Which of the following pairs of elements would be most likely to form an ionic compound?

a. P and Br
b. Zn and K
c. C and O
d. Al and Rb
e. F and Ca

3. What is the name of NaI?

a. sodium iodide
b. sodium(I) iodide
c. sodium monoiodide
d. sodious iodide
e. sodium iodine

4. Which of the following combinations of names and formulas is incorrect?

a. H_3PO_4 phosphoric acid
b. HNO_3 nitric acid
c. $NaHCO_3$ sodium carbonate
d. H_2CO_3 carbonic acid
e. KOH potassium hydroxide

5. Calculate the concentration of calcium ions in a saturated calcium phosphate solution. (K_{sp} for $Ca_3(PO_4)_2 = 1.3 \times 10^{-26}$)

a. 1.2×10^{-5} mol/L
b. 2.0×10^{-5} mol/L
c. 6.6×10^{-6} mol/L
d. 7.8×10^{-6} mol/L
e. 8.3×10^{-6} mol/L

6. Calculate the molar solubility of silver carbonate in 1.0 *M* sodium carbonate solution. (K_{sp} for $Ag_2CO_3 = 8.1 \times 10^{-12}$)

a. 8.1×10^{-12} *M*
b. 2.8×10^{-6} *M*
c. 1.4×10^{-6} *M*
d. 1.4×10^{-8} *M*
e. 2.0×10^{-4} *M*

7. Calculate the pH of a solution necessary to just begin the precipitation of $Mg(OH)_2$ when $[Mg^{2+}] = 0.001$ *M*. (K_{sp} for $Mg(OH)_2 = 1.2 \times 10^{-11}$)

a. 11
b. 10
c. 9
d. 8
e. 4

8. In qualitative analysis, the metals of Ion Group 1 can be separated from other ions by precipitating them as chloride salts. A solution initially contains Ag^+ and Pb^{2+} at a concentration of 0.10 *M*. Aqueous HCl is added to this solution until the Cl^- concentration is 0.10 *M*. What will the concentrations of Ag^+ and Pb^{2+} be at equilibrium? (K_{sp} for AgCl = 1.8×10^{-10}; K_{sp} for $PbCl_2 = 1.7 \times 10^{-5}$)

a. $[Ag^+] = 1.8 \times 10^{-11}$ *M*; $[Pb^{2+}] = 1.7 \times 10^{-6}$ *M*
b. $[Ag^+] = 1.8 \times 10^{-7}$ *M*; $[Pb^{2+}] = 1.7 \times 10^{-4}$ *M*
c. $[Ag^+] = 1.8 \times 10^{-11}$ *M*; $[Pb^{2+}] = 8.5 \times 10^{-5}$ *M*
d. $[Ag^+] = 1.8 \times 10^{-9}$ *M*; $[Pb^{2+}] = 1.7 \times 10^{-3}$ *M*
e. $[Ag^+] = 1.8 \times 10^{-9}$ *M*; $[Pb^{2+}] = 8.5 \times 10^{-6}$ *M*

9. Silver chloride is relatively insoluble in water (K_{sp} for AgCl = 1.8×10^{-10}) but it is soluble in aqueous ammonia, due to the formation of the complex ion $Ag(NH_3)_2^+$. How many moles of AgCl will dissolve in 1.00 L of solution containing 6.0 moles of free NH_3? (K_f for $Ag(NH_3)_2^+ = 1.7 \times 10^7$)

a. 9.1×10^{-6} mol
b. 2.9×10^{-4} mol
c. 0.0091 mol
d. 0.084 mol
e. 0.33 mol

10. What is the mass of $C_{12}H_{22}O_{11}$ in 60.0 mL of 0.0880 *M* solution?

a. 0.181 g
b. 1.81 g
c. 5.02 g
d. 5.28 g
e. none of the above

11. The freezing point of pure camphor is 178.4 °C, and its molal freezing-point constant, K*f* is 40.0 °C/*m*. Find the freezing point of a solution containing 3.00 g of a compound of molar mass 125 g/mol in 45.0 g of camphor.

a. 174.1 °C
b. 157.1 °C
c. 135.2 °C
d. 140.4 °C
e. 11.6 °C

Acids and Bases

1. Calculate the **hydroxide** ion concentration of a solution if its pH is 6.389.

a. 1.00×10^{-14} mol/L
b. 4.08×10^{-7} mol/L
c. 9.92×10^{-7} mol/L
d. 2.45×10^{-8} mol/L
e. none of the above

2. Which of the following is a correct description of the natural direction of a Brønsted-Lowry acid-base reaction?

a. weaker acid + weaker base → stronger acid + stronger base
b. weaker acid + stronger base → stronger acid + weaker base
c. stronger acid + weaker base → weaker acid + stronger base
d. stronger acid + stronger base → weaker acid + weaker base
e. None of the above statements is always correct.

3. In a 0.100 *M* HF solution, the percent dissociation is determined to be 9.5%. Calculate the K_a for HF based on this data.

a. 9.5×10^{-2}
b. 1.0×10^{-3}
c. 3.1×10^{-3}
d. 7.6×10^{-4}
e. 9.5×10^{-4}

4. What is the pH of a solution prepared from 0.250 mol of NH_3 dissolved in sufficient water to make 1.00 L of solution? ($K_b = 1.8 \times 10^{-5}$)

a. 2.12
b. 2.67
c. 8.92
d. 11.33
e. 13.40

5. Which of the following reactions illustrate $Al(OH)_3$ acting as a Lewis acid?

a. $Al(OH)_3 \rightarrow Al^{3+} + 3OH^-$
b. $Al(OH)_3 + OH^- \rightarrow Al(OH)_2O^- + H_2O$
c. $Al(OH)_3 + OH^- \rightarrow Al(OH)_4^-$
d. $Al(OH)_3 + 3H^+ \rightarrow Al^{3+} + 3H_2O$
e. $Al^{3+} + 3OH^- \rightarrow Al(OH)_3$

6. Which of the following pairs of species is **not** a conjugate acid-base pair?

a. HCl and H^+
b. HSO_4^- and SO_4^{2-}
c. H_2SO_4 and HSO_4^-
d. H_2O and OH^-
e. NH_3 and NH_2^-

7. Consider each of the following pairs of acids. Which statement is correct?

a. $HClO_2$ is a stronger acid than $HClO_4$.
b. H_2SO_4 is a stronger acid than H_2SeO_4.
c. H_2O is a stronger acid than HF.
d. H_2S is a stronger acid than H_2Se.
e. HS^- is a stronger acid than H_2S.

8. Consider the reaction
$CH_3NH_2 + H_2O \rightarrow CH_3NH_3^+ + OH^-$
where CH_3NH_2 is methylamine and $CH_3NH_3^+$ is the methylammonium ion. Select the correct description of this reaction in terms of Lewis acid-base theory.

a. Methylamine serves as a Lewis acid in the forward reaction and methylammonium ion serves as a Lewis base in the reverse reaction.
b. Water serves as a Lewis base in the forward reaction and the hydroxide ion serves as a Lewis base in the reverse reaction.
c. Methylamine serves as a Lewis base in the forward reaction and hydroxide ion serves as a Lewis acid in the reverse reaction.
d. Water serves as a Lewis acid in the forward reaction and methylammonium ion serves as a Lewis base in the reverse reaction.
e. Methylamine serves as a Lewis base in the forward reaction and hydroxide ion serves as a Lewis base in the reverse reaction.

9. What is the pH of a buffer prepared by adding 180 mL of 0.100 *M* NaOH to 200 mL of 0.100 *M* acetic acid?
(K_a for $CH_3COOH = 1.8 \times 10^{-5}$)

a. 3.79
b. 4.34
c. 4.74
d. 5.04
e. 5.70

10. Consider the titration of 50.00 mL of 0.1000 *M* HBr with 0.1000 *M* KOH. Calculate the pH after 49.00 mL of the base has been added to the 50.00 mL of HBr.

a. 2.0
b. 3.0
c. 4.0
d. 6.0
e. 7.0

11. An aqueous solution of a weak acid, HA, is titrated with NaOH solution. The pH at the midpoint of the buffer region is 4.5. What is the K_a of the acid?

a. 3.2×10^{-5}
b. 3.2×10^{-10}
c. 1.8×10^{-3}
d. 7.0×10^{-7}
e. 4.5

Electrochemistry

1. Which of the following statements is incorrect?

a. In an electrolytic cell, reduction occurs at the anode.
b. Aluminum metal would form at the cathode during the electrolysis of molten $AlBr_3$.
c. The cathode is labeled "+" in a voltaic cell.
d. Oxidation occurs at the anode in a voltaic cell.
e. Electrons flow from the anode to the cathode in all electrochemical cells.

2. Consider the following notation for an electrochemical cell
$Zn|Zn^{2+}$ (1*M*)||Fe^{2+} (1*M*), Fe^{3+} (1*M*)|Pt
What is the balanced equation for the cell reaction?

a. $Zn(s) + 2Fe^{3+}(aq) \rightarrow 2Fe^{2+}(aq) + Zn^{2+}(aq)$
b. $Zn^{2+}(aq) + 2Fe^{2+}(aq) \rightarrow Zn(s) + 2Fe^{3+}(aq)$
c. $Zn(s) + 2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Zn^{2+}(aq)$
d. $Zn(s) + Fe^{3+}(aq) \rightarrow Fe^{2+}(aq) + Zn^{2+}(aq)$
e. $Zn(s) + Fe^{2+}(aq) \rightarrow Fe(s) + Zn^{2+}(aq)$

3. Consider the following half-reactions and select the strongest oxidizing agent present:
$Sr^{2+}(aq) + 2e^- \rightarrow Sr(s)$ $E° = -2.89$ V
$Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$ $E° = -0.913$ V
$Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ $E° = -0.28$ V

a. $Cr^{2+}(aq)$
b. $Sr^{2+}(aq)$
c. $Co^{2+}(aq)$
d. $Sr(s)$
e. $Co(s)$

4. In an electrolytic cell, how many grams of Cu could be plated out of a $CuSO_4$ solution at a current of 5.00 A for 2.00 min? (F = 96500 C/mol)

a. 318 g
b. 0.395 g
c. 0.329×10^{-3} g
d. 0.198 g
e. 5.31 g

5. A voltaic cell is based on the following two half-reactions:
$Ni^{+2}(aq) + 2e^- \rightarrow Ni(s)$ $E° = -0.25$ V
$Cr^{+3}(aq) + 3e^- \rightarrow Cr(s)$ $E° = -0.74$ V
Sketch the cell and then select the correct statement about it.

a. Cr serves as the cathode.
b. The direction of electron flow through the external wire is from the Ni to the Cr electrode.
c. Anions in solution will migrate **toward** the Ni^{+2}/Ni electrode.
d. The net cell reaction is $3Ni^{+2}(aq) + 2Cr(s) \rightarrow 3Ni(s) + 2Cr^{+3}(aq)$
e. $E°_{cell} = 0.99$ V

6. Consider the following two electrode reactions and their standard electrode potentials:
$Al^{+3}(aq) + 3e^- \rightarrow Al\ (s)$ $E° = -1.66$ V
$Cd^{+2}(aq) + 2e^- \rightarrow Cd(s)$ $E° = -0.40$ V
Write the cell reaction for a voltaic cell based on these two electrodes, and calculate the standard cell potential, $E°_{cell}$.

a. $2Al^{+3}(aq) + 3Cd^{+2}(aq) \rightarrow 2Al(s) + 3Cd(s)$ $E°_{cell} = 2.10$ V
b. $2Al(s) + 3Cd^{+2}(aq) \rightarrow 2Al^{+3}(aq) + 3Cd(s)$ $E°_{cell} = 1.26$ V
c. $2Al(s) + 3Cd^{+2}(aq) \rightarrow 2Al^{+3}(aq) + 3Cd(s)$ $E°_{cell} = 3.78$V
d. $2Al^{+3}(aq) + 3Cd(s) \rightarrow 2Al(s) + 3Cd^{+2}(aq)$ $E°_{cell} = 1.26$ V
e. $2Al^{+3}(aq) + 3Cd(s) \rightarrow 2Al(s) + 3Cd^{+2}(aq)$ $E°_{cell} = 2.10$ V

7. Use the following standard electrode potentials to predict the species formed at the electrodes in the electrolysis of aqueous $CuSO_4$.
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ $E° = +1.23$ V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $E° = +0.34$ V
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O(l)$ $E° = 0.20$ V
$2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E° = 0.00$ V

a. H_2, O_2, H^+
b. Cu, O_2, H^+
c. Cu, H_2
d. H_2, H_2SO_3, H_2O
e. H_2SO_3, H_2O, O_2, H^+

8. A constant current was passed through a solution of $KAuCl_4$ between gold electrodes. Over a period of 20.00 min, the cathode increased in mass by 2.664 g. What was the current in amperes?
($F = 96500$ C/mol)
Cathode half-reaction: $AuCl_4^-(aq) + 3e^- \rightarrow Au(s) + 4Cl^-(aq)$

a. 1.08 A
b. 3.26 A
c. 2.17 A
d. 6.52 A
e. 3.48 A

9. A voltaic cell is constructed from the following half-cells, linked by a KCl salt bridge:
(a) an Fe electrode in 1.0 *M* $FeCl_2$ solution
(b) a Ni electrode in 1.0 *M* $Ni(NO_3)_2$ solution
Use the table of standard electrode potentials in your textbook to decide which one of the following statements is correct.

a. The Ni electrode is the anode.
b. Electrons flow from the iron electrode to the nickel electrode.
c. The iron electrode is positively charged.
d. The iron electrode will gain mass when current flows.
e. The salt bridge conducts electrons through solution.

10. Which one of the following reactions must be carried out in an electrolytic cell, rather than a voltaic cell?

a. $Zn + Cd^{2+} \rightarrow Cd + Zn^{2+}$
b. $Al + 3/2Br_2 \rightarrow Al^{3+} + 3Br^-$
c. $2Al^{3+} + 3Fe \rightarrow 2Al + 3Fe^{2+}$
d. $H_2 + I_2 \rightarrow 2H^+ + 2I^-$
e. $2H_2 + O_2 \rightarrow 2H_2O$

11. How many minutes does it take to plate 0.800 g of silver metal onto a serving tray from an aqueous solution of $AgNO_3$ at a current of 2.50 A? (F = 96500 C/mol)

a. < 2 minutes
b. 2.38 minutes
c. 4.77 minutes
d. 9.54 minutes
e. 23.8 minutes

General Chemistry Answers

Electronic Structure and Periodic Table

1. (e)

2. (e)

3. (a)

4. (c)

5. (c)

6. (b)

7. (a)

8. (a)

9. (a)

10. (d)

11. (b)

Bonding

1. (d)

2. (c)

3. (c)

4. (c)

5. (c)

6. (b)

7. (c)

8. (c)

9. (e)

10. (b)

11. (d)

Phases and Phase Equilibria

1. (b)

2. (e)

3. (b)

4. (c)

5. (c)

6. (d)

7. (d)

8. (c)

9. (d)

10. (e)

Stoichiometry

1. (e)

2. (b)

3. (c)

4. (a)

5. (c)

6. (b)

7. (c)

8. (c)

9. (d)

10. (d)

11. (c)

Thermodynamics and Thermochemistry

1. (d)

2. (c)

3. (d)

4. (d)

5. (c)

6. (d)

7. (c)

8. (d)

9. (c)

10. (a)

11. (a)

Rate Processes in Chemical Reactions—Kinetics and Equilibrium

1. (c)

2. (c)

3. (d)

4. (a)

5. (d)

6. (e)

7. (c)

8. (c)

9. (b)

10. (b)

11. (d)

Solution Chemistry

1. (e)

2. (e)

3. (a)

4. (c)

5. (d)

6. (c)

7. (b)

8. (d)

9. (e)

10. (b)

11. (b)

Acids and Bases

1. (d)

2. (d)

3. (b)

4. (d)

5. (c)

6. (a)

7. (b)

8. (e)

9. (e)

10. (b)

11. (a)

Electrochemistry

1. (a)

2. (a)

3. (c)

4. (d)

5. (d)

6. (b)

7. (b)

8. (b)

9. (b)

10. (c)

11. (c)