Exam 2

Multiple Choice

Identify the choice that best completes the statement or answers the question.

1. A compound dissolves in water to form an aqueous solution that is a poor conductor of electricity. The compound may be composed of
   a. covalent molecules.
   b. a strong base.
   c. ions.
   d. a soluble salt.
   e. a strong acid.

2. Which of the following compounds is insoluble in water?
   a. NH₄Br
   b. KBr
   c. ZnCl₂
   d. PbBr₂
   e. LiBr

3. When a solution of lithium chloride and a solution of ammonium sulfate are mixed,
   a. a new salt is formed.
   b. no reaction occurs.
   c. a precipitate forms.
   d. an acid and a base are formed.
   e. a gas is evolved.

4. Aqueous solutions of sodium sulfide and copper(II) chloride are mixed together. Which statement is correct?
   a. CuS will precipitate from solution.
   b. NaCl will precipitate from solution.
   c. No precipitate will form.
   d. Both NaCl and CuS will precipitate from solution.
   e. No reaction will occur.

5. Which of the following reactions involves neither oxidation nor reduction?
   a. N₂ + 3H₂ → 2NH₃
   b. NH₄NO₂ → N₂ + 2H₂O
   c. Cu + 2Ag⁺ → Cu²⁺ + 2Ag
   d. 2CrO₄²⁻ + 2H⁺ → Cr₂O₇²⁻ + H₂O
   e. C₂H₄ + H₂ → C₂H₆
6. The sum of the oxidation numbers of all the atoms in the dichromate ion, \( \text{Cr}_2\text{O}_7^{2-} \), is
   a. 0.
   b. +4.
   c. +2.
   d. +6.
   e. -2.

7. Which of the following species would be expected to function as a reducing agent?
   a. \( \text{Ba}^{2+} \)
   b. \( \text{Zn}^{2+} \)
   c. \( \text{ClO}_4^- \)
   d. \( \text{Cs}^+ \)
   e. \( \text{V}^{2+} \)

8. All of the following reactions are described as decomposition reactions except
   a. \( \text{PCl}_5(g) \rightarrow \text{PCl}_3(g) + \text{Cl}_2(g) \).
   b. \( 2\text{H}_2\text{O}(g) \rightarrow 2\text{H}_2(g) + \text{O}_2(g) \).
   c. \( \text{BaCl}_2 \cdot 2\text{H}_2\text{O}(s) \rightarrow \text{BaCl}_2(s) + 2\text{H}_2\text{O}(g) \).
   d. \( \text{CH}_4(g) + \text{Cl}_2(g) \rightarrow \text{CH}_3\text{Cl}(g) + \text{HCl}(g) \).
   e. \( \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \).

9. In basic solution, \( \text{H}_2\text{O}_2 \) oxidizes \( \text{Cr}^{3+} \) to \( \text{CrO}_4^{2-} \) and is reduced to \( \text{OH}^- \). What is the lowest whole-number coefficient for \( \text{OH}^- \) in the balanced net ionic equation?
   a. 6
   b. 10
   c. 4
   d. 16
   e. 8

10. A 29.0-g sample of NaOH is dissolved in water, and the solution is diluted to give a final volume of 1.60 L. The molarity of the final solution is
    a. 18.1 \( M \).
    b. 0.453 \( M \).
    c. 0.725 \( M \).
    d. 0.0552 \( M \).
    e. 0.862 \( M \).
11. In a volumetric analysis experiment, a solution of sodium oxalate (Na$_2$C$_2$O$_4$) in acidic solution is titrated with a solution of potassium permanganate (KMnO$_4$) according to the following balanced chemical equation:

$$2\text{KMnO}_4(aq) + 8\text{H}_2\text{SO}_4(aq) + 5\text{Na}_2\text{C}_2\text{O}_4(aq) \rightarrow 2\text{MnSO}_4(aq) + 8\text{H}_2\text{O}(l) + 10\text{CO}_2(g) + 5\text{Na}_2\text{SO}_4(aq) + \text{K}_2\text{SO}_4(aq)$$

It required 25.0 mL of 0.0448 $M$ KMnO$_4$ to reach the endpoint. What mass of Na$_2$C$_2$O$_4$ was present initially?

a. 2.40 g  
b. 0.0600 g  
c. 15.0 g  
d. 0.150 g  
e. 0.375 g

12. The energy associated with a motionless rock on the top of Mount Shasta is

a. potential energy.  
b. temperature.  
c. heat.  
d. internal energy.  
e. kinetic energy.

13. For which of the following reactions is $\Delta H$ not equal to $\Delta U$?

a. $\text{C(s)} + \text{O}_2(g) \rightarrow \text{CO}_2(g)$  
b. $\text{Hg(s)} \rightarrow \text{Hg}(l)$  
c. $2\text{HF}(g) \rightarrow \text{H}_2(g) + \text{F}_2(g)$  
d. $\text{CH}_4(g) + \text{C}_2\text{H}_2(g) \rightarrow 2\text{C}_2\text{H}_6(g)$  
e. $\text{I}_2(s) \rightarrow \text{I}_2(g)$

14. Given:

$$60\text{C}(s) \rightarrow \text{C}_{60}(s); \Delta H = 2320 \text{ kJ}$$

what is $\Delta H$ for the following thermochemical equation?

$$\frac{1}{60}\text{C}_{60}(s) \rightarrow \text{C}(s)$$

a. +38.7 kJ  
b. +2320 kJ  
c. −139 MJ  
d. −2320 kJ  
e. −38.7 kJ
15. It is relatively easy to change the temperature of a substance that
   a. is very massive.
   b. is an insulator.
   c. has a high specific heat capacity.
   d. has a low specific heat capacity.
   e. is brittle.

16. How much heat is gained by iron when 22.6 g of iron is warmed from 20.7°C to 65.9°C? The specific heat of iron is 0.449 J/(g · °C).
   a. 2.10 \times 10^2 J
   b. 29.59 J
   c. 20.29 J
   d. 4.59 \times 10^2 J
   e. 6.69 \times 10^2 J

17. A 170.0-g sample of metal at 79.0°C is added to 170.0 g of H_2O(l) at 17.0°C in an insulated container. The temperature rises to 19.9°C. Neglecting the heat capacity of the container, what is the specific heat of the metal? The specific heat of H_2O(l) is 4.18 J/(g · °C).
   a. 4.18 J/(g · °C)
   b. 84.8 J/(g · °C)
   c. 0.206 J/(g · °C)
   d. −0.206 J/(g · °C)
   e. 20.3 J/(g · °C)

18. What is ΔH^o for the following phase change?

NaCl(s) → NaCl(l)

<table>
<thead>
<tr>
<th>Substance</th>
<th>ΔH^o_f (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl(s)</td>
<td>−411.12</td>
</tr>
<tr>
<td>NaCl(l)</td>
<td>−385.92</td>
</tr>
</tbody>
</table>

   a. 797.04 kJ
   b. 25.20 kJ
   c. −797.04 kJ
   d. −25.20 kJ
   e. 0 kJ
19. What is $\Delta H^\circ$ for the following reaction?

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta H^\circ_f$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$C_2H_2(g)$</td>
<td>+226.7</td>
</tr>
<tr>
<td>$CO_2(g)$</td>
<td>-393.5</td>
</tr>
<tr>
<td>$H_2O(l)$</td>
<td>-285.8</td>
</tr>
</tbody>
</table>

a. +1692.2 kJ  
b. -452.6 kJ  
c. -1692.2 kJ  
d. +2599.0 kJ  
e. -2599.0 kJ

20. What is the standard enthalpy change for the following reaction?

$$3CH_4(g) + 4O_3(g) \rightarrow 3CO_2(g) + 6H_2O(g)$$

<table>
<thead>
<tr>
<th>Substance</th>
<th>$\Delta H^\circ_f$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$CH_4(g)$</td>
<td>-74.87</td>
</tr>
<tr>
<td>$O_3(g)$</td>
<td>+142.7</td>
</tr>
<tr>
<td>$CO_2(g)$</td>
<td>-393.5</td>
</tr>
<tr>
<td>$H_2O(g)$</td>
<td>-241.8</td>
</tr>
</tbody>
</table>

a. -2285.1 kJ  
b. -2977.5 kJ  
c. +2977.5 kJ  
d. +2285.1 kJ  
e. -3426.5 kJ

21. Calculate the change in enthalpy when 52.0 g of solid chromium at 25°C and 1 atm pressure is oxidized. ($\Delta H^\circ_f$ for $Cr_2O_3(s)$ is -1135 kJ/mol.)

$$4Cr(s) + 3O_2(g) \rightarrow 2Cr_2O_3(s)$$

a. -1135 kJ  
b. -284 kJ  
c. -568 kJ  
d. +1135 kJ  
e. +568 kJ

22. A photon of red light has a ____ frequency and a ____ wavelength than a photon of blue light.

a. lower, longer  
b. higher, shorter  
c. lower, shorter  
d. higher, longer  
e. lower, lower
23. Which of the following statements is a valid conclusion from the Heisenberg uncertainty principle?
   a. The square of the wave function is proportional to the probability of finding a particle in space.
   b. Particles can exhibit wavelike behavior.
   c. The orbits proposed by Bohr's model of the atom are correct.
   d. An electron in a 2p orbital is always closer to the nucleus than an electron in a 3p orbital.
   e. The act of measuring a particle's position changes its momentum, and vice versa.

24. All the following statements about the quantum numbers are true except
   a. \( m_l \) has \( 2l + 1 \) possible values.
   b. \( n \) may take integral values from 1 to \( \infty \).
   c. \( m_l \) may take integral values of \( +l \) to \( -l \), including zero.
   d. \( l \) may take integral values from 1 to \( n - 1 \).
   e. \( m_s \) may take only the values of \( +\frac{1}{2} \) and \( -\frac{1}{2} \).

25. Which of the following is a representation of a 1s orbital?
   a. 
   b. 
   c. 
   d. 
   e. 

Exam 2  
Answer Section  

MULTIPLE CHOICE  

1. ANS: A  
PTS: 1   
DIF: easy   
REF: 4.1   
OBJ: Explain how an electrolyte makes a solution electrically conductive.   
TOP: chemical reactions | ions in aqueous solution   
KEY: electrolyte | nonelectrolyte   
MSC: general chemistry  

2. ANS: D  
PTS: 1   
DIF: moderate   
REF: 4.1   
OBJ: Use the solubility rules. (Example 4.1)   
TOP: chemical reactions | ions in aqueous solution   
KEY: solubility rules   
MSC: general chemistry  

3. ANS: B  
PTS: 1   
DIF: difficult   
REF: 4.3   
OBJ: Recognize precipitation (exchange) reactions.   
TOP: chemical reactions | types of chemical reactions   
KEY: precipitation reaction   
MSC: general chemistry  

4. ANS: A  
PTS: 1   
DIF: easy   
REF: 4.3   
OBJ: Decide whether a precipitation reaction will occur. (Example 4.3)   
TOP: chemical reactions | types of chemical reactions   
KEY: precipitation reaction   
MSC: general chemistry  

5. ANS: D  
PTS: 1   
DIF: moderate   
REF: 4.5   
OBJ: Define an oxidation-reduction reaction.   
TOP: chemical reactions | types of chemical reactions   
KEY: oxidation-reduction reaction   
MSC: general chemistry  

6. ANS: E  
PTS: 1   
DIF: moderate   
REF: 4.5   
OBJ: Learn the oxidation-number rules.   
TOP: chemical reactions | types of chemical reactions   
KEY: oxidation-reduction reaction   
MSC: general chemistry  

7. ANS: E  
PTS: 1   
DIF: moderate   
REF: 4.5   
OBJ: Determine the species undergoing oxidation and reduction.   
TOP: chemical reactions | types of chemical reactions   
KEY: oxidation-reduction reaction | oxidation number   
MSC: general chemistry  

8. ANS: D  
PTS: 1   
DIF: easy   
REF: 4.5   
OBJ: Recognize combination reactions, decomposition reactions, displacement reactions, and combustion reactions.   
TOP: chemical reactions | types of chemical reactions   
KEY: oxidation-reduction reaction | common oxidation-reduction reactions   
MSC: general chemistry  

9. ANS: B  
PTS: 1   
DIF: difficult   
REF: 4.6   
OBJ: Balance simple oxidation-reduction reactions by the half-reaction method. (Example 4.8)   
TOP: chemical reactions | types of chemical reactions   
KEY: balancing oxidation-reduction equations | half-reaction method   
MSC: general chemistry  

10. ANS: B  
PTS: 1   
DIF: easy   
REF: 4.7   
OBJ: Calculate the molarity from mass and volume. (Example 4.9)   
TOP: chemical reactions | working with solutions   
KEY: concentration   
MSC: general chemistry  


11. ANS: E PTS: 1 DIF: moderate REF: 4.10
OBJ: Calculate the quantity of substance in a titrated solution. (Example 4.14)
TOP: chemical reactions | quantitative analysis
MSC: general chemistry

12. ANS: A PTS: 1 DIF: easy REF: 6.1
OBJ: Define energy, kinetic energy, potential energy, and internal energy.
TOP: thermochemistry | heats of reaction
MSC: general chemistry

OBJ: Explain how enthalpy and internal energy are related.
TOP: thermochemistry | heats of reaction
MSC: general chemistry

14. ANS: E PTS: 1 DIF: easy REF: 6.4
OBJ: Manipulate a thermochemical equation using these rules. (Example 6.3)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

15. ANS: D PTS: 1 DIF: easy REF: 6.6
OBJ: Relate the heat absorbed or evolved to the specific heat, mass, and temperature change.
TOP: thermochemistry | heats of reaction
MSC: general chemistry

16. ANS: D PTS: 1 DIF: moderate REF: 6.6
OBJ: Calculate using this relation between heat and specific heat. (Example 6.5)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

17. ANS: C PTS: 1 DIF: difficult REF: 6.6
OBJ: Calculate using this relation between heat and specific heat. (Example 6.5)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

18. ANS: B PTS: 1 DIF: moderate REF: 6.8
OBJ: Calculate the heat of a phase transition using standard enthalpies of formation for the different phases. (Example 6.8)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

19. ANS: E PTS: 1 DIF: moderate REF: 6.8
OBJ: Calculate the heat (enthalpy) of reaction from the standard enthalpies of formation of the substances in the reaction. (Example 6.9)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

20. ANS: B PTS: 1 DIF: moderate REF: 6.8
OBJ: Calculate the heat (enthalpy) of reaction from the standard enthalpies of formation of the substances in the reaction. (Example 6.9)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

21. ANS: C PTS: 1 DIF: moderate REF: 6.8
OBJ: Calculate the heat (enthalpy) of reaction from the standard enthalpies of formation of the substances in the reaction. (Example 6.9)
TOP: thermochemistry | heats of reaction
MSC: general chemistry

22. ANS: A PTS: 1 DIF: moderate REF: 7.1
OBJ: Describe the different regions of the electromagnetic spectrum.
TOP: atomic theory | light
MSC: general chemistry
23. ANS: E  PTS: 1  DIF: moderate  REF: 7.4
OBJ: State Heisenberg's uncertainty principle.
KEY: wave functions | Heisenberg's uncertainty principle
TOP: atomic theory | quantum mechanics
MSC: general chemistry

24. ANS: D  PTS: 1  DIF: easy  REF: 7.5
OBJ: State the rules for the allowed values for each quantum number.
TOP: atomic theory | quantum mechanics
KEY: quantum numbers
MSC: general chemistry

25. ANS: C  PTS: 1  DIF: easy  REF: 7.5
OBJ: Describe the shapes of s, p, and d orbitals.
TOP: atomic theory | quantum mechanics
KEY: atomic orbital shape
MSC: general chemistry
Exam 2 sample 2

Multiple Choice

Identify the choice that best completes the statement or answers the question.

1. Which compounds will not dissolve in water in large amounts at 20°C?
   
   \[
   \begin{array}{cccc}
   \text{Ca(NO}_3\text{)}_2 & \text{MgCO}_3 & \text{PbI}_2 & \text{K}_2\text{SO}_4 \\
   \text{I.} & \text{II.} & \text{III.} & \text{IV.} \\
   \end{array}
   \]
   
   a. II & III
   b. III & IV
   c. III & IV
   d. I & II
   e. I & IV

2. Which substance is oxidized in the reaction below?
   \[
   \text{NaNO}_3 + \text{Pb} \rightarrow \text{NaNO}_2 + \text{PbO}
   \]
   
   a. Na+
   b. NaNO2
   c. NaN03
   d. PbO
   e. none of these choices

3. In which of the following compounds does the sulfur atom have the highest (i.e., most positive) oxidation number?
   
   a. Na2SO4
   b. H2S
   c. SO2
   d. H2SO3
   e. S8

4. How many moles of NaOH are present in 25.0 mL of a 0.1000 M NaOH solution?
   
   a. 100 mol
   b. 2.50 mol
   c. 2.50 \times 10^{-3} \text{ mol}
   d. 25.0 mol
   e. 0.100 mol

5. Determine the ammonium ion concentration of a solution that results when 4.53 g of (NH4)2SO4 is dissolved in water and diluted to exactly 100.0 mL.
   
   a. 1.37 M
   b. 0.686 M
   c. 2.51 M
   d. 1.03 M
   e. 0.343 M
6. If a 45.0 mL sample of 2.20 M Na₂SO₄ is diluted to yield a final solution that is 0.110 M in sodium ions, what is the volume of the final solution?
   a. 4500 mL
   b. 900 mL
   c. 450 mL
   d. 1800 mL
   e. 110 mL

7. How many grams of KCl are in 125.0 mL of 0.375 M KCl?
   a. 3.49 x 10⁻³ g
   b. 46.9 g
   c. 0.938 g
   d. 3.49 g
   e. 0.0469 g

8. A solution is prepared by dissolving 20.0 g of NaI in enough water to make 300.0 mL of solution. How many moles of ions are in 25.0 mL of this solution?
   a. 0.0222 mol
   b. 0.445 mol
   c. 0.0445 mol
   d. 0.111 mol
   e. 0.0111 mol

9. A solution is made by dissolving 60.0 g of AlCl₃ in enough water to make 250.0 mL of solution. How many moles of chloride ions are in 5.00 mL of solution?
   a. 1.25 x 10⁻³ mol
   b. 9.00 x 10⁻³ mol
   c. 3.00 x 10⁻³ mol
   d. 1.80 x 10⁻² mol
   e. 2.70 x 10⁻² mol

10. A 25.00 mL sample of HCl solution is neutralized by exactly 31.22 mL of 0.08152 M Ca(OH)₂. What is the molarity of the HCl solution?
    a. 0.08152 M
    b. 0.1018 M
    c. 0.09453 M
    d. 0.1021 M
    e. 0.2036 M

11. A 25.00 mL sample of H₃PO₄ solution is neutralized by exactly 54.93 mL of 0.04345 M Ca(OH)₂. What is the molarity of the H₃PO₄ solution?
    a. 0.09546 M
    b. 0.2897 M
    c. 0.2148 M
    d. 0.0636 M
    e. 0.1432 M
12. Which of the following is an example of potential energy?
   a. a rock rolling down a mountain
   b. atoms vibrating back and forth around a specific location
   c. electrons flowing through an electrical conductor
   d. a fat molecule stored in the body
   e. the sound of a dog barking

13. A 20.0 g sample of aluminum (specific heat \(0.902 \text{ J g}^{-1} \text{ °C}^{-1}\)) with an initial temperature of 48.6°C is heated with 427 J of energy. What is the final temperature of the sample?
   a. 23.7°C
   b. 74.8°C
   c. 72.3°C
   d. 26.2°C
   e. 24.9°C

14. What is the enthalpy change when 175 g of \(\text{C}_3\text{H}_8\) are burned in excess \(\text{O}_2\)?

   \[
   \text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{l}) \quad \Delta H^\circ = -2220 \text{ kJ}
   \]

   a. \(-3.89 \times 10^5 \text{ kJ}\)
   b. \(-1.79 \times 10^3 \text{ kJ}\)
   c. \(-3.47 \times 10^0 \text{ kJ}\)
   d. \(-1.71 \times 10^7 \text{ kJ}\)
   e. \(-8.82 \times 10^3 \text{ kJ}\)

15. Determine the heat of reaction for the process

   \[
   \text{C}_2\text{H}_4(\text{g}) + 6\text{HCl}(\text{g}) \rightarrow 2\text{CHCl}_3(\text{g}) + 4\text{H}_2(\text{g})
   \]

   using the information given below:

   \[
   2\text{C}(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g}) \quad \Delta H^\circ = 52.3 \text{ kJ}
   \]

   \[
   \text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g}) \quad \Delta H^\circ = -184.6\text{kJ}
   \]

   \[
   \text{C}(\text{s}) + 1/2\text{H}_2(\text{g}) + 3/2\text{Cl}_2(\text{g}) \rightarrow \text{CHCl}_3(\text{g}) \quad \Delta H^\circ = -103.1 \text{ kJ}
   \]

   a. \(-29.2 \text{ kJ}\)
   b. \(-295.3 \text{ kJ}\)
   c. \(+398.4 \text{ kJ}\)
   d. \(+295.3 \text{ kJ}\)
   e. \(+29.2 \text{ kJ}\)

16. Light has a wavelength of 582 nm. What is its frequency? The speed of light is \(3.00 \times 10^8 \text{ m/s}\).

   a. \(1.94 \times 10^3 \text{ Hz}\)
   b. \(5.15 \times 10^{14} \text{ Hz}\)
   c. \(1.75 \times 10^2 \text{ Hz}\)
   d. \(1.75 \times 10^{20} \text{ Hz}\)
   e. \(1.94 \times 10^{13} \text{ Hz}\)
17. What is the phenomenon that occurs when excited gaseous elements emit only a few colored lines?
   a. Planck's constant
   b. photoelectric effect
   c. electromagnetic spectrum
   d. quantum theory
   e. line spectrum

18. According to the Bohr model for the hydrogen atom, the energy necessary to excite an electron from \( n = 2 \) to \( n = 3 \) is ______ the energy necessary to excite an electron from \( n = 3 \) to \( n = 4 \).
   a. either more than or equal to
   b. equal to
   c. less than
   d. more than
   e. either less than or equal to

19. If \( l = 2 \), what value can \( m_l \) have?
   a. \( m_l = +1 \)
   b. \( m_l = 0, +1, +2 \)
   c. \( m_l = +2 \)
   d. \( m_l = -1, 0, +1 \)
   e. \( m_l = -2, -1, 0, +1, +2 \)

20. How many electrons can the third principal quantum level hold?
   a. 18
   b. 8
   c. 2
   d. 16
   e. 32

21. What is the correct electron configuration for beryllium (Be)?
   a. \( 1s^22s^22p^6 \)
   b. \( 1s^22s^2 \)
   c. \( 1s^22s^22s^2 \)
   d. \( 1s^22s^22p^2 \)
   e. \( 1s^22s^22p^6 \)

22. What is the electron configuration of arsenide ion, As\(^{3-}\)?
   a. \( 1s^22s^22p^63s^23p^63d^104s^24p^6 \)
   b. \( 1s^22s^22p^63s^23p^64s^24d^10 \)
   c. \( 1s^22s^22p^63s^23p^63d^104s^2 \)
   d. \( 1s^22s^22p^63s^23p^63d^10 \)
   e. \( 1s^22s^22p^63s^23p^64s^24d^104p^6 \)

23. What is the electron configuration of Li\(^+\)?
   a. \( 1s^1 \)
   b. \( 1s^22s^12p^1 \)
   c. \( 1s^22s^2 \)
   d. \( 1s^2 \)
   e. \( 1s^22s^1 \)
24. Substances that have the same electron configuration are:
   a. lanthanides.
   b. paramagnetic.
   c. isoelectronic.
   d. diamagnetic.
   e. ferromagnetic.

25. Which statement is false?
   a. Mg$^{2+}$ is smaller than Ca$^{2+}$.
   b. Cl$^-$ is smaller than S$^{2-}$.
   c. N$^{3-}$ is larger than Ne.
   d. Br$^-$ is larger than Kr.
   e. K$^+$ is smaller than Ar.
Exam 2 sample 2
Answer Section

MULTIPLE CHOICE

1. ANS: A   PTS: 1
   TOP: 5.1 Exchange Reactions: Precipitation and Net Ionic Equations
2. ANS: E   PTS: 1
   TOP: 5.3 Oxidation-Reduction Reactions
3. ANS: A   PTS: 1
   TOP: 5.4 Oxidation Numbers and Redox Reactions
4. ANS: C   PTS: 1
   TOP: 5.6 Solution Concentration
5. ANS: B   PTS: 1
   TOP: 5.6 Solution Concentration
6. ANS: D   PTS: 1
   TOP: 5.6 Solution Concentration
7. ANS: D   PTS: 1
   TOP: 5.6 Solution Concentration
8. ANS: A   PTS: 1
   TOP: 5.6 Solution Concentration
9. ANS: E   PTS: 1
   TOP: 5.6 Solution Concentration
10. ANS: E  PTS: 1
    TOP: 5.8 Aqueous Solution Titrations
11. ANS: D  PTS: 1
    TOP: 5.8 Aqueous Solution Titrations
12. ANS: D  PTS: 1
    TOP: 6.1 The Nature of Energy
13. ANS: C  PTS: 1
    TOP: 6.3 Heat Capacity
14. ANS: E  PTS: 1
    TOP: 6.6 Enthalpy Changes for Chemical Reactions
15. ANS: D  PTS: 1
    TOP: 6.9 Hess's Law
16. ANS: B  PTS: 1
    TOP: 7.1 Electromagnetic Radiation and Matter
17. ANS: E  PTS: 1
    TOP: 7.2 Planck's Quantum Theory
18. ANS: D  PTS: 1
    TOP: 7.3 The Bohr Model of the Hydrogen Atom
19. ANS: E  PTS: 1
    TOP: 7.5 Quantum Numbers, Energy Levels and Atomic Orbitals
20. ANS: A  PTS: 1
    TOP: 7.5 Quantum Numbers, Energy Levels and Atomic Orbitals
21. ANS: B  PTS: 1
    TOP: 7.7 Atom Electron Configurations
22. ANS: A  PTS: 1
    TOP: 7.8 Ion Electron Configurations
23. ANS: D  PTS: 1
    TOP: 7.8 Ion Electron Configurations
24. ANS: C  PTS: 1
    TOP: 7.8 Ion Electron Configurations
25. ANS: E  PTS: 1
    TOP: 7.10 Periodic Trends: Ionic Radii