Chemistry

Description of the Examination

The Chemistry examination covers material that is usually taught in a one-year college course in general chemistry. Understanding of the structure and states of matter, reaction types, equations and stoichiometry, equilibrium, kinetics, thermodynamics, and descriptive and experimental chemistry is required, as is the ability to interpret and apply this material to new and unfamiliar problems. During this examination, an online scientific calculator function and a periodic table are available as part of the testing software.

The examination contains approximately 75 questions to be answered in 90 minutes. Some of these are pretest questions that will not be scored. Any time spent on tutorials and providing personal information is in addition to the actual testing time.

Knowledge and Skills Required

Questions on the Chemistry examination require candidates to demonstrate one or more of the following abilities.

- **Recall** — remember specific facts; demonstrate straightforward knowledge of information and familiarity with terminology
- **Application** — understand concepts and reformulate information into other equivalent terms; apply knowledge to unfamiliar and/or practical situations; use mathematics to solve chemistry problems
- **Interpretation** — infer and deduce from data available and integrate information to form conclusions; recognize unstated assumptions

The subject matter of the Chemistry examination is drawn from the following topics. The percentages next to the main topics indicate the approximate percentage of exam questions on that topic.

**Structure of Matter**

- Atomic theory and atomic structure
  - Evidence for the atomic theory
  - Atomic masses; determination by chemical and physical means
  - Atomic number and mass number: isotopes and mass spectroscopy
  - Electron energy levels: atomic spectra, quantum numbers, atomic orbitals
  - Periodic relationships, including, for example, atomic radii, ionization energies, electron affinities, oxidation states
- Chemical bonding
  - Binding forces
    - Types: covalent, ionic, metallic, macromolecular (or network), dispersion, hydrogen bonding
    - Relationships to structure and to properties
    - Polarity of bonds, electronegativities
  - Geometry of molecules, ions and coordination complexes; structural isomerism, dipole moments of molecules, relation of properties to structure
  - Molecular models
    - Valence bond theory; hybridization of orbitals, resonance, sigma and pi bonds
    - Other models; for example, molecular orbital
- Nuclear chemistry: nuclear equations, half lives, and radioactivity; chemical applications
<table>
<thead>
<tr>
<th>19%</th>
<th>States of Matter</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Gases</strong></td>
<td>Laws of ideal gases: equations of state for an ideal gas</td>
</tr>
<tr>
<td></td>
<td>Kinetic molecular theory</td>
</tr>
<tr>
<td></td>
<td>- Interpretation of ideal gas laws on the basis of this theory</td>
</tr>
<tr>
<td></td>
<td>- The mole concept; Avogadro’s number</td>
</tr>
<tr>
<td></td>
<td>- Dependence of kinetic energy of molecules on temperature: Boltzmann distribution</td>
</tr>
<tr>
<td></td>
<td>- Deviations from ideal gas laws</td>
</tr>
<tr>
<td><strong>Liquids and solids</strong></td>
<td>Liquids and solids from the kinetic-molecular viewpoint</td>
</tr>
<tr>
<td></td>
<td>Phase diagrams of one component systems</td>
</tr>
<tr>
<td></td>
<td>Changes of state, critical phenomena</td>
</tr>
<tr>
<td></td>
<td>Crystal structure</td>
</tr>
<tr>
<td><strong>Solutions</strong></td>
<td>Types of solutions and factors affecting solubility</td>
</tr>
<tr>
<td></td>
<td>Methods of expressing concentration</td>
</tr>
<tr>
<td></td>
<td>Colligative properties; for example, Raoult’s law</td>
</tr>
<tr>
<td></td>
<td>Effect of interionic attraction on colligative properties and solubility</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>10%</th>
<th>Equations and Stoichiometry</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ionic and molecular species present in chemical systems: net ionic equations</td>
<td></td>
</tr>
<tr>
<td>Stoichiometry: mass and volume relations with emphasis on the mole concept</td>
<td></td>
</tr>
<tr>
<td>Balancing of equations, including those for redox reactions</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>7%</th>
<th>Equilibrium</th>
</tr>
</thead>
<tbody>
<tr>
<td>Concept of dynamic equilibrium, physical and chemical; LeChâtelier’s principle: equilibrium constants</td>
<td></td>
</tr>
<tr>
<td>Quantitative treatment</td>
<td></td>
</tr>
<tr>
<td>- Equilibrium constants for gaseous reactions in terms of both molar concentrations and partial pressure (K_p, K_c)</td>
<td></td>
</tr>
<tr>
<td>- Equilibrium constants for reactions in solutions</td>
<td></td>
</tr>
<tr>
<td>- Constants for acids and bases: pK; pH</td>
<td></td>
</tr>
<tr>
<td>- Solubility product constants and their application to precipitation and the dissolution of slightly soluble compounds</td>
<td></td>
</tr>
<tr>
<td>- Constants for complex ions</td>
<td></td>
</tr>
<tr>
<td>- Common ion effect; buffers</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>4%</th>
<th>Kinetics</th>
</tr>
</thead>
<tbody>
<tr>
<td>Concept of rate of reaction</td>
<td></td>
</tr>
<tr>
<td>Order of reaction and rate constant; their determination from experimental data</td>
<td></td>
</tr>
<tr>
<td>Effect of temperature change on rates</td>
<td></td>
</tr>
<tr>
<td>Energy of activation: the role of catalysts</td>
<td></td>
</tr>
<tr>
<td>The relationship between the rate determining step and a mechanism</td>
<td></td>
</tr>
</tbody>
</table>
5% Thermodynamics
State functions
First law: heat of formation; heat of reaction; change in enthalpy. Hess's law: heat capacity; heats of vaporization and fusion
Second law: free energy of formation; free energy of reaction; dependence of change in free energy on enthalpy and entropy changes
Relationship of change in free energy to equilibrium constants and electrode potentials

14% Descriptive Chemistry
The accumulation of certain specific facts of chemistry is essential to enable students to comprehend the development of principles and concepts, to demonstrate applications of principles, to relate fact to theory and properties to structure, and to develop an understanding of systematic nomenclature that facilitates communication. The following areas are normally included on the examination:
• Chemical reactivity and products of chemical reactions
• Relationships in the periodic table: horizontal, vertical and diagonal
• Chemistry of the main groups and transition elements, including typical examples of each
• Organic chemistry, including such topics as functional groups and isomerism (may be treated as a separate unit or as exemplary material in other areas, such as bonding)

9% Experimental Chemistry
Some questions are based on laboratory experiments widely performed in general chemistry and ask about the equipment used, observations made, calculations performed, and interpretation of the results. The questions are designed to provide a measure of understanding of the basic tools of chemistry and their applications to simple chemical systems.
Sample Test Questions

The following sample questions do not appear on an actual CLEP examination. They are intended to give potential test takers an indication of the format and difficulty level of the examination and to provide content for practice and review. Knowing the correct answers to all of the sample questions is not a guarantee of satisfactory performance on the exam.

Note: For all questions involving solutions and/or chemical equations, assume that the system is in pure water and at room temperature unless otherwise stated.

Part A

Directions: Each set of lettered choices below refers to the numbered questions or statements immediately following it. Select the one lettered choice that best answers each question or best fits each statement. A choice may be used once, more than once, or not at all in each set.

Questions 1–3

(A) F
(B) S
(C) Mg
(D) Ar
(E) Mn

1. Forms monatomic ions with –2 charge in solutions

2. Forms a compound having the formula KXO₄

3. Forms oxides that are common air pollutants and that yield acidic solutions in water

Questions 4–6

(A) Hydrofluoric acid
(B) Carbon dioxide
(C) Aluminum hydroxide
(D) Ammonia
(E) Hydrogen peroxide

4. Is a good oxidizing agent

5. Is used extensively for the production of fertilizers

6. Has amphoteric properties

Questions 7–8

(A) A network solid with covalent bonding
(B) A molecular solid with London (dispersion) forces only
(C) A molecular solid with hydrogen bonding
(D) An ionic solid
(E) A metallic solid

7. Solid ethyl alcohol, C₂H₅OH

8. Silicon dioxide, SiO₂
Questions 9–11

(A) CO₃²⁻
(B) MnO₄⁻
(C) NH₄⁺
(D) Ba²⁺
(E) Al³⁺

Assume that you have several "unknowns," each consisting of an aqueous solution of a salt that contains one of the ions listed above. Which ion must be present if the following observations are made of that unknown?

9. The solution is colored.

10. An odor can be detected when a sample of the solution is added drop by drop to a warm solution of sodium hydroxide.

11. A precipitate is formed when a dilute solution of H₂SO₄ is added to a sample of the solution.

Questions 12–13

The spontaneous reaction that occurs when the cell above operates is

\[ 2 \text{Ag}^+ + \text{Cd}(s) \rightarrow 2 \text{Ag}(s) + \text{Cd}^2+ \]

(A) Voltage increases.
(B) Voltage decreases but remains above zero.
(C) Voltage becomes zero and remains at zero.
(D) No change in voltage occurs.
(E) Direction of voltage change cannot be predicted without additional information.

Which of the above occurs for each of the following circumstances?

12. The silver electrode is made larger.

13. The salt bridge is replaced by a platinum wire.
Questions 17–18

\[
\text{H}_3\text{AsO}_4 + 3 \text{I}^- + 2 \text{H}_3\text{O}^+ \rightarrow \text{H}_3\text{AsO}_3 + \text{I}_3^- + 3 \text{H}_2\text{O}
\]

The oxidation of iodide ions by arsenic acid in acidic aqueous solution occurs according to the balanced equation shown above. The experimental rate law for the reaction at 25°C is

\[
\text{Rate} = k [\text{H}_3\text{AsO}_4] [\text{I}^-] [\text{H}_3\text{O}^+].
\]

17. What is the order of the reaction with respect to I⁻?
   (A) 1
   (B) 2
   (C) 3
   (D) 5
   (E) 6

18. According to the rate law for the reaction, an increase in the concentration of the hydronium ion has what effect on the reaction at 25°C?
   (A) The rate of reaction increases.
   (B) The rate of reaction decreases.
   (C) The value of the equilibrium constant increases.
   (D) The value of the equilibrium constant decreases.
   (E) Neither the rate nor the value of the equilibrium constant is changed.

19. The critical temperature of a substance is the
   (A) temperature at which the vapor pressure of the liquid is equal to the external pressure
   (B) temperature at which the vapor pressure of the liquid is equal to 760 mm Hg
   (C) temperature at which the solid, liquid, and vapor phases are all in equilibrium
   (D) temperature at which the liquid and vapor phases are in equilibrium at 1 atmosphere
   (E) lowest temperature above which a substance cannot be liquefied at any applied pressure
20. \[ \text{Cu}(s) + 2 \text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2 \text{Ag}(s) \]

If the equilibrium constant for the reaction above is \(3.7 \times 10^{15}\), which of the following correctly describes the standard voltage, \(E^o\), and the standard free energy change, \(\Delta G^o\), for this reaction?

(A) \(E^o\) is positive and \(\Delta G^o\) is negative.
(B) \(E^o\) is negative and \(\Delta G^o\) is positive.
(C) \(E^o\) and \(\Delta G^o\) are both positive.
(D) \(E^o\) and \(\Delta G^o\) are both negative.
(E) \(E^o\) and \(\Delta G^o\) are both zero.

21. When \(^{214}\text{Po}\) decays, the emission consists consecutively of an \(\alpha\) particle, then two \(\beta\) particles, and finally another \(\alpha\) particle. The resulting stable nucleus is

(A) \(^{206}_{83}\text{Bi}\)
(B) \(^{210}_{83}\text{Bi}\)
(C) \(^{208}_{82}\text{Pb}\)
(D) \(^{208}_{82}\text{Pb}\)
(E) \(^{210}_{81}\text{Ti}\)

22. The pH of 0.1 \(M\) ammonia is approximately

(A) 1
(B) 4
(C) 7
(D) 11
(E) 14

23. \[ \ldots \text{CrO}_2^- + \ldots \text{OH}^- \rightarrow \]
\[ \ldots \text{CrO}_4^{2-} + \ldots \text{H}_2\text{O} + \ldots e^- \]

When the equation for the half reaction above is balanced, what is the ratio of the coefficients \(\text{OH}^- : \text{CrO}_2^-\)?

(A) 1 : 1
(B) 2 : 1
(C) 3 : 1
(D) 4 : 1
(E) 5 : 1

24. \[ \text{CuO}(s) + \text{H}_2(g) \leftrightharpoons \text{Cu}(s) + \text{H}_2\text{O}(g) \quad \Delta H = -2.0 \text{ kJ} \]

The substances in the equation above are at equilibrium at pressure \(P\) and temperature \(T\). The equilibrium can be shifted to favor the products by

(A) increasing the pressure by means of a moving piston at constant \(T\)
(B) increasing the pressure by adding an inert gas such as nitrogen
(C) decreasing the temperature
(D) allowing some gases to escape at constant \(P\) and \(T\)
(E) adding a catalyst

25. The molality of the glucose in a 1.0 \(M\) glucose solution can be obtained by using which of the following?

(A) Solubility of glucose in water
(B) Degree of dissociation of glucose
(C) Volume of the solution
(D) Temperature of the solution
(E) Density of the solution

26. The geometry of the \(\text{SO}_3\) molecule is best described as

(A) trigonal planar
(B) trigonal pyramidal
(C) square pyramidal
(D) bent
(E) tetrahedral

27. Which of the following molecules has the shortest bond length?

(A) \(\text{N}_2\)
(B) \(\text{O}_2\)
(C) \(\text{Cl}_2\)
(D) \(\text{Br}_2\)
(E) \(\text{I}_2\)
28. What number of moles of O_2 is needed to produce 14.2 grams of P_4O_{10} (molar mass 284 g) from P_4?
   (A) 0.0500 mole 
   (B) 0.0625 mole 
   (C) 0.125 mole 
   (D) 0.250 mole 
   (E) 0.500 mole 

29. If 0.060 faraday is passed through an electrolytic cell containing a solution of In^{3+} ions, the maximum number of moles of In that could be deposited at the cathode is 
   (A) 0.010 mole 
   (B) 0.020 mole 
   (C) 0.030 mole 
   (D) 0.060 mole 
   (E) 0.18 mole 

30. CH_4(g) + 2 O_2(g) → CO_2(g) + 2 H_2O(l) 
    \[ \Delta H^{\circ}_{\text{rxn}} = -889.1 \text{ kJ mol}^{-1} \] 
    \[ \Delta H^{\circ}_{\text{H}_2\text{O}(l)} = -285.8 \text{ kJ mol}^{-1} \] 
    \[ \Delta H^{\circ}_{\text{CO}_2(g)} = -393.3 \text{ kJ mol}^{-1} \] 
    What is the standard heat of formation, \( \Delta H^{\circ} \), of methane, CH_4(g), as calculated from the data above? 
   (A) -210.0 kJ mol^{-1} 
   (B) -107.5 kJ mol^{-1} 
   (C) -75.8 kJ mol^{-1} 
   (D) 75.8 kJ mol^{-1} 
   (E) 210.0 kJ mol^{-1} 

31. Each of the following can act as both a Bronsted acid and a Bronsted base EXCEPT 
   (A) HCO_3^- 
   (B) H_3PO_4^- 
   (C) NH_4^+ 
   (D) H_2O 
   (E) HS^- 

32. Two flexible containers for gases are at the same temperature and pressure. One holds 0.50 gram of hydrogen and the other holds 8.0 grams of oxygen. Which of the following statements regarding these gas samples is FALSE? 
   (A) The volume of the hydrogen container is the same as the volume of the oxygen container. 
   (B) The number of molecules in the hydrogen container is the same as the number of molecules in the oxygen container. 
   (C) The density of the hydrogen sample is less than that of the oxygen sample. 
   (D) The average kinetic energy of the hydrogen molecules is the same as the average kinetic energy of the oxygen molecules. 
   (E) The average speed of the hydrogen molecules is the same as the average speed of the oxygen molecules. 

33. \( \pi \) bonding occurs in each of the following species EXCEPT 
   (A) CO_2 
   (B) C_2H_4 
   (C) CN^- 
   (D) C_6H_6 
   (E) CH_4 

34. 3 Ag(s) + 4 HNO_3 → 3 AgNO_3 + NO(g) + 2 H_2O 
    The reaction of silver metal and dilute nitric acid proceeds according to the equation above. If 0.10 mole of powdered silver is added to 10. milliliters of 6.0 molar nitric acid, the number of moles of NO gas that can be formed is 
   (A) 0.015 mole 
   (B) 0.020 mole 
   (C) 0.030 mole 
   (D) 0.045 mole 
   (E) 0.090 mole
35. Which, if any, of the following species are in the greatest concentration in a 0.100 M solution of H₂SO₄ in water?

(A) H₂SO₄ molecules
(B) H₃O⁺ ions
(C) HSO₄⁻ ions
(D) SO₄²⁻ ions
(E) All species are in equilibrium and therefore have the same concentrations.

36. At 20.°C, the vapor pressure of toluene is 22 mm Hg and that of benzene is 75 mm Hg. An ideal solution, equimolar in toluene and benzene, is prepared. At 20.°C, what is the mole fraction of benzene in the vapor in equilibrium with this solution?

(A) 0.23
(B) 0.29
(C) 0.50
(D) 0.77
(E) 0.83

37. Which of the following aqueous solutions has the highest boiling point?

(A) 0.10 M potassium sulfate, K₂SO₄
(B) 0.10 M hydrochloric acid, HCl
(C) 0.10 M ammonium nitrate, NH₄NO₃
(D) 0.10 M magnesium sulfate, MgSO₄
(E) 0.20 M sucrose, C₁₂H₂₂O₁₁

38. When 70 milliliters of 3.0 M Na₂CO₃ is added to 30 milliliters of 1.0 M NaHCO₃, the resulting concentration of Na⁺ is

(A) 2.0 M
(B) 2.4 M
(C) 4.0 M
(D) 4.5 M
(E) 7.0 M

39. Which of the following species CANNOT function as an oxidizing agent?

(A) Cr₂O₇²⁻
(B) MnO₄⁻
(C) NO₃⁻
(D) S
(E) I⁻

40. A student wishes to prepare 2.00 liters of 0.100 M KO₃ solution. The proper procedure is to weigh out

(A) 42.8 grams of KO₃ and add 2.00 kilograms of H₂O
(B) 42.8 grams of KO₃ and add H₂O until the final homogeneous solution has a volume of 2.00 liters
(C) 21.4 grams of KO₃ and add H₂O until the final homogeneous solution has a volume of 2.00 liters
(D) 42.8 grams of KO₃ and add 2.00 liters of H₂O
(E) 21.4 grams of KO₃ and add 2.00 liters of H₂O

41. A 20.0-milliliter sample of 0.200 M K₂CO₃ solution is added to 30.0 milliliters of 0.400 M Ba(NO₃)₂ solution. Barium carbonate precipitates. The concentration of barium ion, Ba²⁺, in solution after reaction is

(A) 0.150 M
(B) 0.160 M
(C) 0.200 M
(D) 0.240 M
(E) 0.267 M
42. One of the outermost electrons in a strontium atom in the ground state can be described by which of the following sets of four quantum numbers?

(A) 5, 2, 0, \( \frac{1}{2} \)

(B) 5, 1, 1, \( \frac{1}{2} \)

(C) 5, 1, 0, \( \frac{1}{2} \)

(D) 5, 0, 1, \( \frac{1}{2} \)

(E) 5, 0, 0, \( \frac{1}{2} \)

43. Which of the following reactions does NOT proceed significantly to the right in aqueous solutions?

(A) \( \text{H}_3\text{O}^+ + \text{OH}^- \rightarrow 2 \text{H}_2\text{O} \)

(B) \( \text{HCN} + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{CN}^- \)

(C) \( \text{Cu}^{2+} + 4 \text{NH}_3 \rightarrow \text{Cu}^{2+} + 4 \text{H}_2\text{O} + 4 \text{NH}_3 \)

(D) \( \text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{O}^+ + \text{HSO}_4^- \)

(E) \( \text{H}_2\text{O} + \text{HSO}_4^- \rightarrow \text{H}_2\text{SO}_4 + \text{OH}^- \)

44. A compound is heated to produce a gas whose molar mass is to be determined. The gas is collected by displacing water in a water filled flask inverted in a trough of water. Which of the following is necessary to calculate the molar mass of the gas but does not need to be measured during the experiment?

(A) Mass of the compound used in the experiment

(B) Temperature of the water in the trough

(C) Vapor pressure of the water

(D) Barometric pressure

(E) Volume of water displaced from the flask

45. A 27.0 gram sample of an unknown hydrocarbon was burned in excess oxygen to form 88.0 grams of carbon dioxide and 27.0 grams of water. What is a possible molecular formula of the hydrocarbon?

(A) \( \text{CH}_4 \)

(B) \( \text{C}_2\text{H}_2 \)

(C) \( \text{C}_4\text{H}_6 \)

(D) \( \text{C}_2\text{H}_6 \)

(E) \( \text{C}_2\text{H}_{10} \)

46. If the acid dissociation constant, \( K_a \), for an acid HA is \( 8 \times 10^{-4} \) at 25°C, what percent of the acid is dissociated in a 0.50 M solution of HA at 25°C?

(A) 0.08%

(B) 0.2%

(C) 1%

(D) 2%

(E) 4%

47. The organic compound represented above is an example of

(A) an alcohol

(B) an aldehyde

(C) an ether

(D) an organic acid

(E) a ketone
48. Equal numbers of moles of $\text{H}_2(g)$, $\text{Ar}(g)$, and $\text{N}_2(g)$ are placed in a glass vessel at room temperature. If the vessel has a pinhole sized leak, which of the following will be true regarding the relative values of the partial pressures of the gases remaining in the vessel after some of the gas mixture has effused?

(A) $P_{\text{H}_2} < P_{\text{N}_2} < P_{\text{Ar}}$
(B) $P_{\text{H}_2} < P_{\text{Ar}} < P_{\text{N}_2}$
(C) $P_{\text{N}_2} < P_{\text{Ar}} < P_{\text{H}_2}$
(D) $P_{\text{Ar}} < P_{\text{H}_2} < P_{\text{N}_2}$
(E) $P_{\text{H}_2} = P_{\text{Ar}} = P_{\text{N}_2}$

49. Which of the following is a correct interpretation of the results of Rutherford's experiments in which gold atoms were bombarded with alpha particles?

(A) Atoms have equal numbers of positive and negative charges.
(B) Electrons in atoms are arranged in shells.
(C) Neutrons are at the center of an atom.
(D) Neutrons and protons in atoms have nearly equal mass.
(E) The positive charge of an atom is concentrated in a small region.

50. A 0.1 $M$ solution of which of the following ions is orange?

(A) $\text{Fe(H}_2\text{O)}_4^{2+}$
(B) $\text{Cu(NH}_3)_4^{2+}$
(C) $\text{Zn(OH)}_4^{2-}$
(D) $\text{Zn(NH}_3)_4^{2+}$
(E) $\text{Cr}_5\text{O}_4^{2-}$

51. In the formation of 1.0 mole of the following crystalline solids from the gaseous ions, the most energy is released by

(A) NaF
(B) MgF$_2$
(C) MgBr$_2$
(D) AlF$_3$
(E) AlBr$_3$

52. If 1 mole of a nonvolatile nonelectrolyte dissolves in 9 moles of water to form an ideal solution, what is the vapor pressure of this solution at 25°C? (The vapor pressure of pure water at 25°C is 23.8 mm Hg.)

(A) 23.8 mm Hg
(B) $\frac{9}{10}$ 23.8 mm Hg
(C) $\frac{10}{9}$ 23.8 mm Hg
(D) $\frac{1}{10}$ 23.8 mm Hg
(E) It cannot be determined from the information given.

53. $\text{MnO}_4^- (aq) + \text{NO}_2^- (aq) + \text{H}_2\text{O}(l) \rightarrow \text{MnO}_2(s) + \text{NO}_3^- (aq) + \text{OH}^- (aq)$

When the redox equation shown above is balanced by using coefficients reduced to lowest whole numbers, the coefficient for $\text{MnO}_4^-$ is

(A) 1
(B) 2
(C) 3
(D) 4
(E) 6

54. If a certain solid solute dissolves in water with the evolution of heat, which of the following is most likely to be true?

(A) The temperature of the solution decreases as the solute dissolves.
(B) The resulting solution is ideal.
(C) The solid has a large lattice energy.
(D) The solid has a large heat of fusion.
(E) The solid has a large energy of hydration.
55. A 0.1 molar aqueous solution of which of the following is neutral?

(A) NaNO₃
(B) Na₂CO₃
(C) NH₃Br
(D) KCN
(E) AlCl₃

56. Which of the following is a true statement about the halogens?

(A) Fluorine is the weakest oxidizing agent.
(B) Bromine is more electronegative than chlorine.
(C) The halide ions are larger than their respective halogen atoms.
(D) Adding I₂(s) to a solution containing Br⁻(aq) will produce Br₂(l).
(E) The first ionization energies increase as the atomic number increases.

57. Considering the structures of the three compounds, X, Y, and Z, shown above, the ranking of their solubility in water from least to greatest is which of the following?

(A) X < Y < Z
(B) X < Z < Y
(C) Z < Y < X
(D) Y < Z < X
(E) Y < X < Z

58. Of the following compounds, which is involved in the environmental problem known as acid rain?

(A) CO₂
(B) CF₂Cl₂
(C) SO₂
(D) H₂S
(E) SiO₂

59. When the chemical equation above is balanced in terms of lowest whole number coefficients, the coefficient for H₂O is

(A) 1
(B) 2
(C) 3
(D) 6
(E) 8

60. Which of the following best describes the role of a catalyst in a chemical reaction?

(A) The catalyst lowers the activation energy by changing the mechanism of the reaction.
(B) The catalyst increases the strength of the chemical bonds in the reactant molecules.
(C) The catalyst increases the value of the equilibrium constant.
(D) The catalyst provides kinetic energy to reactant molecules to increase the reaction rate.
(E) The catalyst bonds to the reaction products and drives the equilibrium toward the products.

61. On the basis of trends in the periodic table, an atom of which of the following elements is predicted to have the lowest first ionization energy?

(A) Ar
(B) Cl
(C) K
(D) Rb
(E) I
CHEMISTRY

\[ X(g) + Y(g) \rightleftharpoons Z(g) \]

62. Which of the following statements is true for the chemical system represented above when the system has reached a state of equilibrium at a constant temperature and pressure?

(A) The forward and reverse reactions have stopped.
(B) The forward and reverse reactions occur at the same rate.
(C) The rate of formation of \(Z(g)\) is equal to half the rate of consumption of \(X(g)\).
(D) Introducing a catalyst will result in an increased amount of \(Z(g)\) at equilibrium.
(E) Introducing more \(Y(g)\) to the system will cause more \(X(g)\) to form.

63. If a 1.0 \(M\) solution of HA, a weak acid, has a pH of 2.0, then the value of \(K_a\), the acid dissociation constant, for HA is closest to

(A) \(1.0 \times 10^{-4}\)
(B) \(1.4 \times 10^{-4}\)
(C) \(1.0 \times 10^{-2}\)
(D) \(1.4 \times 10^{-2}\)
(E) \(1.4 \times 10^{-1}\)

64. Which of the following elements is never found pure (i.e., chemically uncombined with one or more other elements) in Earth's crust?

(A) S
(B) K
(C) Cu
(D) Pt
(E) Au

66. Which of the following single covalent bonds is the most polar?

(A) B - F
(B) F - F
(C) Cl - F
(D) P - Br
(E) Si - Cl

67. In which of the following are the compounds listed correctly in order of increasing strength of their oxygen-to-oxygen bonds?

(A) \(O_2 < O_3 < H_2O_2\)
(B) \(O_2 < H_2O_2 < O_3\)
(C) \(O_3 < O_2 < H_2O_2\)
(D) \(H_2O_2 < O_3 < O_2\)
(E) \(H_2O_2 < O_2 < O_3\)

68. An atom of which of the following elements has the smallest radius?

(A) K
(B) Ca
(C) Br
(D) Rb
(E) Sr

69. Which of the following is a Brønsted-Lowry acid-base pair?

(A) \(H^+\) and \(Cl^-\)
(B) \(Na^+\) and \(Cl^-\)
(C) HCl and NaOH
(D) \(H_2SO_4\) and \(SO_4^{2-}\)
(E) \(HCO_3^-\) and \(CO_3^{2-}\)

70. A sample of gas has a volume of 1.0 L at 300 K and 2.0 atm. If the volume and the absolute temperature are both doubled, what is the final pressure of the sample?

(A) 0.50 atm
(B) 1.0 atm
(C) 2.0 atm
(D) 4.0 atm
(E) 8.0 atm
Study Resources

Most textbooks used in college-level chemistry courses cover the topics in the outline given earlier, but the approaches to certain topics and the emphases given to them may differ. To prepare for the Chemistry exam, it is advisable to study one or more college textbooks, which can be found in most college bookstores. When selecting a textbook, check the table of contents against the knowledge and skills required for this test.

Visit www.collegeboard.org/clepprep for additional chemistry resources. You can also find suggestions for exam preparation in Chapter IV of the Official Study Guide. In addition, many college faculty post their course materials on their schools' websites.

Answer Key

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>B</td>
</tr>
<tr>
<td>2</td>
<td>E</td>
</tr>
<tr>
<td>3</td>
<td>B</td>
</tr>
<tr>
<td>4</td>
<td>E</td>
</tr>
<tr>
<td>5</td>
<td>D</td>
</tr>
<tr>
<td>6</td>
<td>C</td>
</tr>
<tr>
<td>7</td>
<td>C</td>
</tr>
<tr>
<td>8</td>
<td>A</td>
</tr>
<tr>
<td>9</td>
<td>B</td>
</tr>
<tr>
<td>10</td>
<td>C</td>
</tr>
<tr>
<td>11</td>
<td>D</td>
</tr>
<tr>
<td>12</td>
<td>D</td>
</tr>
<tr>
<td>13</td>
<td>C</td>
</tr>
<tr>
<td>14</td>
<td>E</td>
</tr>
<tr>
<td>15</td>
<td>E</td>
</tr>
<tr>
<td>16</td>
<td>B</td>
</tr>
<tr>
<td>17</td>
<td>A</td>
</tr>
<tr>
<td>18</td>
<td>A</td>
</tr>
<tr>
<td>19</td>
<td>E</td>
</tr>
<tr>
<td>20</td>
<td>A</td>
</tr>
<tr>
<td>21</td>
<td>C</td>
</tr>
<tr>
<td>22</td>
<td>D</td>
</tr>
<tr>
<td>23</td>
<td>D</td>
</tr>
<tr>
<td>24</td>
<td>C</td>
</tr>
<tr>
<td>25</td>
<td>E</td>
</tr>
<tr>
<td>26</td>
<td>A</td>
</tr>
<tr>
<td>27</td>
<td>A</td>
</tr>
<tr>
<td>28</td>
<td>D</td>
</tr>
<tr>
<td>29</td>
<td>B</td>
</tr>
<tr>
<td>30</td>
<td>C</td>
</tr>
<tr>
<td>31</td>
<td>C</td>
</tr>
<tr>
<td>32</td>
<td>E</td>
</tr>
<tr>
<td>33</td>
<td>E</td>
</tr>
<tr>
<td>34</td>
<td>A</td>
</tr>
<tr>
<td>35</td>
<td>B</td>
</tr>
</tbody>
</table>