

## MULTIPLE CHOICE ANSWERS AND EXPLANATIONS FOR PRACTICE TEST 1

1. C. Ethanol, a covalent substance, does not ionize in water, so the pH is approximately 7.
2. D. Acids react with carbonates to form carbon dioxide gas and water:  
$$\text{H}^+(aq) + \text{HCO}_3^-(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l)$$
3. A. Generally, anions are basic except anions of strong acids. Phosphate is a base:  
$$\text{PO}_4^{3-}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{HPO}_4^{2-}(aq) + \text{OH}^-(aq)$$
4. A. Phosphate ion reacts with acid to form hydrogen phosphate, a base.  
$$\text{PO}_4^{3-}(aq) + \text{H}^+(aq) \rightarrow \text{HPO}_4^{2-}(aq)$$

(Note: Hydrogen phosphate ion is amphoteric, and without more information, it might be an acid in aqueous solution. However, sodium phosphate is the only substance listed that has a possibility to be basic when acid is added.)
5. E. Ammonium ion reacts with hydroxide ion to form ammonia, a water-soluble gas and water:  
$$\text{NH}_4^+(aq) + \text{OH}^-(aq) \rightarrow \text{NH}_3(aq) + \text{H}_2\text{O}(l)$$
6. D. Generally, first ionization energy increases from left to right along a period and from bottom to top within a group.
7. B. Group 2 atoms have relatively low first and second ionization energies because their two valence electrons are well screened from the nuclear charge. However, their third ionization energies are relatively high because the third electron comes from a poorly screened inner noble gas core and experiences a high-effective nuclear charge.
8. A. Generally, electron affinity increases from left to right along a period and from bottom to top within a group. However, chlorine has a higher electron affinity than fluorine because fluorine's small size contributes to considerable electron-electron repulsion.
9. D. The noble gases are the smallest atoms in their respective periods because their valence electrons experience the highest effective nuclear charges.
10. A. Chlorine displays oxidation states of  $1-$ ,  $0$ ,  $1+$ ,  $3+$ ,  $5+$ , and  $7+$ .
11. C. The change in solubility of NaCl with temperature is minimal compared to the others, indicating that solid NaCl dissolves in water with little or no heat change.
12. A. As temperature increases, the solubility of cerium sulfate decreases, indicating that cerium sulfate dissolves exothermically. Adding heat to an exothermic process will cause the reaction to shift toward reactants, in this case decreasing solubility.

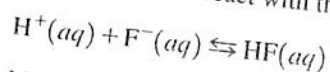
13. E. The two driving forces that influence chemical and physical change are a decrease in enthalpy and an increase in entropy. Most salts dissolve endothermically, but the accompanying increase in entropy (randomness of the system) overcomes the increase in enthalpy associated with an endothermic reaction. All salts that dissolve endothermically fit into this category.
14. D. Carbon dioxide is a nonpolar molecule, held together only by London dispersion forces. Sublimation is the process where a solid changes directly into a gas. Energy is required to overcome the forces of attraction that hold the molecules together in the solid phase.
15. D. The liquid will continue to evaporate without stopping. The pressure will rise until the partial pressure of the vapor equals the vapor pressure of the liquid at the given temperature. At that point, a dynamic equilibrium is established in which the rate of evaporation equals the rate of condensation.
16. A. The  $\text{C}=\text{O}$  bond in trigonal planar acetone is very polar, making the molecule a strong dipole. The absence of  $\text{H}-\text{O}$  bonds rules out hydrogen bonding, but strong dipole-dipole interactions exist.
17. B. The balanced equation is  $2\text{Br}_2\text{O}_7(g) \rightarrow 2\text{Br}_2(g) + 7\text{O}_2(g)$ .
18. D. The only products obtained from the complete combustion are carbon dioxide and water. The assumptions are that all the carbon in the sample is converted to carbon dioxide and all the hydrogen in the sample is converted to water. The amounts of carbon and hydrogen can be determined by measuring the masses of carbon dioxide and water produced.
19. E. The molar mass of anhydrous  $\text{CuSO}_4$  is 159.5 g/mol. 15.95 g is 0.100 mol  $\text{CuSO}_4$ . 9.00 g of water is 0.5 mol water or 5 times as many moles of water as copper(II) sulfate. The formula of the hydrate is  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ .
20. C. The molecular formula of a compound is an integer multiple of the empirical formula. The molecular formula,  $\text{C}_6\text{H}_{12}\text{O}_3$ , is 3 times the given empirical formula and has a molar mass of 132 g/mol.
21. C. Atom Y loses three valence electrons to form a  $\text{Y}^{3+}$  ion. Atom Z gains two electrons to form a  $\text{Z}^{2-}$  ion. The two ions combine to form  $\text{Y}_2\text{Z}_3$ .
22. D. The reaction in question is the reverse of the given reaction and balanced with half the coefficients. The change in enthalpy of the reaction in question is half the value of the given reaction with a change in sign from negative to positive.
23. E. Ideal gases are assumed to be composed of tiny particles having no attractive forces. Therefore, the chemical composition will have no effect on their behavior.
24. B. Temperature is a measure of average kinetic energy. Two gases at the same temperature will have the same kinetic energy.

25. B. Nonideal gases are real gases with significant attractive forces between particles and significant particle volumes. Low temperatures slow the particles sufficiently so that the attractive forces become more dominant. High pressures crowd the particles so that their volumes become significant.
26. C. The balanced equation is  $\text{O}_2(g) + 2\text{SO}_2(g) \rightarrow 2\text{SO}_3(g)$ . Oxygen is the limiting reactant, so twice as many moles of  $\text{SO}_3$  can be formed from the available moles of oxygen. At the same conditions of temperature and pressure, liters are proportional to moles.
- $$x \text{ L SO}_3 = 1 \text{ L O}_2 (2 \text{ L SO}_3 / 1 \text{ L O}_2) = 2 \text{ L}$$
27. C. Ideal gases have insignificant volumes and no attractive forces. Deviations from ideal behavior become greater for larger molecules with high attractive forces. Water is a polar molecule with very large dipole-dipole interactions and hydrogen bonds.
28. B. Effective nuclear charge increases from left to right along any period. Neither electron is screened from the charge of the two protons on He, making helium's effective nuclear charge double that of hydrogen. Consequently, helium is smaller than hydrogen.
29. B. Generally, first ionization energy increases from left to right along any period because the effective nuclear charge increases. However, gallium's  $3p$  electron is screened by the  $4s$  level and the  $3d$  level, giving gallium a slightly smaller effective nuclear charge than calcium.
30. C. Generally, electronegativity increases from the lower left to the upper right of the periodic table. Elements that are farthest apart along the diagonal from lower left to upper right probably have the largest difference in electronegativity. Li and Br also have a large difference, but Rb has a lower electronegativity than does Li and chlorine's electronegativity is larger than that of bromine.
31. A. Se has ten electrons around it, four bonding pairs, and two nonbonding pairs. Octet structures can be written for all the others.
32. A. Lewis structures show that the nitrogen molecule has a triple bond, the oxygen molecule has a double bond, and all the rest have single bonds. Generally, triple bonds are stronger and shorter than double bonds, and double bonds are stronger and shorter than single bonds.
33. B. Carbonate is a planar ion, which is associated with  $sp^2$  hybridization.
34. E. Carbon dioxide has two polar bonds but its linear geometry causes the dipoles to cancel. All the others have polar bonds and nonlinear geometry.
35. B. A linear plot of  $\ln[A]$  versus time is characteristic of a simple first-order process. A simple second-order process would yield a straight line for  $1/[A]$  versus time.



36. E. If just  $[Y]$  is halved, then the initial rate would decrease by a factor of  $(1/2)^2 = 1/4$ . To make the rate remain constant,  $[X]$  is quadrupled because  $(4)^1 = 4$ .
37. B. An endothermic reaction has a higher activation energy for the forward reaction than for the reverse reaction and the rate of reaction decreases with increasing activation energy. The enthalpy of products for an endothermic reaction is higher than the enthalpy of reactants.
38. D. If 97% is decayed, 3% of the original sample remains. Three percent represents the amount left after about five half-lives:  $100\% \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} = 3.1\%$ . Five half-lives  $= 5 \times 30 \text{ years} = 150 \text{ years}$ .
39. A. The balanced equation is
- $$\text{H}_2\text{O} + \text{MnO}_4^- + 3\text{NO}_2 \rightarrow \text{MnO}_2 + 3\text{NO}_3^- + 2\text{H}^+$$
40. C. An easily reduced substance has a large positive standard reduction potential,  $E^\circ$ , and an easily oxidized substance has the highest oxidation potential. Only reduction potentials are shown in the table and fluorine has the largest standard reduction potential, so it is most easily reduced. Aluminum is easily oxidized because the reverse reaction of the ones shown in the table has the highest oxidation potential at  $+1.66 \text{ V}$ .
41. B. Reduction of copper(II) ions is more favorable than the reduction of aluminum ions because the voltage of the copper half-reaction is more positive. Therefore, the aluminum reaction will be the oxidation and the sign of its voltage will change from  $-$  to  $+$ .
- $$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s}) \quad E^\circ_{\text{red}} = 0.337 \text{ V}$$
- $$\text{Al}(\text{s}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{e}^- \quad E^\circ_{\text{ox}} = +1.66 \text{ V}$$
- $$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}} = 0.337 + 1.66 = 2.00 \text{ V}$$
42. C. Generally, protonated anions and water are amphoteric. Amphoteric substances can both donate and accept a proton. Protonated anions have protons to donate and negative charges that can attract protons.  $\text{NH}_4^+$  can donate a proton, but because like charges repel, it is unlikely to accept a positively charged proton.
43. B. Phosphoric acid is a weak polyprotic acid, which ionizes only slightly. While all five of the listed species exist in solution,  $\text{H}_3\text{PO}_4$  exists to the greatest extent.
44. D. KOH is a strong base, which dissociates completely in aqueous solution. HCl is a strong acid,  $\text{CH}_3\text{COOH}$  is a weak acid and weak electrolyte,  $\text{NH}_4\text{NO}_3$  is a weak acid and strong electrolyte, and  $\text{CH}_3\text{OH}$  is a neutral nonelectrolyte.
45. A. In general, for similar formulas having the same central atom, acid strength increases with increasing oxidation number of the central atom. The oxidation number of chlorine in HClO is  $+1$ , the lowest of the compounds listed.

46. A. Fluoride ion is the conjugate base of the weak acid, HF. The solubility of calcium fluoride will increase as pH decreases, because at low pH the available protons will react with the basic fluoride ion.



Methanol is completely miscible in water and all the other choices are ionic compounds containing only neutral cations and anions.

47. D. The weakest acid has the strongest conjugate base ( $K_w = K_a \times K_b$ ). At the equivalence point, all that is left is weak base.
48. B. All are bases except  $\text{NH}_4^+$ . It is unlikely that the already positively charged ammonium ion will accept a proton.

49. B. Removing  $\text{PCl}_5(g)$  will slow the reverse reaction, causing the forward reaction to be faster than the reverse reaction. The equilibrium will shift right and consume  $\text{PCl}_3$ . Adding a catalyst or an inert gas will not affect the position of the equilibrium. The other responses will cause the amount of  $\text{PCl}_3$  to increase.

50. C. The equilibrium-constant expression equals the ratio of the concentrations of products to reactants, each raised to the power of the coefficient that balances the equation.

$$51. A. K_c = [\text{NH}_3]^2 / [\text{N}_2][\text{H}_2]^3 = (2.0)^2 / (2.0)(2.0)^3 = 0.25$$

(Remember that no calculators are allowed on the multiple choice section, so the arithmetic required for complex quantitative problems is relatively simple.)

52. D. Temperature is the only parameter that will change the value of the equilibrium constant. Decreasing the temperature will favor the exothermic or, in this case, the reverse reaction. The value of the equilibrium constant will decrease because the reaction shifts left, reducing the amount of Z and reducing  $K_c$ .  $K_c = [\text{Z}]^2 / [\text{X}][\text{Y}]^2$ .

53. C. An ICE table would look like this:

	$\text{H}_2(g)$	$+$	$\text{I}_2(g)$	$\rightarrow$	$2\text{HI}(g)$
I	1.0 M		1.0 M		0
C	$-x$		$-x$		$+2x$
E	$1.0 - x$		$1.0 - x$		$2x$

$$K_c = 36 = (2x)^2 / (1.0 - x)(1.0 - x)$$

If we take the square root of both sides of the equation, we obtain:

$$6 = 2x / (1.0 - x) \text{ or } 6 - 6x = 2x, x = 0.750 = [\text{H}_2]$$

54. D. Generally, electronegativity increases from left to right along a period. The noble gases do not attract electron pairs and are not included in the trend.
55. A)  $\text{AgCl}(s) = \text{Ag}^+(aq) + \text{Cl}^-(aq)$

$$K_{sp} = 1.8 \times 10^{-10} = [\text{Ag}^+][\text{Cl}^-]$$

Let  $x$  = molar solubility of AgCl.

At equilibrium,  $[\text{Ag}^+] = x$ ,  $[\text{Cl}^-] = 0.050 + x \approx 0.050 \text{ M}$

$$1.8 \times 10^{-10} = x(0.050)$$

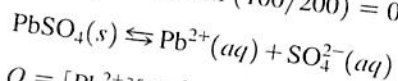
$$x = (1.8/0.050) \times 10^{-10}$$

56. C. The concentrations of the ions are not sufficient to form calcium sulfate but are sufficient to form lead(II) sulfate.

$$[\text{Pb}^{2+}] = 0.0020 \text{ M} (100/200) = 0.0010 \text{ M}$$

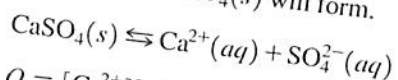
$$[\text{Ca}^{2+}] = 0.0020 \text{ M} (100/200) = 0.0010 \text{ M}$$

$$[\text{SO}_4^{2-}] = 0.002 \text{ M} (100/200) = 0.0010 \text{ M}$$



$$Q = [\text{Pb}^{2+}][\text{SO}_4^{2-}] = (0.0010)(0.0010) = 1 \times 10^{-6}$$

$K_{sp} < Q$ , so  $\text{PbSO}_4(s)$  will form.



$$Q = [\text{Ca}^{2+}][\text{SO}_4^{2-}] = (0.0010)(0.0010) = 1 \times 10^{-6}$$

$K_{sp} > Q$ , so  $\text{CaSO}_4(s)$  will not form.

57. A. Oxide ion has just eight protons in its nucleus, two of which are screened by its two  $1s$  electrons. Its effective nuclear charge is the smallest of the species listed.
58. C. A valid octet structure can be written for  $\text{N}_2\text{O}$ .
59. B. As a reaction continues, its rate decreases as the concentrations of reactants decrease. The half-life remains the same for any first-order reaction.
60. C. Experiments 1 and 2 show that when  $[\text{A}]$  is constant and  $[\text{B}]$  doubles, the rate doubles, meaning the exponent of  $[\text{B}]$  is 1. Experiments 2 and 3 show that when both  $[\text{A}]$  and  $[\text{B}]$  double, the rate increases by a factor of 8. Doubling  $[\text{B}]$  accounts for a factor of 2, so doubling  $[\text{A}]$  accounts for a factor of 4. Therefore, the exponent of  $[\text{A}]$  is 2.