Practice Test 2

These questions are representative of the AP Chemistry examination, but keep in mind that it is impossible to predict exactly how well you will do on the actual exam. The first section of this test is 50% of your total test grade. Time yourself so that you finish this part in 90 minutes.

AP CHEMISTRY EXAMINATION
Section I: Multiple-Choice Questions
Time: 90 minutes
Number of Questions: 60

No calculators can be used in this section. A periodic table and a formula chart with constants is provided.

Directions: Each of the questions or incomplete statements below is followed by four suggested answers or completions. Select the one that is best in each case.

1. The element that has atoms with the largest atomic radius is
   (A) He
   (B) Na
   (C) Si
   (D) Mg

2. The element that has the highest first ionization energy is
   (A) Na
   (B) Si
   (C) Mg
   (D) Ar

3. The element that is the most metallic in character is
   (A) He
   (B) C
   (C) Si
   (D) Mg

4. The element that most easily forms an ion with a 2+ charge is
   (A) Na
   (B) Si
   (C) Mg
   (D) Ar

5. If two atoms have different atomic numbers but the same mass number, what must be true?
   (A) They must be atoms of the same element.
   (B) Each must contain the same total number of neutrons and protons.
   (C) The number of neutrons in both must be the same.
   (D) The number of protons in each atom must be the same.

6. An element has two naturally occurring isotopes. One isotope has an abundance of 80% and a mass of 122.0 amu. The other has a mass of 120.0 amu. What is the atomic mass of the element?
   (A) 121.6
   (B) 122.0
   (C) 120.6
   (D) 120.8

Periodic Table of Elements

1 H 1.008																	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226	89 Ac† (227)	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg (272)							
			*Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
			†Actinides	90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

7. A monoatomic ion contains 15 protons, 16 neutrons, and 18 electrons. What ion of what isotope is this?
   (A) $^{31}P^{3+}$
   (B) $^{34}Se^{2-}$
   (C) $^{31}P^{3-}$
   (D) $^{34}As^{3-}$

8. Which one of the following is an example of a physical property whose value is not dependent on the amount of substance?
   (A) Aluminum burns in bromine.
   (B) A balloon of hydrogen and oxygen explodes.
   (C) A bottle of rubbing alcohol has a mass of 900 grams.
   (D) The molar mass of sodium is 22.99 g/mol.

9. Of the colors listed below that are part of the visible portion of the electromagnetic spectrum, which one has the highest frequency?
   (A) blue
   (B) green
   (C) yellow
   (D) red

10. Atomic size increases going from top to bottom in a group of the periodic table because
    (A) the number of protons increases
    (B) the number of electron shells increases
    (C) the atomic mass increases
    (D) the atomic volume increases

Questions 11–15 refer to the structure and geometry of molecules:

11. What is the arrangement of electron pairs around the central atom in the molecule krypton difluoride, KrF_2?
    (A) linear
    (B) trigonal planar
    (C) tetrahedral
    (D) trigonal bipyramidal

12. Which of the following require(s) resonance to reconcile the Lewis electron representations with the actual or real structure?
    CO_2 O_3 SO_4^{2-} NO_3^-
    (A) CO_2 only
    (B) O_3 only
    (C) O_3 and NO_3^- only
    (D) O_3 and SO_4^{2-} only

13. What is the condensed (shorthand) electron configuration for the phosphide ion, P^{3-}?
    (A) $[Ne]3s^2$
    (B) [Ar]
    (C) $[Ar]4s^2$
    (D) $[Ne]3s^23p^5$

14. What is the average bond order in the nitrate ion, NO_3^-?
    (A) 1.0
    (B) 1.33
    (C) 1.5
    (D) 2.0

15. In the Lewis dot structure for the molecule S_2O, how many lone pairs of electrons are there around the central atom?
    (A) 4
    (B) 3
    (C) 2
    (D) 1

Questions 16–19 refer to the system of reactions below:

$2\ NO(g) \rightarrow N_2O_2(g)$ *Fast*

$N_2O_2(g) + H_2(g) \rightarrow N_2O(g) + H_2O(g)$ *Slow*

$N_2O(g) + H_2(g) \rightarrow N_2(g) + H_2O(g)$ *Fast*

16. What is (are) reaction intermediate(s) in this reaction mechanism?
    (A) NO and N_2O
    (B) N_2O and N_2O_2
    (C) N_2O_2
    (D) N_2O and H_2

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17. What is the rate law for the reaction according to this mechanism?
(A) Rate = $k[N_2O_2][H_2]$
(B) Rate = $k[NO]^2$
(C) Rate = $k[N_2O_2]^2[H_2]^2$
(D) Rate = $k[NO]^2[H_2]$

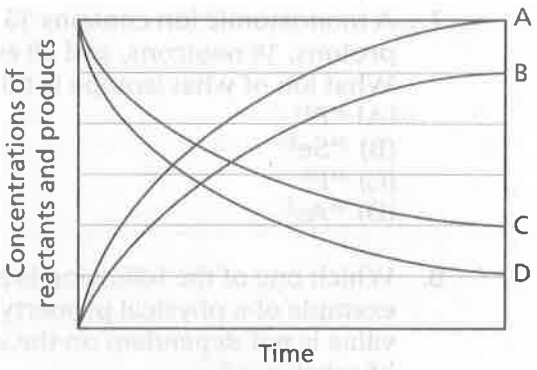
18. Adding a catalyst to this reaction
(A) increases the amount of product at equilibrium
(B) changes the route the reaction takes between reactants and products
(C) increases the activation energy required for the reaction
(D) shifts the equilibrium toward the product side

19. The overall equation for the reaction is
(A) $N_2O(g) + H_2(g) \rightarrow N_2(g) + H_2O(g)$
(B) $2\ NO(g) + 2\ H_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$
(C) $2\ NO(g) + H_2(g) \rightarrow N_2(g) + H_2O(g)$
(D) $NO(g) + H_2(g) \rightarrow N(g) + H_2O(g)$

Questions 20–22 refer to the reaction of ammonia with nitrogen monoxide, which is described by the following equation:

$$4\ NH_3(g) + 6\ NO(g) \rightarrow 5\ N_2(g) + 6\ H_2O(g)$$

20. In one particular experiment, equal molar amounts of the reactants were mixed and concentrations of the reactants and products were plotted against time. The graph below was obtained. Identify the components A, B, C, and D, and choose the line where all the components are identified properly.



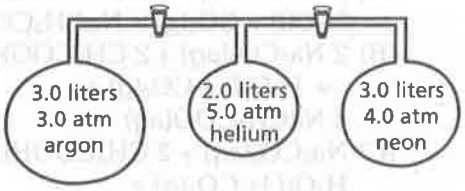
	A	B	C	D
(A)	NH_3	NO	N_2	H_2O
(B)	H_2O	NO	N_2	NH_3
(C)	NH_3	N_2	NO	H_2O
(D)	H_2O	N_2	NH_3	NO

21. At one stage in the reaction, nitrogen is produced at a rate of 30 mol $L^{-1}\ s^{-1}$. At what rate is nitrogen monoxide used up at this stage?
(A) 36 mol $L^{-1}\ s^{-1}$
(B) 30 mol $L^{-1}\ s^{-1}$
(C) 24 mol $L^{-1}\ s^{-1}$
(D) 6.0 mol $L^{-1}\ s^{-1}$

22. A proposed rate law for this reaction is: Rate of reaction = $k[NO]^2[NH_3]$ where k is the rate constant for the reaction. What happens to the rate of the reaction if the concentration of NO is doubled and the concentration of NH_3 is halved? (The temperature remains unchanged.)
(A) The rate halves.
(B) The rate stays the same.
(C) The rate doubles.
(D) The rate increases by four.

Questions 23–27 refer to the system of gases below:

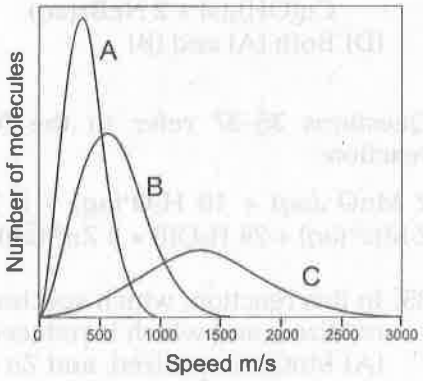
23.



With the stopcocks closed, the three bulbs in the apparatus shown above are filled with the quantities indicated of argon, helium, and neon, respectively. Assume the gases are ideal and that the volumes include the volumes of the connecting tubes. What is the partial pressure of helium inside the apparatus when the taps are opened?

(A) 1.00 atm
(B) 1.25 atm
(C) 1.50 atm
(D) 2.00 atm

24. The graph below shows the speed distributions for the three gases, Ar, He, and Ne, at the same temperature.



Match the gas with the distribution.

	A	B	C
(A)	Ar	Ne	He
(B)	Ar	He	Ne
(C)	Ne	He	Ar
(D)	Ar	He	Ne

25. Which of the gases exhibits the greatest deviation from ideal behavior?
(A) He, because it is the smallest atom.
(B) Ar, because it has the most electrons and is therefore the most polarizable.
(C) Ar, because it is the biggest atom.
(D) None, because inert gases are all ideal.

26. The pressure of the system is increased by decreasing the system volume and the temperature is decreased. How will the system conditions as predicted by the ideal gas law be affected?
(A) There will be no effect as long as all of the substances are in the gas phase.
(B) Intermolecular attractions between gas molecules will become more significant.
(C) The volume of the gas molecules will become significantly large relative to the volume of the container.
(D) Both (B) and (C) are true.

27. The temperature of the system is increased from 0°C to 273°C, and half the moles of gas are released. What happens to the pressure?
(A) It is unchanged.
(B) It doubles.
(C) It decreases by one half.
(D) It increases four times.

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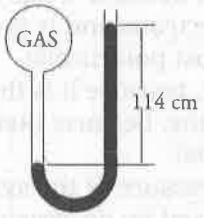
28. A U-tube mercury manometer is open to the atmosphere (1 atm = 760 mm Hg) on the right arm and is connected to a glass vessel containing a gas on the left arm as shown.



The level of the mercury in the tube is 114 cm (1140 mm) higher in the right arm. What is the pressure of the gas in the glass vessel?
(A) 0.50 atm
(B) 1.0 atm
(C) 1.5 atm
(D) 2.5 atm

29. 0.24 mole of sodium carbonate is dissolved in sufficient water to make 300.0-mL of solution. What is the molarity of the sodium carbonate solution?
(A) 0.24 *M*
(B) 0.72 *M*
(C) 0.32 *M*
(D) 0.80 *M*

30. The concentration of sodium ions in a 300.0-mL solution made from 0.24 mole of sodium carbonate dissolved in water is
(A) 3.6 *M*
(B) 1.6 *M*
(C) 1.2 *M*
(D) 4.8 *M*

31. Distilled vinegar purchased at the grocery store is a 0.833 *M* solution of acetic acid. What is the concentration in mole L^{-1} of acetic acid in a solution prepared by taking 100. mL of vinegar and adding it to enough water to make 2.00 L of solution?
(A) 0.0417 *M*
(B) 0.0833 *M*
(C) 0.0975 *M*
(D) 0.167 *M*

32. The correct balanced equation for the reaction of sodium carbonate solution with acetic acid is:
(A) $NaHCO_3(aq) + CH_3COOH(aq) \rightarrow H_2O(l) + CO_2(g) + NaCH_3COO(aq)$
(B) $2\ Na_2CO_3(aq) + 2\ CH_3COOH(aq) \rightarrow H_2O(l) + CO_2(g) + 4\ NaCH_3COO(aq)$
(C) $Na_2CO_3(aq) + 2\ CH_3COOH(aq) \rightarrow H_2O(l) + CO_2(g) + 2\ NaCH_3COO(aq)$
(D) $Na_2CO_3(aq) + CH_3COOH(aq) \rightarrow H_2O(l) + CO_2(g) + 2\ NaCH_3COO(aq)$

33. The reaction of sodium carbonate and hydrochloric acid is an example of
(A) oxidation-reduction
(B) synthesis
(C) acid-base
(D) combustion

34. Which of the following is an oxidation–reduction reaction?
(A) $Ca(OH)_2(aq) + 2\ HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$
(B) $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$
(C) $CuBr_2(aq) + 2\ NaOH(aq) \rightarrow Cu(OH)_2(s) + 2\ NaBr(aq)$
(D) Both (A) and (B)

Questions 35–37 refer to the following reaction:

$2\ MnO^-_4(aq) + 16\ H_3O^+(aq) + 5\ Zn(s) \rightarrow 2\ Mn^{2+}(aq) + 24\ H_2O(l) + 5\ Zn^{2+}(aq)$

35. In this reaction, which species is oxidized, and which is reduced?
(A) MnO_4^- is oxidized, and Zn is reduced.
(B) Zn is oxidized, and MnO_4^- is reduced.
(C) H_3O^+ is oxidized, and Zn is reduced.
(D) Zn^{2+} is oxidized, and Mn^{2+} is reduced.

36. How many electrons are transferred in this reaction?
(A) 5
(B) 7
(C) 10
(D) 14

37. This reaction is typically performed as a titration, but no indicator is added. How is the endpoint identified?
(A) When all the Zn has reacted, the solution will stay purple, the color of the permanganate ion.
(B) A precipitate forms.
(C) When all the Zn has reacted, additional permanganate reacts with water to form a gas.
(D) The pH of the solution increases to 14.

38. What is the change in oxidation number of sulfur in the following half-reaction?
$S_4O_6^{2-} + 10\ H_2O \rightarrow SO_4^{2-} + 2\ OH^- + 14\ e^-$
(A) –2 to –8
(B) +4 to +1
(C) +4 to +6
(D) +2 ½ to +6

39. Balance the following equation for the reaction of dimethylhydrazine with dinitrogen tetroxide:

$\_\_C_2H_8N_2 + \_\_N_2O_4 \rightarrow \_\_N_2 + \_\_CO_2 + \_\_H_2O$

The sum of all the coefficients in the balanced equation is
(A) 9
(B) 11
(C) 12
(D) 14

40. In which compound does nitrogen have the highest (most positive) oxidation number?
(A) NO_2
(B) NH_3
(C) N_2
(D) NO_3^-

Questions 41–43 refer to the following reaction:

$2\ CH_3OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(l)$

41. When this reaction is performed in a calorimeter, the water temperature increases. Therefore, the heat of reaction
(A) is negative
(B) is positive
(C) is equal to zero
(D) cannot be determined from the data

42. From the equation of the reaction, what is most likely happening to entropy during this process?
(A) Entropy of the system decreases, entropy of the surroundings increases
(B) Entropy of the system increases, entropy of the surroundings increases
(C) Entropy of the system decreases, entropy of the surroundings decreases
(D) Entropy of the system increases, entropy of the surroundings decreases

43. Given the following bond energies (kJ/mol), what is the ΔH_{rxn} in kJ for the reaction?

C–H	413	C–O	358	O=O	495
C=O	799	O–H	467		

(A) 2559 kJ
(B) –2559 kJ
(C) –1319 kJ
(D) 1319 kJ

44. For which substance does the standard enthalpy of formation ΔH_f° equal zero?
(A) $H_2O(l)$
(B) $Cu(s)$
(C) $O_3(g)$
(D) $Fe(l)$

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45. Which process is always exothermic?
(A) evaporation of a liquid
(B) dissolving a typical salt in water
(C) breaking a hydrogen molecule into atoms
(D) freezing water

Questions 46–47 refer to the thermodynamics of dissolving ammonium nitrate in water.

46. When crystalline ammonium nitrate dissolves in water to make a solution, the solution gets very cold, dropping in temperature about 20°C. What are the signs of ΔH, ΔS, and ΔG for this process?

	ΔH	ΔS	ΔG
(A)	(–)	(+)	(+)
(B)	(–)	(+)	(–)
(C)	(+)	(–)	(+)
(D)	(+)	(+)	(–)

47. Under what conditions is the process spontaneous?
(A) The process is spontaneous at all temperatures.
(B) The process is not spontaneous at any temperature.
(C) The process is spontaneous at high temperatures and not at low temperatures.
(D) The process is spontaneous at low temperatures and not at high temperatures.

48. Arrange the following substances in order of increasing entropy.
$H_2O(l)$ $CO_2(s)$ $N_2(l)$ $N_2(g)$
(A) $CO_2(s) < H_2O(l) < N_2(l) < N_2(g)$
(B) $N_2(g) < CO_2(s) < N_2(l) < H_2O(l)$
(C) $H_2O(l) < CO_2(s) < N_2(l) < N_2(g)$
(D) $CO_2(s) < N_2(l) < H_2O(l) < N_2(g)$

Questions 49–50 refer to the following reaction:

$$3\ F_2(g) + Cl_2(g) \rightleftharpoons 2\ ClF_3(g)$$

49. When the system is at equilibrium at 298 *K*, $[F_2] = 2.0\ M$; $[Cl_2] = 2.5\ M$; $[ClF_3] = 3.0\ M$. Calculate the value of *K*, the equilibrium constant.
(A) 0.45
(B) 0.90
(C) 1.2
(D) 0.25

50. The system volume is decreased at constant temperature. In which direction, if any, will the system shift to restore equilibrium?
(A) There will be no change.
(B) It will shift to the left.
(C) It will shift to the right.
(D) The value of *K* will change because the volume has decreased.

Questions 51–52 refer to a solution prepared by dissolving hydrogen sulfide in water to form hydrosulfuric acid, $H_2S(aq)$:

51. The Brønsted–Lowry conjugate base of $H_2S(aq)$ is
(A) S^{2-}
(B) HS^-
(C) OH^-
(D) H_3O^+

52. The K_a of $H_2S(aq)$ is equal to 1.0×10^{-7} at 25°C. What is the K_b for its conjugate base?
(A) 1.0×10^{-7}
(B) 1.0×10^{-14}
(C) 1.8×10^{-5}
(D) $1.0 \times 10^{+7}$

53. Which one of the following equimolar solutions will act as a buffer solution?
(A) H_2SO_4 and Na_2SO_4
(B) H_2SO_4 and H_2SO_3
(C) KOH and KCN
(D) HI and KI

54. What is the molar solubility of cadmium sulfide, CdS? (K_{sp} of CdS = 3.6×10^{-29})
(A) 3.6×10^{-16} *M*
(B) 6.0×10^{-15} *M*
(C) 1.8×10^{-16} *M*
(D) 6.0×10^{-14} *M*

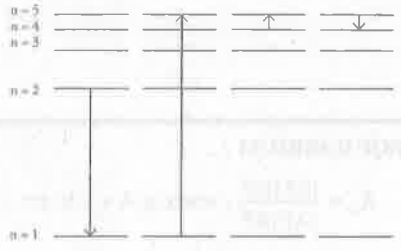
55. A substance with strong intermolecular forces of attraction would be expected to have
(A) a low boiling point
(B) a low vapor pressure
(C) a low melting point
(D) a low boiling point and a low vapor pressure

56. The molecules HF, $CaCO_3$, and $BaSO_4$ are weak electrolytes in aqueous solution. This must be because
(A) the molecules dissolve without breaking up into cations and anions
(B) water is polar and the molecules are nonpolar
(C) the atoms in the molecules are covalently bonded to each other
(D) the forces of attraction between the ions are stronger than those between the ions and water

57. How can the solubility of silver carbonate be increased?
(A) Add silver nitrate, $AgNO_3$
(B) Add sodium carbonate, Na_2CO_3
(C) Add calcium nitrate, $Ca(NO_3)_2$
(D) None of these will increase its solubility.

58. A solution with a low pH
(A) has a low concentration of acid
(B) has a high concentration of hydronium ions
(C) has a low concentration of hydronium ions
(D) is a basic solution

59. Below are representations of energy levels (drawn to scale) in the Bohr model of the hydrogen atom. Which electron making a transition, denoted by the arrows, results in the emission of a photon of shortest wavelength?



(A) (B) (C) (D)

60. Calculate the standard cell potential for the cell

$Sr(s)|Sr^{2+}(aq) \parallel Sn^{2+}(aq)|Sn(s)$ given the half-cell standard potentials:

$E°\ Sr^{2+}(aq)|Sr(s) = -2.89$ V

$E°\ Sn^{2+}(aq)|Sn(s) = -0.14$ V

(A) –3.03 *V*
(B) +3.03 *V*
(C) –2.75 *V*
(D) +2.75 *V*

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Advanced Placement Chemistry Equations and Constants

Throughout the test the following symbols have the definitions specified unless otherwise noted.

L, mL	= liter(s), milliliter(s)	mm Hg	= millimeters of mercury
g	= gram(s)	J, kJ	= joule(s), kilojoule(s)
nm	= nanometer(s)	V	= volt(s)
atm	= atmosphere(s)	mol	= mole(s)

ATOMIC STRUCTURE

$E = h\nu$

$c = \lambda\nu$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s

Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}

Avogadro's number = 6.022×10^{23} mol^{-1}

Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$K_c = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$, where $a\,A + b\,B \rightleftarrows c\,C + d\,D$

$K_p = \dfrac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$

$K_a = \dfrac{[H^+][A^-]}{[HA]}$

$K_b = \dfrac{[OH^-][HB^+]}{[B]}$

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at 25°C

$= K_a \times K_b$

$pH = -\log[H^+]$, $pOH = -\log[OH^-]$

$14 = pH + pOH$

$pH = pK_a + \log\dfrac{[A^-]}{[HA]}$

$pK_a = -\log K_a$, $pK_b = -\log K_b$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$\ln[A]_t - \ln[A]_0 = -kt$

$\dfrac{1}{[A]_t} - \dfrac{1}{[A]_0} = kt$

$t_{1/2} = \dfrac{0.693}{k}$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{M}$$

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

$$D = \frac{m}{V}$$

$$KE \text{ per molecule} = \frac{1}{2}mv^2$$

$$\text{Molarity, } M = \text{moles of solute per liter of solution}$$

$$A = abc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
M = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
a = molar absorptivity
b = path length
c = concentration

Gas constant, $R = 8.314 \text{ J mol}^{-1} \text{K}^{-1}$
$= 0.08206 \text{ L atm mol}^{-1} \text{K}^{-1}$
$= 62.36 \text{ L torr mol}^{-1} \text{K}^{-1}$
$1 \text{ atm} = 760 \text{ mm Hg}$
$= 760 \text{ torr}$
STP = 0.00°C and 1.000 atm

THERMOCHEMISTRY/ ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$
$$= -RT \ln K$$
$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)

Faraday's constant, F = 96,485 coulombs per mole of electrons
$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$

GO ON TO NEXT PAGE

Introduction to Section II: Free-Response Questions

Section II of the AP Chemistry Examination counts for 50% of the total test grade and involves several parts. Answering these questions gives you an opportunity to demonstrate your ability to present your material in clear, orderly, and convincing language. Your answers will be graded on the basis of accuracy, the kinds of information you include to support your responses, and the importance of the descriptive material used. Be specific; general, all-encompassing answers will not be graded as well as detailed answers with examples and equations. CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures. On the AP exam, be sure to write all your answers to the questions on the lined pages following each question in the test booklet. Do not write your answers in the white space between questions.

Section II: Free-Response Questions
Time: 90 minutes
Number of Questions: 7

Allow yourself no more than 90 minutes to answer these questions. You may use a calculator, the equations sheet, and the periodic table throughout this section. All questions must be answered.

1. For the reaction $PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$, $K_p = 0.0870$ at 300 °C. A flask is charged with 0.30 atm PCl_3, 0.60 atm Cl_2, and 0.10 atm PCl_5 at this temperature.
   (a) Determine whether the above conditions (the added gases) are at equilibrium. If not, determine to which direction the reaction must proceed to reach equilibrium.
   (b) Calculate the equilibrium partial pressures of the gases.
   (c) Predict and justify what effect increasing the volume of the system will have on the mole fraction of PCl_5 in the mixture.
   (d) The reaction as written is exothermic. Predict and justify what effect decreasing the temperature of the system will have on the mole fraction of PCl_5 in the mixture.

2. A 0.495 *M* solution of nitrous acid, HNO_2, has a pH of 1.83.
   (a) Find the $[H^+]$ and the percent ionization of nitrous acid in this solution.
   (b) Write the equilibrium expression and calculate the value of K_a for nitrous acid.
   (c) Sketch the pH curve that results when 20.0 mL of a 0.0125 *M* nitrous acid solution are titrated with 0.0125 *M* NaOH solution. Label the axes, the equivalence point, and the buffer region clearly.

3. Given the following data:

$Fe^{2+} + 2\ e^- \rightarrow Fe \qquad E^\circ_{red} = -0.44\ V$

$Ag^+ + e^- \rightarrow Ag \qquad E^\circ_{red} = 0.80\ V$

Answer the following questions with respect to the reaction

$$Fe^{2+}(aq) + 2\ Ag(s) \rightarrow Fe(s) + 2\ Ag^+(aq)$$

(a) What is the cell potential, E°, for the reaction?
(b) Is the reaction spontaneous at standard state conditions? Justify your answer.
(c) What is the value of E° at equilibrium?
(d) What is the value of the equilibrium constant at 25°C?
(e) For the spontaneous reaction, what is the maximum amount of work that can be performed?

4. Answer the following questions about reaction kinetics.
(a) Sketch a potential energy versus reaction progress plot for an exothermic reaction, labeling ΔH and E_a. Use collision theory to justify the shape of the curve and to explain how changing the temperature affects the rate of reaction.
(b) On the same plot, sketch a line representing the potential energy versus reaction progress for the same reaction run using a catalyst. Account for the difference in energy between the catalyzed and uncatalyzed reaction.

5. A student weighs out equal amounts of magnesium hydroxide, calcium carbonate, calcium sulfate, and sodium bicarbonate but carelessly forgets to label the containers in which each sample is placed. If the only chemicals available are a bottle of dilute hydrochloric acid and some distilled water, describe a procedure that could be used to identify each solid. Write the net ionic equations for reactions that assist in identification.

6. Nitrogen forms NF_3 but not NF_5 whereas phosphorus forms PF_3 and PF_5. The trifluorides are both trigonal pyramidal and the pentafluoride is trigonal bipyramidal. Draw Lewis diagrams of the molecules and account for these observations.

7. Calcium oxide has a much higher melting point (2580°C) than potassium fluoride (858°C). Use Coulomb's law to explain these observations.

END OF EXAMINATION

ANSWERS TO PRACTICE TEST 2

SECTION I: MULTIPLE-CHOICE QUESTIONS

Score your test using the table below.

Determine how many questions you answered correctly. You will find explanations of the answers on the following pages.

1. B	2. D	3. D	4. C	5. B
6. A	7. C	8. D	9. A	10. B
11. D	12. C	13. B	14. B	15. D
16. B	17. D	18. B	19. B	20. D
21. A	22. C	23. B	24. A	25. B
26. D	27. A	28. D	29. D	30. B
31. A	32. C	33. C	34. B	35. B
36. C	37. A	38. D	39. C	40. D
41. A	42. A	43. C	44. B	45. D
46. D	47. C	48. D	49. A	50. C
51. B	52. A	53. A	54. B	55. B
56. D	57. C	58. B	59. A	60. D

CALCULATE YOUR SCORE:

Number answered correctly: _____

WHAT YOUR SCORE MEANS:

Each year, since the test is different, the scoring is a little different. But generally, if you scored 20 or more on the multiple-choice questions, you'll most likely get a 3 or better on the test. If you scored 28 or more, you'll probably score a 4 or better. And if you scored a 40 or more, you'll most likely get a 5. Keep in mind that the multiple-choice section is worth 50% of your final grade, and the free-response section is worth 50% of your final grade. To learn more about the scoring for the free-response questions, turn to the last page of this section.

ANSWERS AND EXPLANATIONS

SECTION I: MULTIPLE-CHOICE QUESTIONS

1. ANSWER: B In the periodic table, radius increases going down a group because the number of electron shells increases. The radius decreases across a row because the Z_{eff} (number of positive charges in the nucleus) increases (*Chemistry* 8th ed. pages 318–323, 909–910/9th ed. pages 329–334, 928–929). LO 1.9

2. **ANSWER: D** First ionization energy follows the same trend as atomic radius. Less energy is required to remove electrons that are further from the nucleus. More energy is required to remove electrons from atoms with high Z_{eff} (*Chemistry* 8th ed. pages 318–323/9th ed. pages 329–334). LO 2.28

3. **ANSWER: D** Metallic character refers to ability to lose electrons and conductivity. The other substances are nonmetals (*Chemistry* 8th ed. pages 318–323/9th ed. pages 329–334). LO 1.9

4. **ANSWER: C** Group 2 elements have an ns^2 valence configuration and will form 2+ cations (*Chemistry* 8th ed. pages 318–323/9th ed. pages 329–334). LO 1.9

5. **ANSWER: B** Different atomic numbers, therefore different number of protons. Same mass numbers, therefore same total number of protons and neutrons (*Chemistry* 8th ed. pages 49–52/9th ed. pages 52–55). LO 1.4

6. **ANSWER: A** No calculator is necessary to solve this problem. More than half of the substance is the higher mass isotope, so the weighted average mass should be greater than 121. Weighted average = $(0.80 \times 122) + (0.20 \times 120) = 121.6$ (*Chemistry* 8th ed. pages 341–344/9th ed. pages 352–356). LO 1.4

7. **ANSWER: C**

   15 protons means the atomic number is 15, which is P.

   16 neutrons means that the mass number is 15+16 = 31.

   18 electrons compared to 15 protons means that the charge is 3– (*Chemistry* 8th ed. pages 49–52/9th ed. pages 52–55). LO 1.5

8. **ANSWER: D** Molar mass is a physical property; it does not involve chemical change. It is also independent of the amount present (*Chemistry* 8th ed. pages 244–246/9th ed. pages 253–255). LO 3.10

9. **ANSWER: A** High frequency light is high energy. The visible spectrum is bracketed by UV on the high energy side, and IR on the low energy side. Blue is the closest color to UV (*Chemistry* 8th ed. pages 285–287/9th ed. pages 296–298). LO 1.15

10. **ANSWER: B** Each row on the periodic table represents another electron shell. The more electron shells an atom has, the larger its radius (*Chemistry* 8th ed. pages 322–323/9th ed. pages 333–334). LO 1.9

11. **ANSWER: D**

    Kr: 8 valence electrons

    2 F: 14 valence electrons

    Total: 22 valence electrons. Around Kr there are 2 bonding pairs and 3 lone pairs = 5 pairs total

    Therefore: trigonal bipyramidal (*Chemistry* 8th ed. pages 378–388/9th ed. pages 389–399). LO 2.21

12. **ANSWER: C** O_3 has one single bond and one double bond. The double bond can be drawn in either position. The nitrate ion has two single bonds and one double bond; the double bond is drawn in one of three positions (*Chemistry* 8th ed. pages 373–378/9th ed. pages 384–389). LO 2.21

13. **ANSWER: B** Review the rules for writing electron configurations (*Chemistry* 8th ed. pages 312–318/9th ed. pages 322–329). LO 1.6

14. **ANSWER: B** There are two N–O single bonds and one double bond; the average bond order is 1.3 (*Chemistry* 8th ed. pages 373–378/9th ed. pages 384–389). LO 2.21

15. Answer: D Total valence electrons = 2 × 6 for S + 6 for O = 18 total. Divide 18 by 8, for 2 bonding pairs, leaving 2 electrons or one lone pair (*Chemistry* 8th ed. pages 365–369/9th ed. pages 376–380). LO 2.21

16. **ANSWER: B** Reactants are NO and H_2. Products are N_2 and H_2O. Intermediates are N_2O_2 and N_2O; they are formed and used up (*Chemistry* 8th ed. pages 562–565/9th ed. pages 574–577). LO 4.7

17. **ANSWER: D** Slowest step: Rate = $k[N_2O_2][H_2]$, but N_2O_2 is a reaction intermediate. Assume first step reaches steady state: $K' = [N_2O_2]/[NO]^2$, so $[N_2O_2] = K[NO]^2$ and Rate = $kK[NO]^2[H_2]$ (*Chemistry* 8th ed. pages 562–565/9th ed. pages 574–577). LO 4.2

18. **ANSWER: B** Adding a catalyst causes the reaction to follow a different route with a lower activation energy, allowing the process to proceed at a higher rate. It does not affect equilibrium position (amount of product formed overall) (*Chemistry* 8th ed. pages 570–575/9th ed. pages 583–587). LO 4.9

19. **ANSWER: B** This is a Hess's law problem. Sum the three equations, eliminating terms that are the same on both sides.

    $2\ NO(g) \rightarrow \cancel{N_2O_2(g)}$

    $\cancel{N_2O_2(g)} + H_2(g) \rightarrow \cancel{N_2O(g)} + H_2O(g)$

    $\cancel{N_2O(g)} + H_2(g) \rightarrow N_2(g) + H_2O(g)$

    $2\ NO(g) + 2\ H_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$

    (*Chemistry* 8th ed. pages 562–565/9th ed. pages 574–577) LO 4.7

20. **ANSWER: D** The reactants diminish in concentration (C and D). D [NO] goes down 6/4 faster than C [NH_3]. The products are formed (A and B). A [H_2O] is formed 6/5 times faster than B [N_2] (*Chemistry* 8th ed. pages 540–545/9th ed. pages 553–557). LO 4.2

21. **ANSWER: A** Rate at which NO is used = rate of production of N_2 × (6/5) = 36 mol/L/s (*Chemistry* 8th ed. pages 540–545/9th ed. pages 553–557). LO 4.1

22. **ANSWER: C** Rate = $k[A]^2[B] = (2)^2(1/2) = 4 \times ½ = 2$ (*Chemistry* 8th ed. pages 547–551/9th ed. pages 559–563). LO 4.7

23. **ANSWER: B** The partial pressure of the helium = 5.0 atm × (2.0L/8.0L) = 1.25 atm (*Chemistry* 8th ed. pages 199–213/9th ed. pages 208–223). LO 2.6

24. **ANSWER: A** Average speeds increase as the molecular masses decrease (*Chemistry* 8th ed. pages 205–214/9th ed. pages 214–224). LO 2.4

25. **ANSWER: B** Ar is the largest atom, but the reason it deviates from ideal gas behavior is its larger number of electrons, which makes it more polarizable (*Chemistry* 8th ed. pages 214–216/9th ed. pages 224–226). LO 2.12

26. **ANSWER: D** Be familiar with the assumptions underpinning the ideal gas law, and why changing conditions cause deviation from ideal behavior (*Chemistry* 8th ed. pages 183–193/9th ed. pages 192–203). LO 2.4

27. **ANSWER: A** Be familiar with solving problems using the ideal gas law, and remember always to convert temperature from °C to K. Since the volume is not mentioned in the problem statement, it must be constant.

    $P_2 = P_1\ (n_2/n_1)(T_2/T_1) = P_1(0.5\ n_1/n_1)(546\ K/273K)$ so $P_2 = P_1$

    (*Chemistry* 8th ed. pages 183–193/9th ed. pages 192–203). LO 2.4

28. **ANSWER: D** The difference in levels is 1.5 atm. The pressure outside is 1.0 atm; therefore, the total pressure of the gas is 2.5 atm (*Chemistry* 8th ed. pages 181–183/9th ed. pages 190–192). LO 2.4

29. **ANSWER: D** Molarity = number of moles/volume of solution (in liters) = 0.24 moles/0.300 L = 0.80 *M* (*Chemistry* 8th ed. pages 136–145, 498–500/9th ed. pages 145–153, 511–513). LO 2.9

30. **ANSWER: B** There are two sodium ions per formula unit, so the concentration is twice 8.0×10^{-1} *M* (*Chemistry* 8th ed. pages 136–145/9th ed. pages 145–153). LO 2.9

31. **ANSWER: A** Use the relationship $M_1V_1 = M_2V_2$; rearrange to solve for $M_2 = M_1(V_1/V_2) = 0.833\ M\ (0.100\text{L}/2.00\text{L}) = 0.0417\ M$ (*Chemistry* 8th ed. pages 141–144/9th ed. pages 150–153). LO 2.9

32. **ANSWER: C** Sodium carbonate can neutralize two monoprotic acids, whereas sodium hydrogen carbonate can only neutralize one. Remember that when carbonates react with acid, CO_2 is formed in addition to water and a salt (*Chemistry* 8th ed. pages 154–159/9th ed. pages 163–169). LO 3.7

33. **ANSWER: C** Carbonates act as bases when added to acids (*Chemistry* 8th ed. pages 154–159/9th ed. pages 163–169). LO 3.7

34. **ANSWER: B** The first choice is acid-base; the third choice is precipitation. In answer B, the oxidation numbers change so only this choice is oxidation-reduction (*Chemistry* 8th ed. pages 161–166/9th ed. pages 170–175). LO 3.8

35. **ANSWER: B** The Mn in MnO_4^- goes from oxidation number of 7+ to 2+ in Mn^{2+} so it is reduced. Zn is oxidized to Zn^{2+} (*Chemistry* 8th ed. pages 165–166/9th ed. pages 174–175). LO 3.8

36. **ANSWER: C** For the oxidation, $5\ Zn(s) \rightarrow 5\ Zn^{2+}(aq)$, 10 electrons are lost. Ten electrons are gained in the reduction, so the equation is balanced (*Chemistry* 8th ed. pages 166–169/9th ed. pages 175–177). LO 3.9

37. **ANSWER: A** Permanganate is purple; a solution of Mn^{+2} ions is light yellow (*Chemistry* 8th ed. page 962/9th ed. page 981). LO 3.9

38. **ANSWER: D** In $S_4O_6^{2-}$, $4x + 6(-2) = -2$, so $x = +2\ ½$ (*Chemistry* 8th ed. pages 161–166/9th ed. pages 170–175). LO 3.8

39. **ANSWER: C** $C_2H_8N_2 + 2\ N_2O_4 \rightarrow 3\ N_2 + 2\ CO_2 + 4\ H_2O$

(*Chemistry* 8th ed. pages 99–103/9th ed. pages 105–109). LO 3.2

40. **ANSWER: D** In NO_3^-, N has an oxidation number of +5 whereas its oxidation number is +4 in NO_2. Its oxidation number is 0 in N_2, and –3 in NH_3 (*Chemistry* 8th ed. pages 161–166/9th ed. pages 170–175). LO 3.8

41. **ANSWER: A** An exothermic reaction releases heat to the surroundings, whose temperature will increase (*Chemistry* 8th ed. pages 248–249/9th ed. pages 257–258). LO 5.7

42. **ANSWER: A** Entropy of the system decreases in going from 5 moles of gas to 2 moles of gas and 4 of liquid. The exothermic reaction releases heat to the surroundings, whose entropy then increases (*Chemistry* 8th ed. pages 780–783/9th ed. pages 795–798). LO 5.12

43. **ANSWER: C** Draw the Lewis structures and identify the kind and number of bonds. ΔH_{rxn} will be equal to the difference between the energy required to break bonds and the energy released when bonds are formed.

Reactants			Products	
C–H	413 kJ × 3 × 2	C=O	799 kJ × 2 × 2	
C–O	358 kJ × 1 × 2	O–H	467 kJ × 2 × 4	
O–H	467 kJ × 1 × 2			
O=O	495 kJ × 1 × 3			
TOTALS	5613 kJ	–	6932 kJ =	–1319 kJ

(*Chemistry* 8th ed. pages 361–364/9th ed. pages 373–376). LO 5.8

44. **ANSWER: B** An element in its standard state has a defined standard heat of formation of zero (*Chemistry* 8th ed. pages 255–261/9th ed. pages 264–270). LO 5.8

45. **ANSWER: D** The freezing of water involves the release of energy since the water molecules are moving very close to one another to maximize the formation of hydrogen bonds. Bond formation is always exothermic (*Chemistry* 8th ed. pages 798–802/9th ed. pages 813–817). LO 5.6

46. **ANSWER: D** The process is endothermic, so ΔH = (+). There in an increase in disorder so ΔS = (+). The process happens, so ΔG = (–) (*Chemistry* 8th ed. pages 783–794/9th ed. pages 798–810). LO 5.13

47. **ANSWER: C** Compare the values of the terms in the expression

$$\Delta G = \Delta H - T\Delta S$$

Since ΔS is positive, the value of the term $T\Delta S$ is positive and "$-T\Delta S$" will always be negative. When T decreases, the value of "$-T\Delta S$" decreases until it is less than the positive ΔH. At this point, the process will no longer be spontaneous (*Chemistry* 8th ed. pages 783–794/9th ed. pages 798–810). LO 5.13

48. **ANSWER: D** A gas phase molecule has the most entropy, and a solid phase the least. For the liquids, the more complex molecule, H_2O, has more entropy than N_2 (*Chemistry* 8th ed. pages 773–779/9th ed. pages 788–794). LO 5.12

49. **ANSWER: A**

$$K = \frac{[ClF_3]^2}{[F_2]^3[Cl_2]} = \frac{(3)^2}{(2)^3(2.5)} = 0.45$$

(*Chemistry* 8th ed. pages 545–547/9th ed. pages 557–559). LO 6.5

50. **ANSWER: C** Decreasing volume will force the reaction to the side with fewer moles of gas, the right (*Chemistry* 8th ed. pages 620–626/9th ed. pages 633–639). LO 6.8

51. **ANSWER: B** The only difference in a conjugate acid–base pair is the removal of a H^+ going from the acid to the base or the addition of a H^+ going from the base to acid (*Chemistry* 8th ed. pages 642–647/9th ed. pages 656–661). LO 3.7

52. **ANSWER: A** $K_w/K_a = 1.0 \times 10^{-14}/1.0 \times 10^{-7} = 1.0 \times 10^{-7}$ (*Chemistry* 8th ed. pages 651–666/9th ed. pages 666–680). LO 6.12

53. **ANSWER: A** The pair must be a weak electrolyte and its conjugate partner (*Chemistry* 8th ed. pages 701–710/9th ed. pages 715–724). LO 6.18

54. **ANSWER: B**

$K_{sp} = [Cd^{2+}][S^{2-}] = 3.6 \times 10^{-29}$

$[Cd^{2+}] = \sqrt{3.6 \times 10^{-29}} = \sqrt{3.6 \times 10^{-30}} = 6.0 \times 10^{-15}$

(*Chemistry* 8th ed. pages 744–752/9th ed. pages 759–768). LO 6.21

55. **ANSWER: B** Strong intermolecular forces means that additional energy must be supplied in order to increase the physical distance between the molecules when they move from liquid to vapor state. This results in a greater heat of vaporization. It also means that fewer molecules at a given temperature are able to move into the vapor state; thus the vapor pressure is lower than for other liquids with weaker intermolecular forces (*Chemistry* 8th ed. pages 440–443, 504/9th ed. pages 455–458, 517). LO 2.1

56. **ANSWER: D** Substances that are weak electrolytes ionize to a limited extent in water, so the solutions they form are not strongly conductive. Therefore, the ions must be more attracted to each other than they are to polar water. The fact that the solutions are somewhat conductive indicates that the substances do dissociate to some extent (*Chemistry* 8th ed. pages 132–136/9th ed. pages 141–145). LO 2.14

57. **ANSWER: C** $Ag_2CO_3 \rightarrow 2\ Ag^+ + CO_3^{2-}$

Silver nitrate contains Ag^+, a common ion, so this will suppress solubility. Sodium carbonate contains the carbonate ion, a common ion and so this will also suppress solubility.

Calcium nitrate is soluble and will release Ca^{+2}, which reacts with the carbonate ion to form insoluble calcium carbonate. This will

increase solubility because the carbonate ion will be removed from solution (*Chemistry* 8th ed. pages 606–613, 620–626/9th ed. pages 618–626, 633–639). LO 6.23

58. **ANSWER: B** Review definitions of acids and of pH (*Chemistry* 8th ed. pages 647–650/9th ed. pages 661–664). LO 6.12

59. **ANSWER: A** The arrow must point down to show emission and the largest difference corresponds to the greatest difference in energy which means highest frequency and shortest wavelength (*Chemistry* 8th ed. pages 294–300/9th ed. pages 305–310). LO 1.6

60. **ANSWER: D** Change the sign for the oxidation at the anode and add the two half-cell potentials

+2.89 – 0.14 V = +2.75 V

(*Chemistry* 8th ed. pages 876–883/9th ed. pages 894–902). LO 3.12

SECTION II: FREE-RESPONSE QUESTIONS

Question 1: Answers

(a)

$$Q = \frac{P_{PCl_5}}{P_{PCl_3} \times P_{Cl_2}} = \frac{0.10}{(0.30)(0.60)} = 0.56$$

Since 0.56 (Q) > 0.0870 (K), the reaction proceeds to the left.

(b)

	$PCl_3(g)$	+	$Cl_2(g)$	⇌	$PCl_5(g)$
Initial	0.30 atm		0.60 atm		0.10 atm
Change	$+x$		$+x$		$-x$
Equil	$(0.30 + x)$ atm		$(0.60 + x)$ atm		$(0.10 - x)$

Since the reaction proceeds to the left, PCl_5 must decrease and the reactants increase.

$$K_p = 0.0870 = \frac{(0.10 - x)}{(0.30 + x)(0.60 + x)} = \frac{(0.10 - x)}{0.18 + 0.90x + x^2}$$

Using the quadratic equation and solving for x, $x = 0.078$.
Therefore, at equilibrium, $P_{PCl_5} = (0.10 - 0.078) = 0.022$ atm = 0.02 atm

$P_{PCl_3} = (0.30 + 0.078) = 0.378$ atm = 0.38 atm

$P_{PCl_2} = (0.60 + 0.078) = 0.678$ atm = 0.68 atm

(c) Increasing the volume of the container favors the process where more moles of gas are produced, so the reverse reaction is favored; the equilibrium shifts to the left; the mole fraction of PCl_5 decreases.

(d) For an exothermic reaction decreasing the temperature increases the value of K, favoring the products. The partial pressure of PCl_5 increases. Decreasing the

temperature always favors the exothermic direction of a reaction, and increasing the temperature favors the endothermic direction

(*Chemistry* 8th ed. pages 601–604, 606–626/9th ed. pages 614–616, 618–639). LO 6.4, LO 6.8

Question 2: Answers

(a) pH = 1.83 $[H^+] = 1.5 \times 10^{-2}\ M$

$$\text{So percent ionized} = \frac{0.015 \times 100}{0.495} = 3.0\%$$

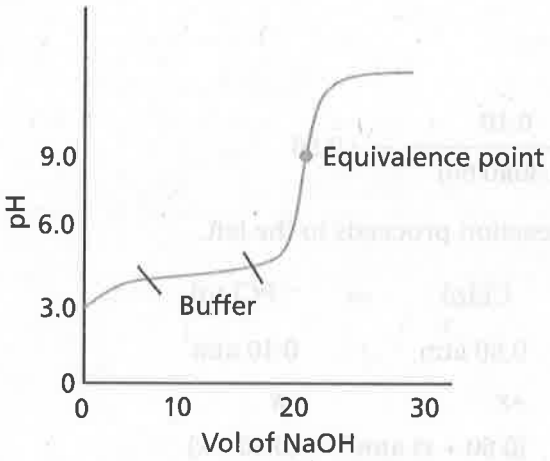
(b)

$$K_a = \frac{[H^+][NO_2^-]}{[HNO_2]} = \frac{(1.5 \times 10^{-2})^2}{0.495} = 4.5 \times 10^{-4}$$

The answer will vary slightly if the ionized acid is subtracted from the initial amount of acid present.

$$K_a = \frac{\left(1.5 \times 10^{-2}\right)^2}{0.480} = 4.7 \times 10^{-4}$$

(c)



(*Chemistry* 8th ed. pages 713–733/9th ed. pages 727–747). LO 6.4, LO 6.12, LO 6.13

Question 3: Answers

(a)

$$Fe^{2+} + 2\,e^- \rightarrow Fe \qquad E^\circ_{red} = -0.44\ V$$

$$2(Ag \rightarrow Ag^+ + e^-) \qquad E^\circ_{oxid} = -0.80\ V$$

$$Fe^{2+} + 2\,Ag \rightarrow 2\,Ag^+ + Fe \qquad E^\circ = -1.24\ V$$

(b) The reaction is not spontaneous because $E^\circ < 0$.

(c) $E = 0$ for any reaction at equilibrium.

(d) You can calculate log K by using the equation

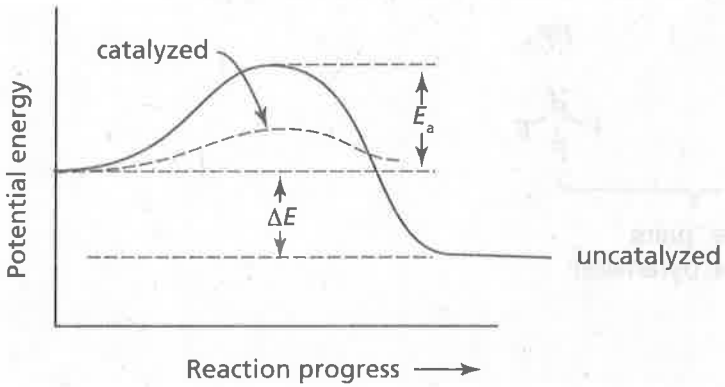
$$\log K = \left(\frac{nE^\circ}{0.0592\ V}\right) V$$

$$= \frac{2(-1.24\ V)}{(0.0592\ V)} = -42.0$$

Taking the antilog of both sides yields $K = 1.0 \times 10^{-42}$

(e) The reaction that is spontaneous is the reverse of the one above, with $E = 1.24$ V.
$W_{max} = -nFE = -(2 \text{ mol } e^-)(96{,}500 \text{ C mol}^{-1})(1.24 \text{ V})(1 \text{J C}^{-1}\text{V}^{-1}) = 2.39 \times 10^5$ J
(*Chemistry* 8th ed. pages 833–842/9th ed. pages 849–858). LO 3.12

Question 4: Answers



(a) Collision theory postulates that molecules must collide in order to react, and that not all collisions lead to reactions. Only when molecules with sufficient kinetic energy collide will there be enough energy to go over the activation energy barrier and create products. In addition, the reactants must approach each other in the proper orientation in order for the collision to result in product formation. Increasing the temperature increases the rate of reaction by increasing the fraction of molecules that have sufficient kinetic energy to overcome the energy barrier (*Chemistry* 8th ed. pages 565–568/9th ed. pages 577–581). LO 4.6

(b) Using a catalyst speeds the rate of reaction by lowering the activation energy barrier, thus increasing the number of collisions with sufficient energy to form products. A catalyst does this by providing a different reaction route, and it is not consumed in the course of the reaction (*Chemistry* 8th ed. pages 570–575/9th ed. pages 583–587). LO 4.9

Question 5: Answers

The net ionic equations for the identification reactions are:

$CaCO_3(s) + 2\ H^+(aq) \rightarrow CO_2(g) + H_2O(l) + Ca^{2+}(aq)$

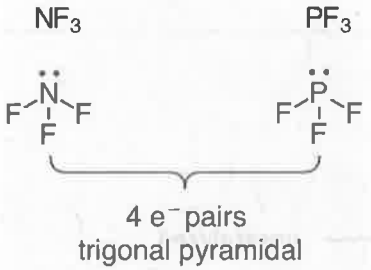
$NaHCO_3(s) + H^+(aq) \rightarrow CO_2(g) + H_2O(l) + Na^+(aq)$

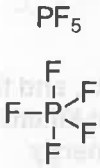
$Mg(OH)_2(s) + 2\ H^+(aq) \rightarrow 2\ H_2O(l) + Mg^{2+}(aq)$

Add water to each solid. The calcium sulfate and the calcium carbonate will be insoluble; however, the magnesium hydroxide and sodium bicarbonate will each be soluble to some degree. Then add the acid to a sample of each insoluble solid; the one which produces bubbles of gas (carbon dioxide) is the calcium carbonate. Then take a sample of each solid that dissolved in water and add acid to each one. The sodium bicarbonate will also produce bubbles of gas (carbon dioxide). The magnesium hydroxide will react with the acid and increase in solubility, but will not produce any gaseous products (*Chemistry* 8th ed. pages 154–160, 218–219, 732/9th ed. pages 163–170, 228–229, 746). LO 2.22

Question 6: Answers

N can form three bonds (NF_3) but not five (NF_5) because it lacks *d* orbitals that are energetically available for the formation of hybrid orbitals (or alternatively, because it is too small to accommodate five atoms). Both NF_3 and PF_3 are trigonal pyramidal because the central atom has three bonding pairs and one lone pair of electrons. PF_5 is trigonal bipyramidal because it has five bonding pairs.





(*Chemistry* 8th ed. pages 378–388, 404–407, 410–411, 932–933/9th ed. pages 389–399, 416–419, 422–423, 952). LO 2.21

Question 7: Answers

Coulomb's law states that the force of attraction between two oppositely charged particles is directly proportional to the product of the absolute value of the charges, and inversely proportional to distance between the ion centers.

$$E = (2.31 \times 10^{-19}\ \text{J} \bullet \text{nm})\left(\frac{Q_1 Q_2}{r}\right)$$

Because of their charges, Ca^{2+} and O^{2-} ions are attracted about four times more strongly than K^+ and F^- ions. Ca^{+2} is smaller than K^+ because it has more protons pulling on the same number of electrons (greater Z_{eff}). Even though O^{2-} is somewhat larger than F^-, overall the calcium-to-oxygen distance is less than the potassium-to-fluoride distance, leading to a smaller distance between ion centers in CaO. Therefore, CaO has a larger numerator and smaller denominator, and a lower solubility than KF (*Chemistry* 8th ed. pages 341–342/9th ed. pages 352–353). LO 2.14

SCORING THE FREE-RESPONSE QUESTIONS

It is difficult to come up with an exact score for this section of test. However, if you compare your answers to the answers in this book, remembering that each part of the test you answer correctly is worth points even if the other parts of the answer are incorrect (see the section titled "Types of Free-Response Questions" on page 12 of this book), you can get a general idea of the percentage of the questions for which you would get credit. If you believe that you got at least one-third of the possible credit, you would probably receive a 3 on this part of the test. If you believe that you would receive close to half or more of the available credit, your score would more likely be a 4 or a 5.