Test 1 Answer Sheet

Answer Sheet: Practice Examination 1

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MC Practice Test 1

SECTION I

Time—90 minutes

Directions: Select the best answer choice and fill in the corresponding oval on the answer sheet. You may NOT use a calculator for this section. You may use the Periodic Table provided on page 469 of this book.

1. Which of the following statements about nuclear stability is correct?
   A. Heavier, more stable nuclei have somewhat larger numbers of protons than neutrons.
   B. A stable nucleus cannot undergo a nuclear reaction, even with the addition of external energy.
   C. Unstable nuclei do not spontaneously change to stable nuclei.
   D. Lighter nuclei tend to have equal numbers of protons and neutrons.
   E. Heavier nuclei have significantly more neutrons than protons.

2. Rank the following in order of increasing acidity: propionic acid ($K_a = 1.3 \times 10^{-5}$), benzoic acid ($K_a = 6.3 \times 10^{-5}$), hypobromous acid ($K_a = 2.6 \times 10^{-9}$).
   A. Benzoic acid < propionic acid < hypobromous acid
   B. Hypobromous acid < propionic acid < benzoic acid
   C. Propionic acid < benzoic acid < hypobromous acid
   D. Hypobromous acid < benzoic acid < propionic acid
   E. Benzoic acid < hypobromous acid < propionic acid

3. Which of the following compounds contain(s) no covalent bonds?

   KCl PH_3 O_2 B_2H_6 H_2SO_4

   A. KCl, PH_3, and B_2H_6 only
   B. KCl and H_2SO_4 only
   C. PH_3, O_2, and B_2H_6 only
   D. KCl only
   E. KCl and B_2H_6 only

4. Which of the following nuclear reactions is *incorrect*?
   A. ${}^{14}_{7}N + {}^{4}_{2}He \rightarrow {}^{17}_{8}O + {}^{1}_{1}H$
   B. ${}^{9}_{4}Be + {}^{4}_{2}He \rightarrow {}^{12}_{6}C + {}^{1}_{0}n$
   C. ${}^{30}_{15}P + {}^{-30}_{-14}Si \rightarrow {}^{0}_{-1}\beta$
   D. ${}^{3}_{1}H + {}^{3}_{2}He \rightarrow {}^{1}_{-1}\beta$
   E. None of the above

5. What is the hydrogen ion concentration in a solution of 0.00200 M potassium hydroxide?
   A. $[H^+] = \dfrac{K_w}{0.00200}$
   B. $[H^+] = K_w(0.00200)$
   C. $[H^+] = \dfrac{0.00200}{K_w}$
   D. $[H^+] = -\log\dfrac{K_w}{0.00200}$
   E. $[H^+] = -\log\dfrac{0.00200}{K_w}$

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6. What is the correct coefficient on the Fe^{3+} ion when the following reaction is balanced?

$$Fe^{2+} + ClO_3^- + H^+ \rightarrow Fe^{3+} + Cl^- + H_2O$$

A. 2
B. 3
C. 5
D. 6
E. 7

7. The van der Waals equation includes 2 terms, "a" and "b," that are not present in the ideal gas law. Which of the following statements is true about these terms?

A. "a" corrects for attractive forces between gas particles and "b" corrects for the volume of the container.

B. "a" corrects for the external pressure of the gas and "b" corrects for the internal pressure of the gas.

C. "a" corrects for interactions between gas particles and "b" corrects for the volume of the gas particles.

D. "a" corrects for the volume of the gas particles and "b" corrects for the repulsive forces between gas particles.

E. None of the above

8. Which of the following molecular structures is not possible?

A. OF_2

B. SF_2

C. OF_4

D. SF_4

E. O_2F_2

9. Which of the following is *not* a postulate of the kinetic molecular theory?

A. Gas molecules travel in random, straight paths.

B. The energy of a gas molecule is determined by quantum mechanics.

C. The collisions between gas molecules are elastic.

D. The absolute temperature of a substance is equal to the average kinetic energy of its particles.

E. The particles of a sample of gas have no volume.

10. Which parameter of a chemical reaction will change with a catalyst?

A. Free energy change
B. Entropy change
C. Equilibrium constant
D. Rate constant
E. Enthalpy change

11. A plot of $1/[NO_2]$ versus time for the decomposition of NO_2 was found to be linear. This means that the reaction:

A. is zero order with respect to NO_2.

B. is first order with respect to NO_2.

C. is second order reaction with respect to NO_2.

D. is third order reaction with respect to NO_2.

E. Order cannot be determined from the information given.

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12. Which of the following is not a characteristic of ionic substances?
    A. High melting point
    B. Fragility
    C. Crystalline (in the solid form)
    D. Deforms when struck
    E. Well-defined three-dimensional structure

13. Which of the following behaves most like an ideal gas?
    A. H_2
    B. He
    C. O_2
    D. CO_2
    E. Ne

14. The change in enthalpy of a system is equal to Q, the heat flow between the system and the surroundings, under which conditions?
    A. Constant volume
    B. Constant pressure
    C. Constant temperature
    D. Absence of pressure-volume work
    E. None of the above

15. Which statement about metals is *incorrect*?
    A. Metals exhibit higher electronegativities than non-metals.
    B. Metals are reducing agents.
    C. Metals form basic hydroxides.
    D. Metals exhibit low ionization potentials.
    E. Metals generally have one to five electrons in their outermost shell.

16. 40 L of an ideal gas at 25°C and 750 mmHg is allowed to expand to 50 L and the pressure is increased to 765 mmHg. What is the final temperature of the gas?
    A. $\frac{(298)(750)(50)}{(40)(765)}$
    B. $\frac{(298)(765)(50)}{(40)(750)}$
    C. $\frac{(298)(750)(40)}{(50)(765)}$
    D. $\frac{(750)(40)}{(50)(298)(765)}$
    E. $\frac{(298)(765)(40)}{(50)(750)}$

17. Rutherford's scattering experiments demonstrated:
    A. the existence of X-rays.
    B. the existence of α-particles.
    C. the nature of blackbody radiation.
    D. the mass-to-charge ration of the electron.
    E. the nuclear model of the atom.

18. Elements not found in nature, synthesized in nuclear reactions and involving the completion of the 5*f* atomic orbitals are known as:
    A. lanthanides.
    B. halogens.
    C. actinides.
    D. transition metals.
    E. rare gases.

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19. Which of the following statements is (are) true about the half-reactions shown below?

I. As written, the standard potential for the overall reaction is –1.56 V.

II. As written, the overall reaction is spontaneous.

III. The sign of the potential of the overall reaction indicates whether or not the reaction is spontaneous.

$Ag \rightarrow Ag^+ + e^- \quad E° = -0.80\ V$

$Zn^{2+} + 2\ e^- \rightarrow Zn \quad E° = -0.76\ V$

A. I only
B. II only
C. III only
D. I and III only
E. II and III only

20. K_p for the reaction below is 1.36 at 499 K. Which of the following equations can be used to calculate K_c for this reaction?

$$PCl_3\ (g) + Cl_2\ (g) \rightarrow PCl_5\ (g)$$

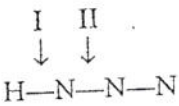
A. $K_c = \dfrac{[(0.0821)(499)]^{-1}}{1.36}$

B. $K_c = \dfrac{1.36(0.0821)}{499}$

C. $K_c = \dfrac{1.36}{[(0.0821)(499)]^{-1}}$

D. $K_c = 1.36 \cdot [(0.0821)(499)]^{-1}$

E. $K_c = \dfrac{1.36(499)}{0.0821}$



21. The Third Law of Thermodynamics states that at absolute zero temperature, all perfect crystals have:

A. the same crystal lattice.
B. the same lattice energy.
C. the same enthalpy.
D. the same free energy.
E. the same entropy.

22. Which of the following has the smallest ionic radius?

A. Li^+
B. Na^+
C. K^+
D. Rb^+
E. Cs^+

23. The van der Waals equation for non-ideal gases differs from the ideal gas law in that it accounts for:

I. the mass of each particle of gas.
II. the volume of each particle of gas.
III. the attractive forces between particles of gas.

A. I only
B. II only
C. III only
D. I and III only
E. II and III only

24. What is the molarity of a sulfuric acid solution if 50.0 mL completely neutralizes 1.00 L of a 0.10 M potassium hydroxide solution?

A. 1.0 M
B. 0.10 M
C. 2.0 M
D. 0.20 M
E. 10.0 M

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25. Which of the following is an intensive property of a system?
    A. Pressure
    B. Mass
    C. Enthalpy
    D. Volume
    E. None of the above

26. Which of the following salts produces the most basic aqueous solution?
    A. $Al(CN)_3$
    B. $KC_2H_3O_2$
    C. $FeCl_3$
    D. KCl
    E. $Pb(C_2H_3O_2)_2$

27. How many molecules are present in 0.20 g of hydrogen gas?
    A. $\frac{0.20}{1.008} \cdot 6.02 \times 10^{23}$
    B. $0.20 \cdot 1.008$
    C. $0.20 \cdot 2.016$
    D. $\frac{0.20}{2.016} \cdot 6.02 \times 10^{23}$
    E. $\frac{0.20}{6.02 \times 10^{23}} \cdot 2.016$

28. The least accurate of the volumetric measuring devices is a:
    A. pipet.
    B. buret.
    C. volumetric flask.
    D. graduated cylinder.
    E. beaker.

29. Which of the following describes the equilibrium constant for a spontaneous reaction?
    A. $K = 0$
    B. $K < 0$
    C. $K = 1$
    D. $K < 1$
    E. $K > 1$

30. Which of the following oxides is amphoteric?
    A. Na_2O
    B. ZnO
    C. MgO
    D. Cl_2O_7
    E. P_2O_5

31. It takes 250.0 J to raise the temperature of a 50.0 g sample of a metal by 10.0°C. What is the specific heat capacity of this metal?
    A. 5.00×10^{-4} J/g·K
    B. 1.25×10^{5} J/g·K
    C. 0.500 J/g·K
    D. 1.25×10^{2} kJ/g·K
    E. 50.0 J/g·K

32. What is the molecular geometry of IF_5?
    A. Tetrahedral
    B. Trigonal bipyramidal
    C. Square pyramidal
    D. Octahedral
    E. Seesaw

33. How much barium nitrate is required to prepare 250.0 mL of a 0.100 M solution? (The molar mass of barium nitrate is 199.344.)

A. $\frac{(250.0)(199.344)}{(1000)(0.100)}$

B. $\frac{(250.0)(0.100)}{(1000)(199.344)}$

C. $\frac{(250.0)(0.100)(199.344)}{(1000)}$

D. $\frac{(250.0)}{(1000)(199.344)(0.100)}$

E. None of the above

34. What is the oxidation number of chlorine in ClO_4^-?

A. +1
B. +3
C. +5
D. +7
E. +8

35. A neutral atom has the ground-state electron configuration $1s^2 2s^2 2p^6 3s^1$. It will gain or lose electrons to form an ion of charge:

A. –2.
B. –1.
C. +1.
D. +2.
E. +3.

36. The emission of an alpha particle from $^{226}_{88}Ra$ will yield:

A. $^{223}_{86}Rn$.

B. $^{222}_{86}Rn$.

C. $^{223}_{87}Fr$.

D. $^{222}_{87}Fr$.

E. $^{222}_{88}Ra$.

37. The Rydberg equation was a very useful result of Bohr's model of the hydrogen atom. The Rydberg equation gives:

A. the velocity of electrons as they move through spectral lines.

B. the rate of absorption of hydrogen atoms in the ultraviolet region.

C. the rate of emission of hydrogen atoms in the Lyman series.

D. the rate of emission of heated hydrogen atoms.

E. the frequencies of the series of lines in the hydrogen spectrum.

38. Which of the characteristics below is *not* necessary for a reaction to be used in a titration?

A. The reaction can have no side reactions.

B. The equilibrium constant of the reaction must be very large.

C. The reaction should proceed according to a definite chemical equation.

D. The reaction should proceed very slowly so that the endpoint is readily observable.

E. A method should be available to indicate when to stop the titration.

39. Which of the following molecules or ions is linear?

A. H_2O

B. ClO_2^-

C. NO_2^-

D. NO_2

E. NO_2^+

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40. Which of the following statements about equilibrium is correct?

I. Equilibrium is reached when $\Delta G = 0$.

II. $\Delta G° = -RT\ln K_{eq}$

III. At equilibrium, $\Delta G°$ is dependent on pressure.

A. I only
B. II only
C. III only
D. I and II only
E. I, II, and III

41. The triple point pressure of water is 4.58 mmHg and the triple point temperature is 273.16 K. From this we can conclude that:

A. steam cannot exist at temperatures below 273.16 K.

B. the vapor pressure of ice is 4.58 mmHg for temperatures below 273.16 K.

C. the vapor pressure of water is less than 4.58 mmHg in most cases.

D. liquid water cannot exist at pressures below 4.58 mmHg.

E. ice cannot exist at pressures below 4.58 mmHg.

42. The molecular geometry of the ammonium ion, NH_4^+, is:

A. trigonal planar.
B. trigonal pyramidal.
C. square planar.
D. tetrahedral.
E. octahedral.

43. Which of the following statements is (are) true about an oxygen atom in the ground state?

I. Electrons in the 1*s* atomic orbital may be described using the quantum numbers (1, 0, 0, +1/2) and (1, 0, 0, −1/2).

II. Electrons in the 2*s* atomic orbital may be described using the quantum numbers (2, 1, 1, +1/2) and (2, 1, 1, −1/2).

III. The fourth quantum number, m_s, describes the ways that an electron may be aligned with a magnetic field.

A. I only
B. I and II only
C. I and III only
D. II and III only
E. I, II, and III

44. LeChâtelier's principle states that:

A. equilibrium is only reached under certain conditions of temperature and pressure.

B. when stress is applied to a system at equilibrium, the reaction shifts in the direction that minimizes the stress.

C. increasing the temperature while decreasing the pressure increases the equilibrium constant.

D. neither temperature nor pressure has a major effect on equilibrium.

E. equilibrium is eventually obtained, regardless of reaction conditions.

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45. Which of the following statements about carbon-containing compounds is (are) true?

I. Carbon monoxide is produced when carbon is burned with insufficient oxygen.

II. Carbon dioxide may undergo sublimation.

III. Carbonic acid is a diprotic acid.

A. I only
B. II only
C. III only
D. I and III only
E. I, II, and III

46. An ionic bond is formed between two ions. Which of the following has no effect on the strength of the bond?

I. Doubling the charge on both ions
II. Doubling the temperature
III. Doubling the radii of both ions

A. I only
B. II only
C. III only
D. Both I and II
E. Both II and III

47. Which of the alkali metals is most electronegative?

A. Li
B. Na
C. K
D. Rb
E. Cs

48. How many carbon atoms are there in 27.3 g of trichloroacetic acid? (The molar mass of trichloroacetic acid is 163.5.)

A. $\frac{(27.3)(2)(6.02 \times 10^{23})}{163.5}$

B. $\frac{(27.3)(2)}{163.5}$

C. $\frac{(27.3)(163.5)}{(2)(6.02 \times 10^{23})}$

D. $\frac{(27.3)(6.02 \times 10^{23})(163.5)}{(2)(12.01)}$

E. $\frac{(27.3)(6.02 \times 10^{23})}{163.5}$

49. The primary weakness of the Bohr model of the atom is that:

A. it only works for the hydrogen atom.

B. it treats the electron as a wave rather than a particle.

C. it doesn't consider the role of the neutron.

D. it neglects the radiation emitted by accelerating charged particles.

E. it only allows for certain energy levels.

50. Which of the following statements is not true for the reaction shown below?

$$Fe^{3+} + 1\ e^- \rightarrow Fe^{2+}$$

A. Fe^{3+} is being reduced.

B. The oxidation state of Fe has changed.

C. Fe^{3+} is the oxidizing agent in this reaction.

D. The reaction is similar to the reaction between magnesium metal and hydrogen gas.

E. Both Fe^{3+} and Fe^{2+} are anions.

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51. How much heat is released when the temperature of 100 g of water decreases from 25°C to 5°C? (The specific heat of water is 4.18 J/g · K.)

A. $(100)(4.18)(25 - 5)$

B. $(100)(4.18)(5 - 25)$

C. $\frac{2000}{4.18}$

D. $\frac{-2000}{4.18}$

E. $\frac{-20(4.18)}{100}$

52. Increasing the temperature of a reaction increases:

A. the reaction order.

B. the activation energy.

C. the number of collisions in the correct orientation.

D. the kinetic energy of the molecules.

E. None of the above.

53. Glass is an example of an amorphous solid which can be characterized as:

A. a malleable solid.
B. crystal-like in structure.
C. a good conductor.
D. a molecular solid.
E. a very viscous fluid.

54. The K_a of the ammonium ion is 5.6×10^{-10} at 25°C. What is the approximate pH of a 1.0 M ammonium chloride solution?

A. 9
B. 7
C. 5
D. 3
E. 1

55. Which of the following pairs of elements does not have approximately the same electronegativity?

A. C and S
B. Co and Ni
C. B and Al
D. U and Pu
E. Fe and Ni

56. Which of the following salts will produce a basic solution when dissolved in water?

A. NH_4Cl

B. NaCl

C. $NaNO_3$

D. Na_2SO_4

E. Na_2CO_3

57. A pH 7 buffer solution contains H_2CO_3 and $NaHCO_3$. What must the ratio of $[NaHCO_3]/[H_2CO_3]$ be in order to maintain the solution at pH 7? The K_a of H_2CO_3 is 4.3×10^{-7}.

A. 43
B. 4.3
C. 0.43
D. 86
E. 1.29

58. For which of the following compounds is hydrogen bonding an important component of the intermolecular forces?

A. CH_3Cl

B. CH_3OCH_3

C. CH_3NH_2

D. CH_3CH_3Cl

E. None of the above

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59. The rate data for the reaction $A + B \rightarrow C$ is shown below.

[A]	[B]	rate
1.0	1.0	0.01
1.0	2.0	0.02
3.0	1.0	0.09

The reaction is:

A. first order in both A and B.
B. second order in both A and B.
C. first order in A and second order in B.
D. second order in A and first order in B.
E. second order in A and zero order in B.

60. During a redox reaction, the oxidizing agent:

A. gains electrons.
B. is oxidized.
C. has an increase in oxidation state.
D. is hydrolyzed.
E. loses electrons.

61. Which of the following statements about semiconductors is (are) true?

I. A p-type semiconductor is formed when silicon or germanium is doped with a Group III element.

II. Doping decreases the conductivity of a silicon or germanium crystal.

III. An n-type semiconductor is formed when silicon or germanium is doped with an element that produces non-bonded electrons.

A. I only
B. II only
C. III only
D. I and II only
E. I and III only

62. What is the frequency of light of wavelength 3×10^{-3} cm?

A. 1×10^{13}
B. 2.2×10^{-31}
C. 1×10^{7}
D. 9×10^{7}
E. 2.64×10^{-36}

63. Which of the following substances is a Lewis acid?

A. CCl_4
B. BF_3
C. I_2
D. NaH
E. $(CH_3)_3N$

64. Which of the following compounds has the shortest carbon-halogen bond?

A. CH_3F
B. CH_3Cl
C. CH_3Br
D. CH_3I
E. They are all equal.

65. All of the following statements about entropy change are true EXCEPT:

A. It is a measure of the energy dispersal.
B. The natural tendency is for it to increase.
C. It is not a state function under all conditions.
D. It can be defined both thermodynamically and statistically.
E. Its calculation is only possible for processes involving no temperature change.

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66. Which of the following is not a method for separating mixtures?

A. Filtration
B. Distillation
C. Selective precipitation
D. Absorption chromatography
E. Solvation

67. An atom containing two electrons which possess the following sets of quantum numbers $(3, 1, 1, -\frac{1}{2})$ and $(3, 1, 1, -\frac{1}{2})$ may not exist based on:

A. Pauli exclusion principle.
B. Lewis's law.
C. Hund's rule.
D. Heisenberg uncertainty principle.
E. Bohr model.

68. The natural logarithm of the rate constant of a reaction is:

A. directly proportional to temperature.
B. inversely proportional to temperature.
C. not affected by changes in temperature.
D. only affected by the activation energy.
E. independent of the activation energy and the temperature.

69. Which of the following compounds does not contain a covalent bond?

A. PH_3

B. $GeCl_4$

C. H_2S

D. CsF

E. CH_3Cl

70. In the most stable resonance form of the molecule whose skeleton structure is shown above, the bond orders of bonds I and II are:

A. I = 2 and II = 1
B. I = 2 and II = 20
C. I > 2 and II < 2
D. I > 2 and II > 2
E. I < 2 and II < 2

Match the descriptions in questions 71–73 with the choices given below. A choice may be used once, more than once, or not at all.

(A) BF_3

(B) C_2H_2

(C) $CHCl_3$

(D) XeF_4

(E) NO_2

71. A molecule that has an unpaired electron within its structure.

72. A molecule whose shape is square planar.

73. A molecule that contains 2 pi (π) bonds.

74. A weak acid has a K_a of 1.0×10^{-10}. What is the $[H_3O^+]$ ion concentration in a 0.01 M solution of this acid?

A. 1.0×10^{-12} M

B. 1.0×10^{-10} M

C. 1.0×10^{-8} M

D. 1.0×10^{-6} M

E. 1.0×10^{-5} M

75. All of the following statements are correct EXCEPT:

A. In all spontaneous processes, $\Delta S_{universe} > 0$.

B. The entropy of a perfect crystal is taken to be 0 at 298 K.

C. During freezing, the entropy of the system decreases.

D. A spontaneous process is accompanied by a negative free energy change.

E. The free energy change of a system at equilibrium is 0.

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Practice Test 1

INFORMATION IN THE TABLE BELOW AND IN THE TABLES ON PAGES 3-5 MAY BE USEFUL IN ANSWERING THE QUESTIONS IN THIS SECTION OF THE EXAMINATION.

DO NOT DETACH FROM BOOK.

PERIODIC TABLE OF THE ELEMENTS

1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.012											5 B 10.811	6 C 12.011	7 N 14.007	8 O 16.00	9 F 19.00	10 Ne 20.179
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.974	16 S 32.06	17 Cl 35.453	18 Ar 39.948
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.938	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	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 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.2	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.02	89 †Ac 227.03	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 § (269)	111 § (272)	112 § (277)						

§Not yet named

*Lanthanide Series	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
†Actinide Series	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	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)

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STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			$E°$(V)
$Li^+ + e^-$	→	$Li(s)$	−3.05
$Cs^+ + e^-$	→	$Cs(s)$	−2.92
$K^+ + e^-$	→	$K(s)$	−2.92
$Rb^+ + e^-$	→	$Rb(s)$	−2.92
$Ba^{2+} + 2\,e^-$	→	$Ba(s)$	−2.90
$Sr^{2+} + 2\,e^-$	→	$Sr(s)$	−2.89
$Ca^{2+} + 2\,e^-$	→	$Ca(s)$	−2.87
$Na^+ + e^-$	→	$Na(s)$	−2.71
$Mg^{2+} + 2\,e^-$	→	$Mg(s)$	−2.37
$Be^{2+} + 2\,e^-$	→	$Be(s)$	−1.70
$Al^{3+} + 3\,e^-$	→	$Al(s)$	−1.66
$Mn^{2+} + 2\,e^-$	→	$Mn(s)$	−1.18
$Zn^{2+} + 2\,e^-$	→	$Zn(s)$	−0.76
$Cr^{3+} + 3\,e^-$	→	$Cr(s)$	−0.74
$Fe^{2+} + 2\,e^-$	→	$Fe(s)$	−0.44
$Cr^{3+} + e^-$	→	Cr^{2+}	−0.41
$Cd^{2+} + 2\,e^-$	→	$Cd(s)$	−0.40
$Tl^+ + e^-$	→	$Tl(s)$	−0.34
$Co^{2+} + 2\,e^-$	→	$Co(s)$	−0.28
$Ni^{2+} + 2\,e^-$	→	$Ni(s)$	−0.25
$Sn^{2+} + 2\,e^-$	→	$Sn(s)$	−0.14
$Pb^{2+} + 2\,e^-$	→	$Pb(s)$	−0.13
$2\,H^+ + 2\,e^-$	→	$H_2(g)$	0.00
$S(s) + 2\,H^+ + 2\,e^-$	→	$H_2S(g)$	0.14
$Sn^{4+} + 2\,e^-$	→	Sn^{2+}	0.15
$Cu^{2+} + e^-$	→	Cu^+	0.15
$Cu^{2+} + 2\,e^-$	→	$Cu(s)$	0.34
$Cu^+ + e^-$	→	$Cu(s)$	0.52
$I_2(s) + 2\,e^-$	→	$2\,I^-$	0.53
$Fe^{3+} + e^-$	→	Fe^{2+}	0.77
$Hg_2^{2+} + 2\,e^-$	→	$2\,Hg(l)$	0.79
$Ag^+ + e^-$	→	$Ag(s)$	0.80
$Hg^{2+} + 2\,e^-$	→	$Hg(l)$	0.85
$2\,Hg^{2+} + 2\,e^-$	→	Hg_2^{2+}	0.92
$Br_2(l) + 2\,e^-$	→	$2\,Br^-$	1.07
$O_2(g) + 4\,H^+ + 4\,e^-$	→	$2\,H_2O(l)$	1.23
$Cl_2(g) + 2\,e^-$	→	$2\,Cl^-$	1.36
$Au^{3+} + 3\,e^-$	→	$Au(s)$	1.50
$Co^{3+} + e^-$	→	Co^{2+}	1.82
$F_2(g) + 2\,e^-$	→	$2\,F^-$	2.87

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ADVANCED PLACEMENT CHEMISTRY EQUATIONS AND CONSTANTS

ATOMIC STRUCTURE

$$\Delta E = h\nu$$
$$c = \lambda\nu$$
$$\lambda = \frac{h}{mv}$$
$$p = mv$$
$$E_n = \frac{-2.178 \times 10^{-18}}{n^2} \text{ joule}$$

E = energy
ν = frequency
λ = wavelength
p = momentum
v = velocity
n = principal quantum number
m = mass

EQUILIBRIUM

$$K_a = \frac{[H^+][A^-]}{[HA]}$$
$$K_b = \frac{[OH^-][HB^+]}{[B]}$$
$$K_w = [OH^-][H^+] = 1.0 \times 10^{-14} \text{ @ } 25°C$$
$$= K_a \times K_b$$
$$pH = -\log[H^+],\ pOH = -\log[OH^-]$$
$$14 = pH + pOH$$
$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$
$$pOH = pK_b + \log\frac{[HB^+]}{[B]}$$
$$pK_a = -\log K_a,\ pK_b = -\log K_b$$
$$K_p = K_c(RT)^{\Delta n},$$

where Δn = moles product gas − moles reactant gas

Speed of light, $c = 3.0 \times 10^8$ m s^{-1}
Planck's constant, $h = 6.63 \times 10^{-34}$ J s
Boltzmann's constant, $k = 1.38 \times 10^{-23}$ J K^{-1}
Avogadro's number = 6.022×10^{23} molecules mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb
1 electron volt per atom = 96.5 kJ mol^{-1}

Equilibrium Constants
K_a (weak acid)
K_b (weak base)
K_w (water)
K_p (gas pressure)
K_c (molar concentrations)

THERMOCHEMISTRY

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$
$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$
$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$
$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$
$$= -RT \ln K = -2.303\,RT \log K$$
$$= -n\mathcal{F}E^\circ$$
$$\Delta G = \Delta G^\circ + RT \ln Q = \Delta G^\circ + 2.303\,RT \log Q$$
$$q = mc\Delta T$$
$$C_p = \frac{\Delta H}{\Delta T}$$

S° = standard entropy
H° = standard enthalpy
G° = standard free energy
E° = standard reduction potential
T = temperature
n = moles
m = mass
q = heat
c = specific heat capacity
C_p = molar heat capacity at constant pressure
1 faraday $\mathcal{F}$ = 96,500 coulombs

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GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT$$

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

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

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

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

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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

$$u_{rms} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{M}}$$

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

$$KE \text{ per mole} = \frac{3}{2}RT$$

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

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

$$\text{molality} = \text{moles solute per kilogram solvent}$$

$$\Delta T_f = iK_f \times \text{molality}$$

$$\Delta T_b = iK_b \times \text{molality}$$

$$\pi = \frac{nRT}{V}i$$

OXIDATION-REDUCTION; ELECTROCHEMISTRY

$$Q = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b}, \text{ where } a\text{ A} + b\text{ B} \rightarrow c\text{ C} + d\text{ D}$$

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

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF}\ln Q = E^\circ_{cell} - \frac{0.0592}{n}\log Q \text{ @ } 25^\circ C$$

$$\log K = \frac{nE^\circ}{0.0592}$$

P = pressure
V = volume
T = temperature
n = number of moles
D = density
m = mass
v = velocity

u_{rms} = root-mean-square speed
KE = kinetic energy
r = rate of effusion
M = molar mass
π = osmotic pressure
i = van't Hoff factor
K_f = molal freezing-point depression constant
K_b = molal boiling-point elevation constant
Q = reaction quotient
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
E° = standard reduction potential
K = equilibrium constant

Gas constant, R = 8.31 J mol^{-1} K^{-1}
= 0.0821 L atm mol^{-1} K^{-1}
= 8.31 volt coulomb mol^{-1} K^{-1}
Boltzmann's constant, k = 1.38×10^{-23} J K^{-1}
K_f for H_2O = 1.86 K kg mol^{-1}
K_b for H_2O = 0.512 K kg mol^{-1}
STP = 0.000°C and 1.000 atm
Faraday's constant, F = 96,500 coulombs per mole of electrons

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2001 AP® CHEMISTRY FREE-RESPONSE QUESTIONS

CHEMISTRY
Section II
(Total time—90 minutes)

Part A
Time—40 minutes
YOU MAY USE YOUR CALCULATOR FOR PART A.

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.

Be sure to write all your answers to the questions on the lined pages following each question in the booklet with the pink cover. Do NOT write your answers on the green insert.

Answer Question 1 below. The Section II score weighting for this question is 20 percent.

1. Answer the following questions relating to the solubility of the chlorides of silver and lead.

(a) At 10°C, 8.9×10^{-5} g of $AgCl(s)$ will dissolve in 100. mL of water.

(i) Write the equation for the dissociation of $AgCl(s)$ in water.

(ii) Calculate the solubility, in mol L^{-1}, of $AgCl(s)$ in water at 10°C.

(iii) Calculate the value of the solubility-product constant, K_{sp}, for $AgCl(s)$ at 10°C.

(b) At 25°C, the value of K_{sp} for $PbCl_2(s)$ is 1.6×10^{-5} and the value of K_{sp} for $AgCl(s)$ is 1.8×10^{-10}.

(i) If 60.0 mL of 0.0400 M $NaCl(aq)$ is added to 60.0 mL of 0.0300 M $Pb(NO_3)_2(aq)$, will a precipitate form? Assume that volumes are additive. Show calculations to support your answer.

(ii) Calculate the equilibrium value of $[Pb^{2+}(aq)]$ in 1.00 L of saturated $PbCl_2$ solution to which 0.250 mole of $NaCl(s)$ has been added. Assume that no volume change occurs.

(iii) If 0.100 M $NaCl(aq)$ is added slowly to a beaker containing both 0.120 M $AgNO_3(aq)$ and 0.150 M $Pb(NO_3)_2(aq)$ at 25°C, which will precipitate first, $AgCl(s)$ or $PbCl_2(s)$? Show calculations to support your answer.

GO ON TO THE NEXT PAGE.

Answer EITHER Question 2 below OR Question 3 printed on page 8. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

$$2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g) \qquad \Delta H^\circ = -114.1\ kJ,\ \Delta S^\circ = -146.5\ J\ K^{-1}$$

2. The reaction represented above is one that contributes significantly to the formation of photochemical smog.

   (a) Calculate the quantity of heat released when 73.1 g of $NO(g)$ is converted to $NO_2(g)$.

   (b) For the reaction at 25°C, the value of the standard free-energy change, ΔG°, is −70.4 kJ.

   (i) Calculate the value of the equilibrium constant, K_{eq}, for the reaction at 25°C.

   (ii) Indicate whether the value of ΔG° would become more negative, less negative, or remain unchanged as the temperature is increased. Justify your answer.

   (c) Use the data in the table below to calculate the value of the standard molar entropy, S°, for $O_2(g)$ at 25°C.

	Standard Molar Entropy, S° ($J\ K^{-1}\ mol^{-1}$)
$NO(g)$	210.8
$NO_2(g)$	240.1

   (d) Use the data in the table below to calculate the bond energy, in $kJ\ mol^{-1}$, of the nitrogen-oxygen bond in NO_2. Assume that the bonds in the NO_2 molecule are equivalent (i.e., they have the same energy).

	Bond Energy ($kJ\ mol^{-1}$)
Nitrogen-oxygen bond in NO	607
Oxygen-oxygen bond in O_2	495
Nitrogen-oxygen bond in NO_2	?

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3. Answer the following questions about acetylsalicylic acid, the active ingredient in aspirin.

   (a) The amount of acetylsalicylic acid in a single aspirin tablet is 325 mg, yet the tablet has a mass of 2.00 g. Calculate the mass percent of acetylsalicylic acid in the tablet.

   (b) The elements contained in acetylsalicylic acid are hydrogen, carbon, and oxygen. The combustion of 3.000 g of the pure compound yields 1.200 g of water and 3.72 L of dry carbon dioxide, measured at 750. mm Hg and 25°C. Calculate the mass, in g, of each element in the 3.000 g sample.

   (c) A student dissolved 1.625 g of pure acetylsalicylic acid in distilled water and titrated the resulting solution to the equivalence point using 88.43 mL of 0.102 *M* NaOH(*aq*). Assuming that acetylsalicylic acid has only one ionizable hydrogen, calculate the molar mass of the acid.

   (d) A 2.00×10^{-3} mole sample of pure acetylsalicylic acid was dissolved in 15.00 mL of water and then titrated with 0.100 *M* NaOH(*aq*). The equivalence point was reached after 20.00 mL of the NaOH solution had been added. Using the data from the titration, shown in the table below, determine

   (i) the value of the acid dissociation constant, K_a, for acetylsalicylic acid and

   (ii) the pH of the solution after a total volume of 25.00 mL of the NaOH solution had been added (assume that volumes are additive).

Volume of 0.100 *M* NaOH Added (mL)	pH
0.00	2.22
5.00	2.97
10.00	3.44
15.00	3.92
20.00	8.13
25.00	?

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CHEMISTRY

Part B
Time—50 minutes

NO CALCULATORS MAY BE USED FOR PART B.

Answer Question 4 below. The Section II score weighting for this question is 15 percent.

4. Write the formulas to show the reactants and the products for any FIVE of the laboratory situations described below. Answers to more than five choices will not be graded. In all cases, a reaction occurs. Assume that solutions are aqueous unless otherwise indicated. Represent substances in solution as ions if the substances are extensively ionized. Omit formulas for any ions or molecules that are unchanged by the reaction. You need not balance the equations.

Example: A strip of magnesium is added to a solution of silver nitrate.

Ex.	$Mg + Ag^+ \rightarrow Mg^{2+} + Ag$

(a) Sulfur dioxide gas is bubbled into distilled water.

(b) A drop of potassium thiocyanate solution is added to a solution of iron(III) nitrate.

(c) A piece of copper wire is placed in a solution of silver nitrate.

(d) Solutions of potassium hydroxide and propanoic acid are mixed.

(e) A solution of iron(II) chloride is added to an acidified solution of sodium dichromate.

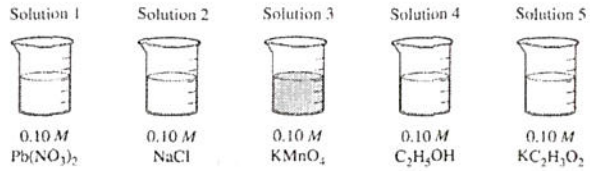
(f) Chlorine gas is bubbled through a solution of potassium bromide.

(g) Solutions of strontium nitrate and sodium sulfate are mixed.

(h) Powdered magnesium carbonate is heated strongly.

Your responses to the rest of the questions in this part of the examination will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

Answer BOTH Question 5 below AND Question 6 printed on page 11. Both of these questions will be graded. The Section II score weighting for these questions is 30 percent (15 percent each).



5. Answer the questions below that relate to the five aqueous solutions at 25°C shown above.

(a) Which solution has the highest boiling point? Explain.

(b) Which solution has the highest pH? Explain.

(c) Identify a pair of the solutions that would produce a precipitate when mixed together. Write the formula of the precipitate.

(d) Which solution could be used to oxidize the $Cl^-(aq)$ ion? Identify the product of the oxidation.

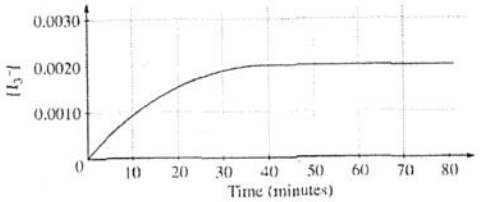
(e) Which solution would be the least effective conductor of electricity? Explain.

$$3\,I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_3^-(aq) + 2\,SO_4^{2-}(aq)$$

6. Iodide ion, $I^-(aq)$, reacts with peroxydisulfate ion, $S_2O_8^{2-}(aq)$, according to the equation above. Assume that the reaction goes to completion.

   (a) Identify the type of reaction (combustion, disproportionation, neutralization, oxidation-reduction, precipitation, etc.) represented by the equation above. Also, give the formula of another substance that could convert $I^-(aq)$ to $I_3^-(aq)$.

   (b) In an experiment, equal volumes of $0.0120\,M$ $I^-(aq)$ and $0.0040\,M$ $S_2O_8^{2-}(aq)$ are mixed at 25°C. The concentration of $I_3^-(aq)$ over the following 80 minutes is shown in the graph below.



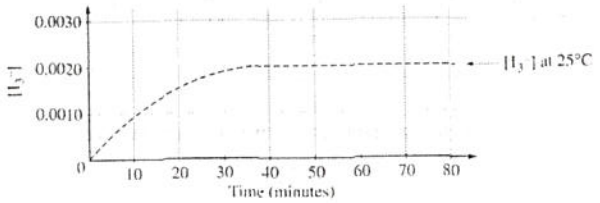
   (i) Indicate the time at which the reaction first reaches completion by marking an "X" on the curve above at the point that corresponds to this time. Explain your reasoning.

   (ii) Explain how to determine the instantaneous rate of formation of $I_3^-(aq)$ at exactly 20 minutes. Draw on the graph above as part of your explanation.

   (c) Describe how to change the conditions of the experiment in part (b) to determine the order of the reaction with respect to $I^-(aq)$ and with respect to $S_2O_8^{2-}(aq)$.

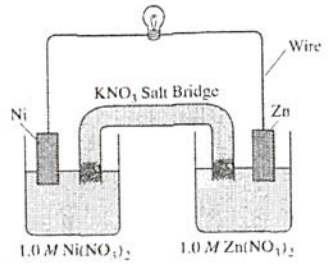
   (d) State clearly how to use the information from the results of the experiments in part (c) to determine the value of the rate constant, k, for the reaction.

   (e) On the graph below (which shows the results of the initial experiment as a dashed curve), draw in a curve for the results you would predict if the initial experiment were to be carried out at 35°C rather than at 25°C.



GO ON TO THE NEXT PAGE.

Answer EITHER Question 7 below OR Question 8 printed on page 13. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 15 percent.



7. Answer the following questions that refer to the galvanic cell shown in the diagram above. (A table of standard reduction potentials is printed on the green insert and on page 4 of the booklet with the pink cover.)

   (a) Identify the anode of the cell and write the half-reaction that occurs there.

   (b) Write the net ionic equation for the overall reaction that occurs as the cell operates and calculate the value of the standard cell potential, E°_{cell}.

   (c) Indicate how the value of E_{cell} would be affected if the concentration of $Ni(NO_3)_2(aq)$ was changed from $1.0\,M$ to $0.10\,M$ and the concentration of $Zn(NO_3)_2(aq)$ remained at $1.0\,M$. Justify your answer.

   (d) Specify whether the value of K_{eq} for the cell reaction is less than 1, greater than 1, or equal to 1. Justify your answer.

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8. Account for each of the following observations about pairs of substances. In your answers, use appropriate principles of chemical bonding and/or intermolecular forces. In each part, your answer must include references to both substances.

   (a) Even though NH_3 and CH_4 have similar molecular masses, NH_3 has a much higher normal boiling point (−33°C) than CH_4 (−164°C).

   (b) At 25°C and 1.0 atm, ethane (C_2H_6) is a gas and hexane (C_6H_{14}) is a liquid.

   (c) Si melts at a much higher temperature (1,410°C) than Cl_2 (−101°C).

   (d) MgO melts at a much higher temperature (2,852°C) than NaF (993°C).

END OF EXAMINATION