The answer key is at the end of this exam.

1. What is the correct expression for $K_c$ for the following reaction?

$$4\text{Fe}(s) + 3\text{O}_2(g) \rightleftharpoons 2\text{Fe}_2\text{O}_3(s) \quad \text{(rust never sleeps)}$$

a. $[\text{Fe}_2\text{O}_3]/\{(\text{[Fe][O}_2]\}$

b. $[\text{Fe}_2\text{O}_3]^2/\{(\text{[Fe]^4}[O_2]^3\}$

c. $1/[O_2]^3$

d. $[O_2]^{1/3}$

e. $\{(\text{[Fe]^4}[O_2]^3\})/[\text{Fe}_2\text{O}_3]^2$

2. If $K_p = 4\times10^{-8}$ for the reaction $3/2\text{O}_2(g) \rightleftharpoons \text{O}_3(g)$, calculate $K_p$ for the reaction:

$$2\text{O}_3(g) \rightleftharpoons 3\text{O}_2(g)$$

a. $2\times10^{16}$

b. $6\times10^{14}$

c. $1\times10^7$

b. $6\times10^{17}$

e. $2\times10^7$

3. Dinitrogen tetroxide in its liquid state was used as one of the fuels in the lunar lander for the NASA Apollo missions. In its gaseous state, it is in equilibrium with nitrogen dioxide:

$$\text{N}_2\text{O}_4(g) = 2\text{NO}_2(g) \quad K_p = 0.133 \text{ at } 450^\circ C$$

Calculate $K_c$ for this equilibrium.

a. $3.60\times10^{-3}$

b. $2.78\times10^2$

c. $0.133$

d. $2.24\times10^{-3}$

e. $7.89$
4. Given the following reactions and $K_c$'s, calculate $K_c$ for: $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$

\[
\begin{align*}
N_2(g) + O_2(g) &= 2NO(g) \quad K_c = 2.7 \times 10^{-18} \\
2NO(g) + O_2(g) &= 2NO_2(g) \quad K_c = 6.0 \times 10^{13}
\end{align*}
\]

a. $2.2 \times 10^{31}$
b. $8.7 \times 10^{13}$
c. $6.2 \times 10^3$
d. $1.6 \times 10^{-4}$
e. $4.5 \times 10^{-32}$

5. A flask contains $NH_4Cl(s)$, $NH_3(g)$, and $HCl(g)$ at equilibrium:

\[NH_4Cl(s) + \text{heat} = NH_3(g) + HCl(g)\]

What will cause the equilibrium to shift towards more products?

a. Cooling the flask.
b. Adding more $NH_4Cl(s)$.
c. Adding an inert gas to the flask.
d. Adding a catalyst.
e. Lowering the pressure in the flask by expansion of its volume.

6. A chemical engineer wants to get maximum yield of hydrogen out of the gas shift reaction:

\[CO(g) + H_2O(g) = CO_2(g) + H_2(g) \quad \Delta H^\circ = -41.2 \text{ kJ}\]

The engineer should ______________ the temperature of the reaction and ______________.
7. Molecular iodine (I₂) is sparingly soluble in water. However, it dissolves readily in solutions containing iodide because of the equilibrium:

I⁻(aq) + I₂(aq) ⇌ I₃⁻(aq) \quad K_c = 710.

A solution at equilibrium contains 0.100 M I⁻ and 0.100 M I₃⁻. What is the molar concentration of I₂?

a. 1.41×10⁻³ M  
b. 71.0 M  
c. 0.710 M  
d. 1.41×10⁻⁴ M  
e. 1.41×10⁻⁵ M

8. A mixture of 4.2 mol of N₂, 2.0 mol H₂, and 10.0 mol NH₃ is introduced into a 20.0 L reaction vessel at 500 K. At this temperature, K_c for the synthesis of ammonia is 1.7×10².

N₂(g) + 3H₂(g) ⇌ 2NH₃(g)

What is the correct conclusion about this mixture?

a. Q = K_c; the mixture is at equilibrium.  
b. Q < K_c; more NH₃ will form.  
c. Q < K_c; more N₂ and H₂ will form.  
d. Q > K_c; more NH₃ will form.  
e. Q > K_c; more N₂ and H₂ will form.

9. 1.0 mol of CO₂ is placed into a 1.0 L flask and heated to 3000K. At equilibrium, 40% of the CO₂ decomposes to CO and O₂:

2CO₂(g) ⇌ 2CO(g) + O₂(g)

Calculate the value of K_c.

a. 0.40  
b. 0.089  
c. 0.11  
d. 0.13  
e. 0.027
10. 3.0 mol of IBr are placed in a 1.0 L flask and heated to 400 K. Calculate the concentration of I₂ at equilibrium.

\[ 2\text{IBr}(g) \rightleftharpoons \text{I}_2(g) + \text{Br}_2(g) \quad K_c = 0.16 \]

a. 1.5 M  
b. 0.86 M  
c. 0.36 M  
d. 0.67 M  
e. 0.48 M

11. The conjugate acid of CO₃²⁻ is ___________ while the conjugate base of NH₄⁺ is ___________.

a. HCO₃⁻, NH₃  
b. H₂CO₃, NH₂⁻  
c. HCO₃⁻, NH₄⁺  
d. CO₂⁺, NH₃  
e. HCO₃⁻, NH₃⁺

12. In the following reaction, pick out the two conjugate acid/base pairs:

\[ \text{HSO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{SO}_3 + \text{OH}^- \]

a. HSO₃⁻/H₂O and H₂SO₃/OH⁻  
b. H₂SO₃/HSO₃⁻ and H₂O/OH⁻  
c. HSO₃⁻/OH⁻ and H₂SO₃/H₂O

13. In the following reaction, pick out the Lewis base:

\[ \text{BCl}_3 + \text{NH}_3 \rightleftharpoons \text{Cl}_3\text{B}⁻\text{NH}_3 \]

a. BCl₃  
b. NH₃  
c. Cl₃B⁻NH₃
14. Which of the following ions are not basic i.e. do not react with water to form OH⁻?

a. OCl⁻

b. PO₄³⁻

c. S²⁻

d. CO₃²⁻

e. I⁻

15. A solution of KNO₃ is __________ while a solution of Ba(CH₃COO)₂ is __________.

a. neutral, acidic

b. acidic, basic

c. acidic, neutral

d. basic, neutral

e. neutral, basic

16. The concentration of OH⁻ in a sample of water from a limestone spring at 25°C is 1.3×10⁻⁶ M. The concentration of H⁺ is __________; the water is __________.

a. 7.7×10⁻⁹ M; basic

b. 8.7×10⁻⁶ M; acidic

c. 1.0×10⁻⁷ M; neutral

d. 7.7×10⁻⁶ M, basic

e. 7.7×10⁻⁹ M; acidic

17. Vinegar has a pH of 2.95. What is the concentration of H⁺ in vinegar?

a. 11.05 M

b. 8.9×10⁻¹² M

c. 1.1×10⁻³ M

d. 5.2×10⁻² M

e. 8.9×10⁴ M
18. The pH of household ammonia solution is 11.80. What is $[\text{OH}^-]$?
   a. $8.5 \times 10^{-2}$ M
   b. $1.6 \times 10^{-12}$ M
   c. $1.2 \times 10^{-13}$ M
   d. $6.3 \times 10^{-3}$ M
   e. $7.5 \times 10^{-6}$ M

19. What is the pH of a 5.5 M HBr solution?
   a. 0.74
   b. 1.70
   c. −0.74
   d. 2.50
   e. 8.50

20. What is the pH of a 0.15 M Ba(OH)$_2$ solution?
   a. 13.48
   b. 13.70
   c. 13.17
   d. 13.85
   e. 14.15

**ANSWERS**

1. c. Neither Fe nor Fe$_2$O$_3$ appear in the expression because they are pure solids. O$_2$ appears in the denominator because it is a reactant. “Rust never sleeps” is a quote to remind scientists and engineers of the overwhelming importance of iron and steel oxidation in an industrialized society.

2. b. Invert the reaction ($O_3 \rightarrow 3/2O_2$) and multiply by 2 ($2O_3 \rightarrow 3O_2$). $K'_p = (1/K_p)^2 = (1/4 \times 10^{-8})^2 = 6 \times 10^{14}$.

3. d. $K_p = K_c(RT)^\Delta n$. $\Delta n = 2 - 1 = +1$. $T = 450 + 273 = 723$ K. $0.133 = K_c((0.08206)(723))^{+1}$; $K_c = 2.24 \times 10^{-3}$.

4. d. The final reaction is the sum of the two given reactions. $K'_c = K_{c_1}K_{c_2} = (2.7 \times 10^{-18})(6.0 \times 10^{13}) = 1.6 \times 10^{-4}$.

5. e. (a) shifts the equilibrium towards reactants; (b)-(d) have no effect on the equilibrium. Lowering the pressure shifts the equilibrium towards the side with the greater number of moles of gas, i.e., the products.

6. e. Raising the temperature of an exothermic equilibrium shifts the equilibrium towards reactants (heat is a product). Changing the pressure has no effect because there are 2
moles of gas on both sides of the equilibrium. Removing CO₂ and lowering the
temperature is the best way to get more H₂.

7. a. A type 1 problem. \( K_c = [I_2]/([I][I_2]) = 710 = (0.100)/((0.100)[I_2]). \) \( [I_2] = 1/710 = 1.41 \times 10^{-3} \) M.

8. e. \( Q = [NH_3]^2/([N_2][H_2]^3). \) \( [NH_3] = (10.0 \text{ mol})/(20. \text{ L}) = 0.50 \) M; \( [N_2] = (4.2 \text{ mol})/(20. \text{ L}) = 0.21 \) M; \( [H_2] = (2.0 \text{ mol})/(20. \text{ L}) = 0.10 \) M. \( Q = (0.50)^2/(0.21)(0.10)^3) = 1.2 \times 10^4. \) Since \( Q > K_c, \) the equilibrium shifts towards reactants.

9. b. \( 2CO_2 \rightleftharpoons 2CO + O_2 \)

\[
\begin{array}{ccc}
\text{Init.} & 1.0 \text{ M} & 0 & 0 \\
\Delta & -0.40 \text{ M} & +0.40 \text{ M} & +0.20 \text{ M} \\
\text{Final} & 0.60 \text{ M} & 0.40 \text{ M} & 0.20 \text{ M} \\
\end{array}
\]

\( \Delta = \text{eq.} \) (follows stoichiometry)

\( K_c = \frac{[CO]^2[O_2]}/[CO_2]^2 = \frac{(0.40)(0.20)}{(0.60)^2} = 0.089 \)

10. d. \( 2IBr \rightleftharpoons I_2 + Br_2 \)

\[
\begin{array}{ccc}
\text{Init.} & 3.0 \text{ M} & 0 & 0 \\
\Delta & -2x & +x & +x \\
\text{Final} & 3.0 - 2x & x & x \\
\end{array}
\]

\( K_c = \frac{[I_2][Br_2]}/[IBr]^2 = 0.16 = x^2/(3.0 - 2x) \) (a perfect square; take square root)

\( 0.40 = x/(3.0 - 2x); \) \( x = [I_2] = 0.67 \) M.

11. a. Add H⁺ to get the conjugate acid; remove H⁺ to get the conjugate base.

12. b. The acid contains the extra H⁺, the base does not.

13. b. NH₃ has a lone pair of non-bonding electrons which it donates to the boron, which has only 3 pairs of electrons, not the usual 4.

14. e. Ions that are not basic are either neutral or acidic. None of the listed ions can be acidic (no H⁺). Neutral anions are the conjugate bases of strong acids. The list of strong acids is: HCl, HBr, HI, HNO₃, HClO₃, and H₂SO₄.

15. e. See above (Q. 14). Neutral cations are the metal cations of Group I and Group II on the Periodic Table. K⁺ is in Group I, Ba²⁺ is in Group II. NO₃⁻ is neutral, CH₃COO⁻ is basic.

16. a. \( [H^+] = K_w/[OH^-] = 1.0 \times 10^{-14}/1.3 \times 10^{-6} = 7.7 \times 10^{-9} \) M. Since \( [H^+] \) is less than \( 1.0 \times 10^{-7} \) M, the solution is basic.

17. c. \( [H^+] = 10^{-2.95} = 1.1 \times 10^{-3} \) M.

18. d. \( \text{pOH} = 14.00 - \text{pH} = 14.00 - 11.80 = 2.20. \) \( [OH^-] = 10^{-2.20} = 6.3 \times 10^{-3} \) M.

19. c. \( \text{pH} = -\log(5.5) = -0.74. \) A negative pH simply means that \( [H^+] \) is greater than 1 M.

20. a. Since Ba(OH)₂ ---→ Ba²⁺ + 2OH⁻, \( [OH^-] = 2(0.15 \text{ M}) = 0.30 \) M. \( \text{pOH} = -\log(0.30) = 0.52. \) \( \text{pH} = 14.00 - 0.52 = 13.48. \)