Use the following information to answer the next two questions.

The following chemical equation represents a reaction that occurs when soft drinks are carbonated.

\[ \text{CO}_2(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq) \]

1. Which statement describes a carbonated beverage at equilibrium (in a soft drink bottle)?
   a. The contents are in a closed system and have variable macroscopic properties.
   b. The contents are in an open system and have variable macroscopic properties.
   c. The contents are in a closed system and have constant macroscopic properties.
   d. The contents are in an open system and have constant macroscopic properties.

2. What occurs when this reaction is at equilibrium?
   a. The decomposition of H_2CO_3(aq) begins.
   b. The formation of H_2CO_3(aq) stops.
   c. The decomposition of H_2CO_3(aq) stops.
   d. The rates of formation and decomposition of H_2CO_3(aq) are equal.

3. The quantities of reactants and products in chemical systems can be related as follows.
   I. More reactants than products
   II. More products than reactants
   III. Equal reactants and products

Which relationship(s) can be true for chemical systems that are at equilibrium?

a. III only
b. II only
c. I, II, and III
d. I and II

4. A beaker contains a saturated aqueous solution of lead (II) chromate in contact with the solid compound. The system is represented by the chemical equation:

\[ \text{PbCrO}_4(s) \rightleftharpoons \text{Pb}^{2+}(aq) + \text{CrO}_4^{2-}(aq) \quad \text{K}_{eq} = 2 \times 10^{-16} \text{ at } 25^\circ\text{C} \]

At 25°C, the solubility of lead (II) chromate is:
   a. moderate
   b. very high
   c. impossible to determine from the information given
   d. very low
5. Chemistry teachers often use the reaction between copper metal and aqueous silver nitrate to demonstrate a single replacement reaction. The net ionic equation for this reaction follows:

\[ \text{Cu(s)} + 2 \text{Ag}^{+}\text{(aq)} \leftrightarrow \text{Cu}^{2+}\text{(aq)} + 2 \text{Ag(s)} \]

The equilibrium constant for the reaction between silver metal and the copper ion at 25°C is

a. 1 X 10^{-14}  

b. 2 X 10^{15}  

c. 5 X 10^{-16}  

d. 2 X 10^{15}  

6. Use the following information to answer the next two questions.

The reddish brown haze that sometimes hangs over urban areas is caused by nitrogen dioxide, a pollutant produced by automobile engines. Nitrogen dioxide exists in equilibrium with dinitrogen tetraoxide as shown in the chemical equation that follows:

\[ \text{N}_2\text{O}_4\text{(g)} + \text{energy} \leftrightarrow 2\text{NO}_2\text{(g)} \]

colourless \leftrightarrow reddish brown

6. In a study of the air pollutant mentioned above, an environmental chemist finds that, at 50°C, a closed flask contains 2.00 mmol/L of NO_2(g) and 5.00 mmol/L of N_2O_4(g). The numerical value for the equilibrium constant for this reaction at 50°C is

a. 0.800  

b. 0.400  

c. 2.50  

d. 1.25  

7. Which conditions will lower the equilibrium constant for the system above?

a. Lower pressure  

b. Lower N_2O_4\text{(g)} concentration  

c. Lower temperature  

d. Addition of a catalyst
Use the following information to answer the next two questions.

The reaction of nitrogen with oxygen in the combustion chamber of an automobile engine produces a variety of toxic oxides of nitrogen. One such reaction can be represented by the equation that follows:

\[ 2 \text{N}_2(g) + \text{O}_2(g) \leftrightarrow 2 \text{N}_2\text{O} (g) \]

At a very high temperature, nitrogen at an initial concentration of 1.20 mol/L is mixed with oxygen at an initial concentration of 1.00 mol/L. This reaction mixture reaches equilibrium when the concentration of the produced dinitrogen oxide, commonly called nitrous oxide, is 1.00 mol/L.

8. For the reaction described above, calculate the numerical value of the equilibrium constant.
   a. 50
   b. 75
   c. $1.0 \times 10^2$
   d. 10

9. For the reaction described above, calculate the equilibrium concentration of \( \text{O}_2(g) \).
   a. 0.50 mol/L
   b. 0.20 mol/L
   c. 1.20 mol/L
   d. 1.00 mol/L

Use the following information to answer the next question.

Propane, \( \text{C}_3\text{H}_8(g) \), a component of natural gas may be cracked thermally to produce ethylene, \( \text{C}_2\text{H}_4(g) \). Propane is introduced into a reaction chamber at 325°C and no catalyst.

\[ \text{C}_3\text{H}_8(g) \leftrightarrow \text{C}_2\text{H}_4(g) + \text{CH}_4(g) \]

When equilibrium is established the concentration of reactant and products are as follows:

\[ [\text{C}_3\text{H}_8(g)] = 0.206 \text{ mol/L} \]
\[ [\text{C}_2\text{H}_4(g)] = [\text{CH}_4(g)] = 0.394 \text{ mol/L} \]

10. The percent yield of the \( \text{C}_2\text{H}_4(g) \) in this system is:
   a. 32.8 %
   b. 52.3 %
   c. 65.7 %
   d. 19.7 %
11. The synthesis of ammonia is achieved world-wide through the Haber-Bosch process. Consider the following research evidence gathered for the synthesis of ammonia:

\[ \text{N}_2(g) + 3 \text{H}_2(g) \leftrightarrow 2 \text{NH}_3(g) \]

The synthesis reaction of ammonia from nitrogen and hydrogen is:
- a. increased by increasing the temperature
- b. decreased by increasing the pressure
- c. endothermic
- d. exothermic

12. Hydrofluoric acid is used to etch glass. A 100 mmol amount of hydrogen fluoride gas is dissolved in water to make 1.0 L of an etching solution. At 20°C, the percent reaction at equilibrium is 6.0%. The chemical equilibrium system is:

\[ \text{HF}(\text{aq}) \leftrightarrow \text{H}^+(\text{aq}) + \text{F}^-(\text{aq}) \]

According to this evidence, the concentration of:
- a. hydrogen ions is 94 mmol/L
- b. fluoride ions is 6.0 mmol/L
- c. hydrofluoric acid is 100 mmol/L
- d. hydrogen fluoride is 6.0 mmol/L

Use the following information to answer the next question.

<table>
<thead>
<tr>
<th>Temperature (K)</th>
<th>Equilibrium Constant (L²/mol²)</th>
</tr>
</thead>
<tbody>
<tr>
<td>500</td>
<td>90</td>
</tr>
<tr>
<td>600</td>
<td>3</td>
</tr>
<tr>
<td>700</td>
<td>0.3</td>
</tr>
<tr>
<td>800</td>
<td>0.04</td>
</tr>
</tbody>
</table>

According to this evidence, the synthesis reaction of ammonia from nitrogen and hydrogen is:
- a. increased by increasing the temperature
- b. decreased by increasing the pressure
- c. endothermic
- d. exothermic

13. If at equilibrium, \([ \text{CO}(g) ] = 0.25 \text{ mol/L}, [ \text{H}_2(g) ] = 0.12 \text{ mol/L} \) and \([ \text{CH}_3\text{OH}(g) ] = 0.0018 \text{ mol/L} \), the value of the equilibrium constant is:
- a. 0.50
- b. 0.060
- c. 17
- d. 2.0

14. The decomposition of hydrogen peroxide to produce oxygen and water is used in some contact lens cleaning procedures. A small piece of platinum is used as a catalyst for this reaction. Assuming a closed system, the purpose of the platinum is to:
- a. Shift the equilibrium towards the reactants
- b. Alter the equilibrium concentrations
- c. Decrease the time required to reach equilibrium
- d. Shift the equilibrium towards the products
15. A catalytic converter in a car exhaust system increases the rate of many chemical reactions in the exhaust gases. If the catalytic converter makes the reaction of carbon monoxide with oxygen to produce carbon dioxide happen much more rapidly, then it is correct to imply that the:
   a. CO\(_2\)(g) concentration in the exhaust gases will decrease greatly
   b. CO\(_g\) and CO\(_2\)(g) reaction system must reach equilibrium
   c. CO\(_g\) concentration is much less than it would be if no converter were used
   d. Equilibrium constant is increased considerably from its normal value.

16. Consider the following hypothetical equilibrium system:
   \[ QA\,(s) + 2\,X\,(g) \rightleftharpoons Q\,(l) + X_2A\,(g) + \text{energy} \]
   This reaction can be shifted to increase the temperature by adding:
   a. X\,(g)
   b. QA\,(s)
   c. X\(_2\)A\,(g)
   d. Q\,(l)

17. Which of the following variables are singly or collectively manipulated in order to use Le Chatelier’s Principle to predict shifts in chemical equilibria?
   a. Pressure, concentration and addition of a catalyst
   b. Temperature, concentration and addition of a catalyst
   c. Temperature, pressure and addition of a catalyst
   d. Temperature, pressure and concentration

Use the following information to answer the next question.

The control of blood pH is important to the transport and delivery of oxygen in the bloodstream. Some important reactions that occur in the circulatory system are:

I. \[ \text{HHb\,(aq)} + \text{O}_2\,(g) \rightleftharpoons \text{HHbO}_2\,(aq) \]
   hemoglobin            oxyhemoglobin

II. \[ \text{HHbO}_2\,(aq) + \text{HCO}_3^-\,(aq) \rightleftharpoons \text{H}_2\text{CO}_3\,(aq) + \text{HbO}_2^-\,(aq) \]

III. \[ \text{H}_2\text{CO}_3\,(aq) \rightleftharpoons \text{CO}_2\,(g) + \text{H}_2\text{O\,(l)} \]

18. When there is a high concentration of O\(_2\)(g) in the lungs, the effect would be to shift:
   a. Reaction I right and Reaction II left
   b. Reaction I right and Reaction III left
   c. All three reactions left
   d. All three reactions right
Large quantities of ammonia are produced by the Haber-Bosh method. The essential reaction in this process involves the equilibrium:

\[ N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g) + 92.2 \text{ kJ} \]

19. The equation that best represents the equilibrium in an aqueous solution of ammonia, producing a basic solution is:
   a. \[ \text{NH}_3(aq) \leftrightarrow 3H^+(aq) + N_3^-(aq) \]
   b. \[ \text{NH}_3(aq) + H_2O(l) \leftrightarrow \text{NH}_4OH(aq) \]
   c. \[ \text{NH}_3(aq) + H_2O(l) \leftrightarrow \text{NH}_4^+(aq) + OH^-(aq) \]
   d. \[ \text{NH}_3(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + \text{NH}_2^- (aq) \]

20. The conditions that theoretically favour the formation of ammonia in the Haber-Bosh process are
   a. High pressure and high temperature
   b. High pressure and low temperature
   c. Low pressure and high temperature
   d. Low pressure and low temperature

Use the following additional information to answer the next two questions.

Data for the equilibrium system used in the Haber-Bosh process was collected and plotted on a graph by a student:

**NR #1:** The minimum time, in minutes, required to establish equilibrium is ______ min.  
(Record your **three-digit** answer in the space provided)

**NR #2:** The equilibrium constant for this system is ______.  
(Record your **three-digit** answer in the space provided)
Air bags are designed to deploy when a car moving in excess of 18 km/h undergoes a sudden decrease in speed. During the deceleration, an electrical circuit ignites a small amount of explosive powder to cause the sodium azide, NaN₃(s), to react inside the air bag. The sodium azide rapidly decomposes, producing nitrogen gas, which fills the bag. The equation for this reaction is:

\[
2\text{NaN}_3(s) \rightarrow 3\text{N}_2(g) + 2\text{Na}(s) + 43.5 \text{ kJ}
\]

21. One problem associated with inflated airbags is that a white powder is produced. It can cause skin and eye irritation. The irritation is most likely due to the sodium metal reacting with moisture in the surroundings to form:
   a. An acid
   b. A base
   c. A salt
   d. An inert, gritty material

22. A 0.100 mol/L unknown acid solution is found to have a \([\text{H}_3\text{O}^+_{(aq)}]\) of 0.0015 mol/L. The pH of this solution is
   a. 1.0
   b. 1.00
   c. 2.8
   d. 2.82

23. 2.0 g of Mg(OH)₂(s) is added to 500 mL of distilled water. Calculate the pH of this solution.
   a. 13.14
   b. 0.86
   c. 12.83
   d. 1.17