You may need a calculator for number 4 (but you actually don't as long as you're clever about how you approach the problem).

1. You're doing an experiment to see how long an <u>open</u> container with 100mL of methanol will last. You also have a meter handy to tell you the amount of methanol left (and once it is gone), in case you decide to store it as a mixture. To decrease the vapor pressure, and thus maximize how long the methanol will last, you could:

I. Keep it on your sunny windowsill
II. Conduct the experiment in your cabin at the mountains
III. Decrease the surface area of the container
IV. Add water to it
V. Store it in a cool place
VI. Pour in some CaO
VII. Get some surfing time in between measurements (be at sea level)

1) III, IV, V and VII
2) IV and V
3) III, V and VI
4) II and III
5) II, IV and VI

2. A 7.0 molal aqueous solution of (CsF/ methanol/SrI_2) would have a higher boiling point. What is the boiling point of this solution? K_{bwater} = 0.5 K*kg*mol^{-1}.

1) CsF, 380K
2) CH_3OH, 376.5K
3) CsF, 383.5K
4) SrI_2, 380K
5) SrI_2, 383.5K
6) CH_3OH, 383.5K
7) CsF, 363K

3. Sodium hypochlorite (NaOCl), aka bleach, is commercially prepared by adding chlorine gas to a solution of sodium hydroxide.

$Cl_2(aq) + 2\ OH^-(aq)$ □ $ClO^-(aq) + H_2O(l) + Cl^-(aq)$

The hypochlorite ion is the active bleaching agent, but can decompose to chloride and chlorate ions in a competing side reaction:

$3\ ClO^-(aq)$ □ $2\ Cl^-(aq) + ClO_3^-(aq)$

Set up the equilibrium constant K for the decomposition of hypochlorite:

1. $K = [Cl^-]^2[ClO^-]^3/[ClO_3^-]$

2. $K = [Cl^-]^2[ClO_3^-]/[ClO^-]^3$

3. $K = [ClO^-]^3 /[Cl^-]^2[ClO_3^-]$

4. $K = [Cl^-][ClO_3^-]/[ClO^-]$

5. $K = [Cl^-][ClO^-]/[Cl_2][OH^-]^2$

4. Suppose H_2 (g) and I_2 (g) are sealed in a container at T= 400 K with partial pressures P_{H2} = 1.32 atm and P_{I2} = 1.14 atm. At this temperature, the gases do not react rapidly to form HI (g), although after a long enough time they would produce HI (g) at its equilibrium partial pressure. Suppose, instead that the gases are heated in the sealed flask to 600 K, a temperature at which equilibrium is quickly established:

$$H_2(g) + I_2(g) \square \; 2\,HI(g)$$

For this reaction at 600K, the equilibrium constant K = 92.6.

Determine the **change in partial pressure** of hydrogen gas.

1. 1.5044 atm
2. 2.3522 atm

5. Consider the following two ways to make the product, C:

I. $A + B \square C \quad K = 4\ M^{-1}$

II. $2A + B \square C \quad K = 2\ M^{-2}$

All other things being equal, at what concentration of A would method I produce more C?

1. [A] > 1M
2. [A] < 1 M
3. All values of [A]
4. No values of [A]

6. Which of the following statements is/are true:
I. K will slowly approach the value of Q while the reaction reaches equilibrium.
II. If Q>K, the reverse reaction is dominant over the forward reaction.
III. Q will never exceed the value of K.

1. I only
2. II only
3. III only
4. I and II
5. I and III

7. Which of the followings statements is not true concerning the Le Chatelier Principle?
1. For the reaction $N_2(g) + 3\,H_2(g) \square\ 2\,NH_3(g)$ (called the Haber process), decreasing the volume of the system would cause ammonia to form.
2. For the reaction $CO(g) + NO_2(g) \square\ CO_2(g) + NO(g)$, changing the volume of the system would have no effect.

3. For the reaction $CO(g) + 2\ H_2(g)$ ☐ $CH_3OH(g)$ adding methanol would cause carbon monoxide to form.

4. For the reaction $N_2(g) + 3\ H_2(g)$ ☐ $2\ NH_3(g)$ ($\Delta H = -92\ kJ{\cdot}mol^{-1}$), increasing the temperature would cause ammonia to form.

8. If a given reaction has an equilibrium constant of 10 at 25 °C (K = 10, T = 302 K) and a standard change in enthalpy of $-10\ kJ{\cdot}mol^{-1}$ ($\Delta H° = -10\ kJ{\cdot}mol^{-1}$), without doing any math, at which of the following temperatures will the reaction most likely be non-spontaneous?

1. 100 °C
2. - 200 °C
3. 500 K
4. 250 K