1. A gaseous chemical equilibrium has an equilibrium constant with the following form.

$$K_p = \frac{P_{HI}^2}{P_{H_2}P_{I_2}}$$

A) Write a balanced chemical equation for this system/reaction.

$H_2\,(g) + I_2\,(g) \rightleftharpoons 2HI\,(g)$

B) Write an expression for K_c and determine the relationship between K_p and K_c.

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

Use ideal gas law: $\Delta n = 0$

$$[x] = \frac{n_x}{V} = \frac{P_x}{RT} \rightarrow K_c = \frac{[HI]^2}{[H_2][I_2]} = \frac{P_{HI}^2}{P_{H_2}P_{I_2}} \times \left(\frac{1}{RT}\right)^{\Delta n} = K_p$$

C) A container holds $[H_2] = 2.95 \times 10^{-3}$ M, $[I_2] = 5.22 \times 10^{-4}$ M, and $[HI] = 1.95 \times 10^{-3}$ M at 25 °C. If $K_c = 48.8$ at 25 °C, in which direction will the reaction proceed in the container?

Determine the reaction quotient (Q):

$$Q = \frac{[HI]^2}{[H_2][I_2]} = \frac{(1.95 \times 10^{-3})^2}{(2.95 \times 10^{-3})(5.22 \times 10^{-4})} = 2.47$$

Because Q < K, the reaction will shift toward the right or products side (HI).

2. In the laboratory you synthesize emerald-colored crystals of trihydrate potassium ferrioxalate ($K_3[Fe(C_2O_4)_3]\cdot 3H_2O$) from aqueous solutions of $FeCl_3$ and $K_2C_2O_4$.

$$FeCl_3\,(aq) + 3K_2C_2O_4\,(aq) + 3H_2O\,(l) \rightarrow K_3[Fe(C_2O_4)_3]\cdot 3H_2O\,(s) + KCl\,(aq)$$

Recrystallization from a saturated aqueous solution of your products, contaminated with by-products and starting materials, served to purify your desired product.

A) Write a (solubility) equilibrium constant for dissolution of the crystals:

$$K_3[Fe(C_2O_4)_3]\cdot 3H_2O\,(s) \rightleftharpoons 3K^+\,(aq) + [Fe(C_2O_4)_3]^{3-}\,(aq) + 3H_2O\,(l)$$

$$K_{sp} = [K^+]^3[\{Fe(C_2O_4)_3\}^{3-}]$$

B) If cooling the saturated solution results in crystal formation, is the dissolution of the $K_3[Fe(C_2O_4)_3]\cdot 3H_2O$ and endothermic or exothermic process?

If cooling shifts the equilibrium to the left, then heat must be a "reactant."
Therefore, this reaction is endothermic: $\Delta H > 0$.

3. Consider the reaction between phosphorus(III) chloride and chlorine gas to produce phosphorus(V) chloride.

$$PCl_3\ (g) + Cl_2\ (g) \rightleftharpoons PCl_5\ (g) \qquad K_p = 24.2$$

A) A 1.00 L container at constant temperature contains $P_{PCl_3} = 1.5$ atm, $P_{Cl_2} = 0.72$ atm, and $P_{PCl_5} = 0$ atm initially. Calculate the partial pressures at equilibrium.

Because no PCl_5 is present initially ($Q < K_p$), equilibrium shifts to the right.

Set up an ICE chart (units of atm):

	PCl_3 (g)	+	Cl_2 (g)	⇌	PCl_5 (g)
I	1.5		0.72		0
C	– x		– x		+ x
E	1.5 – x		0.72 – x		x

Now set up the equilibrium expression and solve for x:

$$K_p = \frac{P_{PCl_5}}{P_{PCl_3}P_{Cl_2}}$$

$$24.2 = \frac{x}{(1.5 - x)(0.72 - x)}$$

$$0 = 24.2x^2 - 54.724x + 26.136$$

$$x = 1.5_8 \ or\ x = 0.68_5$$

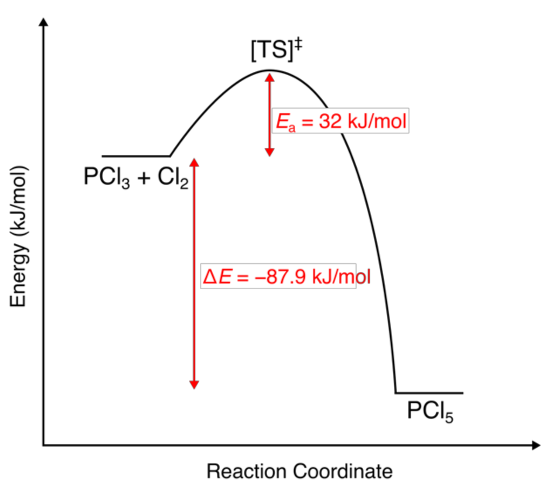
Discard the x = 1.5_8 solution, so the equilibrium partial pressures are:

$$P_{PCl_5} = 0.69 \text{ atm} \quad P_{PCl_3} = 0.8 \text{ atm} \quad P_{Cl_2} = 0.03 \text{ atm}$$

B) Describe some ways in which we can increase the yield of PCl_5 (g).

- Add either of the reactants → shifts right
- Remove products → shifts right
- Increase the pressure → shifts right
- Decrease the volume → shifts right
- Decrease the temperature → see part C

C) The energy diagram for the reaction is shown below. Determine how the equilibrium number of moles of PCl_5 would change if system were heated.



The reaction is exothermic ($\Delta H < 0$), so we can treat heat as a "product." Therefore, increasing the temperature would shift the reaction to the left and the number of moles of PCl_5 would decrease.