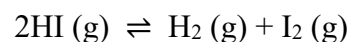


Chapter 14: Equilibrium
Practice Problem Set

Given reaction:



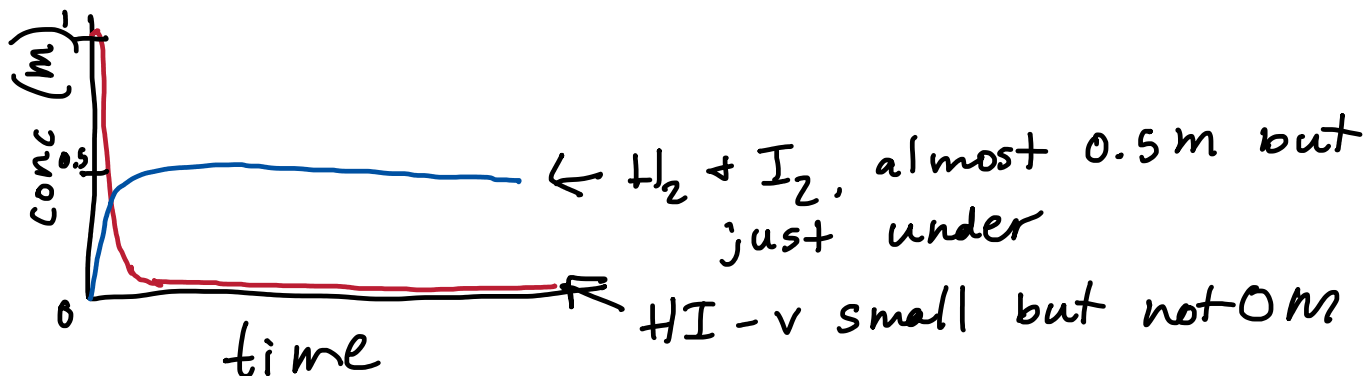
- a. A system of HI, H₂, and I₂ was left to equilibrate at 445 °C. Once equilibrium was established, the concentrations of HI, H₂, and I₂ were determined to be 2.506 x 10⁻⁴ M, 1.24 x 10⁻³ M, and 3.24 x 10⁻³ M, respectively. Calculate K_c at this temperature.

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{[1.24 \times 10^{-3} \text{ M}][3.24 \times 10^{-3} \text{ M}]}{[2.506 \times 10^{-4} \text{ M}]^2} = 64.0$$

- b. What is K_p at this temperature?

Same as K_c, since $K_p = K_c(RT)^{\Delta n}$,
where $\Delta n = \text{moles gaseous prod} - \text{moles gaseous reactants}$
and here, $\Delta n = 0$

- c. Draw a plot of the concentration of the reactants and products over time, assuming we began with 1M of HI.



- d. If the partial pressures of HI and H₂ are found to be 3.01×10^{-3} atm and 2.04×10^{-2} atm at equilibrium, respectively, what would the partial pressure of I₂ be?

$$K_c = K_p = 64 = \frac{[2.04 \times 10^{-2} \text{ atm}] x \text{ atm}}{(3.01 \times 10^{-3} \text{ atm})^2} =$$

$$x = 0.0284 \text{ or } \boxed{2.84 \times 10^{-2} \text{ atm}}$$

- e. Calculate Q when $[HI] = 2.11 \times 10^{-4}$ M and $[H_2] = 2.10 \times 10^{-3}$ M and $[I_2] = 3.4 \times 10^{-3}$ M. Is the system at equilibrium? If no, in which direction is the reaction predominately proceeding?

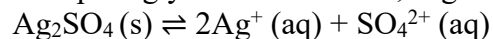
$$Q = \frac{[H_2][I_2]}{[HI]^2} = \frac{[2.10 \times 10^{-3} \text{ M}][3.4 \times 10^{-3} \text{ M}]}{[2.11 \times 10^{-4} \text{ M}]^2} = 160.4$$

$Q > K$ indicating too many products,
Rxn is proceeding towards reactants.

- f. How would the system respond if the volume of the reaction flask was decreased without changing the amount of reactants added?

No change, since $\Delta n = 0$ as above.

The dissolution of Ag_2SO_4 , which is known to be sparingly soluble in water, is given by:



$$K_{sp} = 1.20 \times 10^{-5}$$

What will the concentration of SO_4^{2-} and Ag^+ be at equilibrium when 0.342 g of Ag_2SO_4 is added to 10 L of water?

↳ This is superfluous info.
The conc. @ eq. will always be dictated only by K_{sp} .

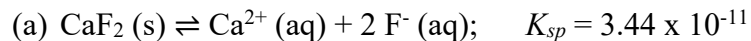
$$K_{sp} = 1.20 \times 10^{-5} = [\text{Ag}^+]^2 [\text{SO}_4^{2-}] = [2x]^2 [x] = 4x^3 \rightarrow x = 0.0144$$

↑
For every 1 unit of Ag_2SO_4 that dissolves, you get 2 Ag^+ + 1 SO_4^{2-}

So:

$[\text{Ag}^+] = 2x = 0.0288 \text{ M}$
$[\text{SO}_4^{2-}] = x = 0.0144 \text{ M}$

For the following scenarios, indicate which direction the equilibrium will shift when the "stressor" indicated is introduced?

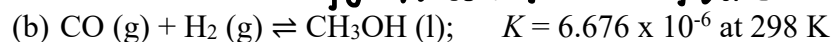


(i) Additional water is added to a system at equilibrium

Right - more CaF_2 will dissolve to reestablish conc. of $\text{Ca}^{2+} + \text{F}^-$.

(ii) $\text{Ca}(\text{NO}_3)_2$ is added

Left - K_{sp} dictates eq. $[\text{Ca}^{2+}]$, & thus some CaF_2 will reform to maintain this conc.



(i) The volume of the reaction chamber is compressed

Right - more liquid product will be produced to relieve the pressure

(ii) More H_2 is added

Right - to use up excess H_2 + reestablish eq. concs of reactants

(c) $NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$; $K_b = 1.8 \times 10^{-5}$

(i) The solution is basified with NaOH

Left - to use up excess OH^- , also NH_3 should predominate at a higher pH.

(ii) The solution is acidified with HCl

Right - H^+ will neutralize some OH^- , and thus reaction will re-equilibrate to generate more OH^- also, NH_4^+ should predominate at lower pHs.

(d) $MgSO_4(s) \rightleftharpoons Mg^{2+}(aq) + SO_4^{2-}(aq)$; $K_{sp} = 5.9 \times 10^{-3}$; $\Delta H = -91.38 \text{ kJ/mol}$

(i) The solution is heated

Left - the reaction is exothermic + can be treated as if heat is a product

(ii) The solution is cooled

Right for the reason above

(iii) Solid $BaSO_4$ is added to the reaction flask

No change - $BaSO_4$ is insoluble in water + thus will produce no common ion to perturb the equilibrium.