Answer on Question #65027 - Chemistry - General Chemistry

Question: Calculate the equilibrium concentration of each compound if 0.200 moles of H_2 and 0.200 moles of H_2S are placed in a 2.50 liter container at 200°. K_c = 14.5 for the reaction at this temperature. Assume an excess of sulfur is present.

Solution

1) Sulfur which is present in excess is in the solid state, so it is not involved into the expression of the equilibrium constant, which involves only gaseous components. The balanced equation for this reaction is

$$H_2(gas) + S(solid) = H_2S(gas)$$

So, the total amount of substance of gaseous components does not change during the reaction (hydrogen which reacts is transformed into the equivalent amount of matter of gaseous hydrogen sulfide), as well as the pressure during the reaction. Thereof, the expression for the equilibrium constant for this reaction can be written as following:

$$K_c = \frac{[H_2S]}{[H_2]} = 14.5$$

2) Find the total concentration of the gaseous components in the container. At the starting point, 0.2 moles of H_2S and 0.2 moles of H_2 were placed into the container. As we already mentioned, the total amount of substance of gaseous compounds does not change during the reaction, as any amount of hydrogen is converted into the same amount of hydrogen sulfide. So, the total amount of gaseous substances in the container will stay constant and equivalent to the amount of H_2 and H_2S initially added:

$$n(gases) = n_0(H_2) + n_0(H_2S) = 0.2\,mol + 0.2\,mol = 0.4\,mol$$

The total equilibrium concentration of all gases will be

$$[gas] = \frac{n(gases)}{V} = \frac{0.4\,mol}{2.5\,l} = 0.16\frac{mol}{l}.$$

3) From the written above equations for the equilibrium concentration of gas and K_c we can derive:

$$[H_2S] + [H_2] = 0.16\frac{mol}{l} \qquad (1);$$

$$[H_2S] = 0.16\frac{mol}{l} - [H_2] \qquad (2);$$

$$[H_2S] = 14.5 * [H_2] \qquad (3).$$

When we put the expression for $[H_2S]$ from (3) into (2), we obtain the following equations:

$$14.5 * [H_2] = 0.16\frac{mol}{l} - [H_2];$$

$$15.5 * [H_2] = 0.16\frac{mol}{l};$$

$$[H_2] = \frac{0.16}{15.5} \approx 1.0323 * 10^{-2}\frac{mol}{l};$$

$$[H_2S] = 14.5 * [H_2] = 14.5 * 1.0323 * 10^{-2} \approx 14.9677 * 10^{-2}\frac{mol}{l}.$$

To check if the calculations are correct, we can sum the equilibrium concentrations of hydrogen and hydrogen sulfide and must obtain 0.16 mol/l:

$$[H_2] + [H_2S] = 1.0323 * 10^{-2}\frac{mol}{l} + 14.9677 * 10^{-2}\frac{mol}{l} = 0.16\frac{mol}{l}.$$

Answer: the equilibrium concentrations of hydrogen and hydrogen sulfide are $1.0323*10^{-2}$ mol/l and $14.9677*10^{-2}$ mol/l, respectively.