expect that the reaction would continue until all of the I_2 is used up. The violet color of the tube would decrease in intensity until all of the iodine reacts. At that time, the tube would be colorless, because both HI and the excess H_2 are colorless gases.

In actuality, the color fades to a constant intensity but does not disappear completely because the reaction is reversible. Hydrogen iodide decomposes to re-form hydrogen and iodine. The rate of this reverse reaction increases as the concentration of hydrogen iodide increases. The rate of the forward reaction decreases accordingly. The concentrations of hydrogen and iodine decrease as they react. As the rates of the opposing reactions become equal, equilibrium is established. The constant color achieved indicates that equilibrium exists among hydrogen, iodine, and hydrogen iodide. The net chemical equation for the reaction system at equilibrium follows.

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

From this chemical equation, the following chemical equilibrium expression can be written. The concentration of HI is raised to the power of 2 because the coefficient of HI in the balanced chemical equation is 2.

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

Chemists have carefully measured the concentrations of H_2, I_2, and HI in equilibrium mixtures at various temperatures. In some experiments, the flasks were filled with hydrogen iodide at known pressure. The flasks were held at fixed temperatures until equilibrium was established. In other experiments, hydrogen and iodine were the original substances. Experimental data, together with the calculated values for K, are listed in **Table 1.** Experiments 1 and 2 began with hydrogen iodide. Experiments 3 and 4 began with hydrogen and iodine. Note the close agreement obtained for the numerical values of the equilibrium constant in all cases.

At 425°C, the equilibrium constant for this equilibrium reaction system has the average value of 54.34. This value for K is constant for any system of H_2, I_2, and HI at equilibrium at this temperature. If the

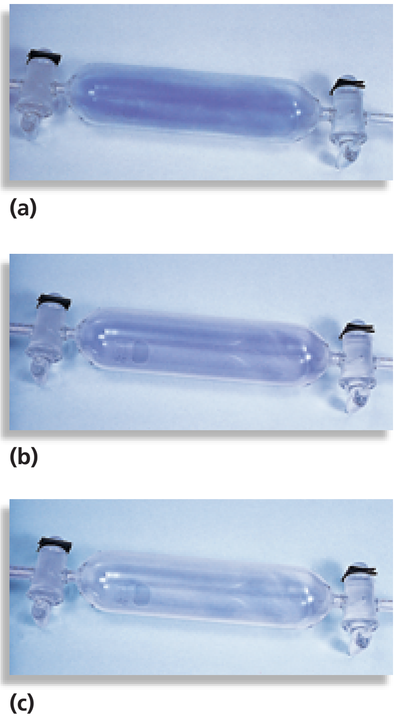

FIGURE 3 Hydrogen iodide gas is produced from gaseous hydrogen and iodine. The violet color of iodine gas (a) becomes fainter as the reaction consumes the iodine (b). The violet does not disappear but reaches a constant intensity when the reaction reaches equilibrium (c).

TABLE 1 ***Typical Equilibrium Concentrations of H_2, I_2, and HI in mol/L at 425°C***

Experiment	$[H_2]$	$[I_2]$	$[HI]$	$K = \frac{[HI]^2}{[H_2][I_2]}$
1	0.4953×10^{-3}	0.4953×10^{-3}	3.655×10^{-3}	54.46
2	1.141×10^{-3}	1.141×10^{-3}	8.410×10^{-3}	54.33
3	3.560×10^{-3}	1.250×10^{-3}	15.59×10^{-3}	54.62
4	2.252×10^{-3}	2.336×10^{-3}	16.85×10^{-3}	53.97