Uses for Hydrogen Gas

Chemical plants make a mixture of hydrogen gas and carbon monoxide gas called synthesis gas.

A shift converter converts carbon monoxide and water into more hydrogen gas and carbon dioxide.

Hydrogen gas

Ammonia for fertilizers, explosives, plastics, and fibers

HCl for cleaning metals acidifying oil wells, food processing, and the manufacture of many other chemicals

Reduction of metal oxides to form pure metals

Methanol, used to make formaldehyde, acetic acid, MTBE, and many other chemicals

Dynamic Equilibrium

$$CO(g) + H_2O(g) \underset{}{\overset{870\ °C}{\rightleftharpoons}} CO_2(g) + H_2(g)$$

- This reaction is significantly reversible.
- When CO(g) and H_2O(g) are combined, they begin to form CO_2(g) and H_2(g).
- The CO_2(g) and H_2(g) react to reform CO(g) and H_2O(g).
- Very quickly, the system comes to a point where although the forward and reverse reactions continue, the rates of these two reactions are equal, so there is no net change in the concentration of CO(g), H_2O(g), CO_2(g), and H_2(g).
- This is called a dynamic equilibrium.

Rates of Forward and Reverse Reactions

$$CO(g) + H_2O(g) \overset{870\ °C}{\rightleftharpoons} CO_2(g) + H_2(g)$$

$$\text{rate of reaction} = \frac{\text{number of moles of product formed}}{\text{liter} \cdot \text{second}}$$

$$\text{rate of forward reaction} = R_f = \frac{\text{mol } CO_2 \text{ formed}}{L \cdot s} = \frac{\text{mol } H_2 \text{ formed}}{L \cdot s}$$

$$\text{rate of reverse reaction} = R_r = \frac{\text{mol CO formed}}{L \cdot s} = \frac{\text{mol } H_2O \text{ formed}}{L \cdot s}$$

Concentration and Rates of Reaction

$$CO(g) + H_2O(g) \overset{870\ °C}{\rightleftharpoons} CO_2(g) + H_2(g)$$

- Increased concentration of a gaseous reactant will decrease the average distance between reactant particles, leading to more collisions between reactant molecules per second, which increases the rate of the reaction.
- For our reversible reaction, if the concentration of hydrogen gas and/or carbon dioxide gas increases, the rate of collisions between hydrogen molecules and carbon dioxide molecules will increase, and the rate of the reverse reaction will increase.

Concentration and Rates of Reaction

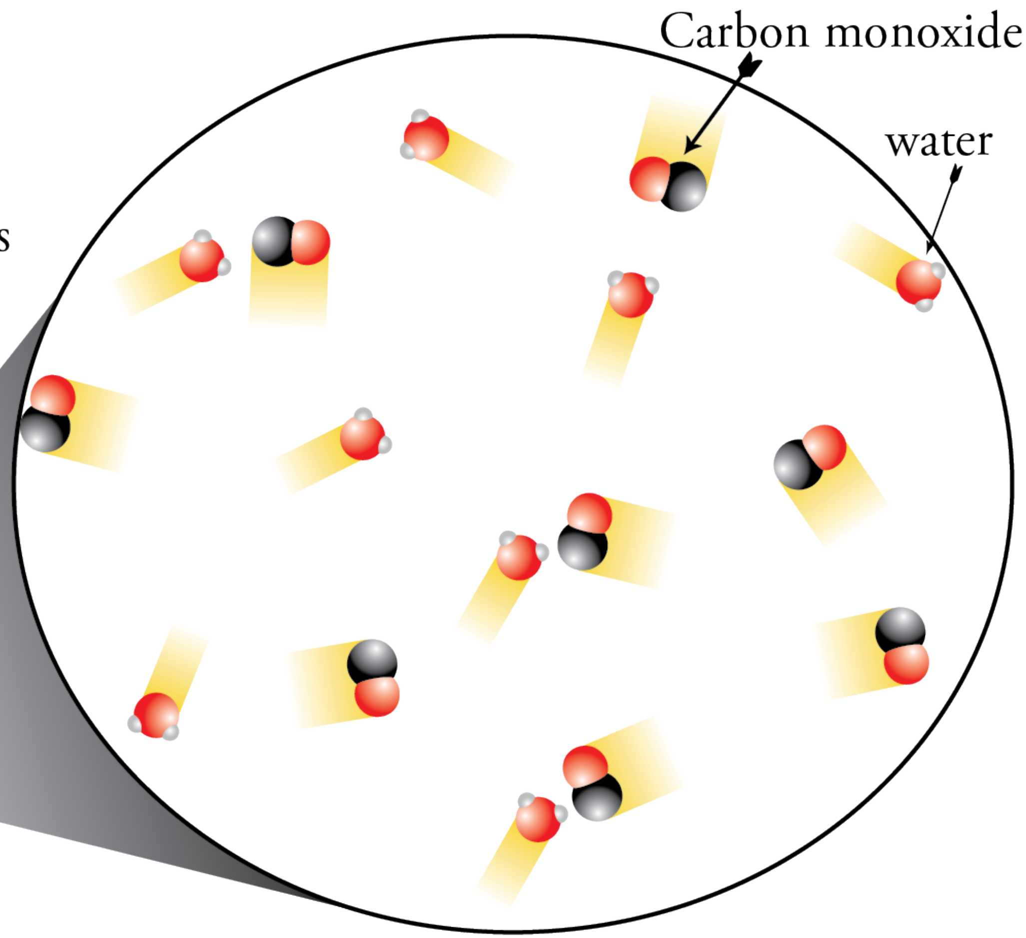
$$CO(g) + H_2O(g) \underset{}{\overset{870\ °C}{\rightleftharpoons}} CO_2(g) + H_2(g)$$

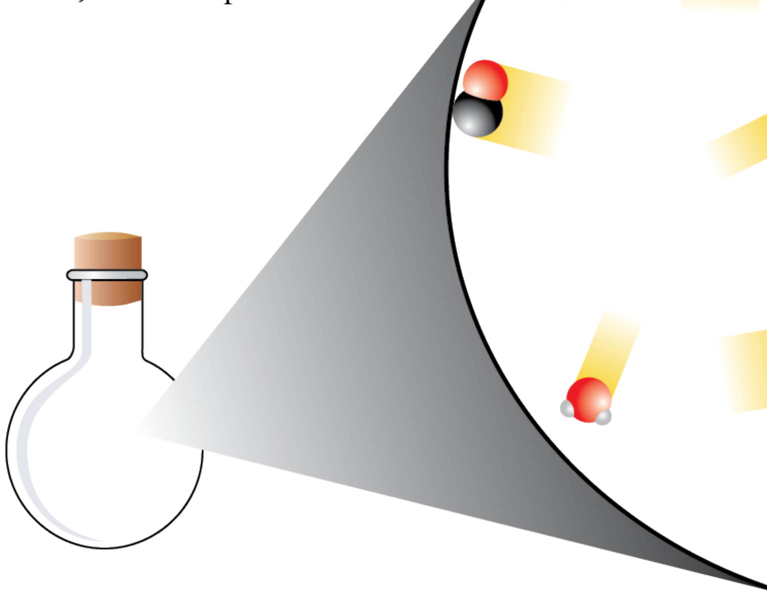
- Decreased concentration of a gaseous reactant in a reaction will increase the average distance between reactant particles, leading to fewer collisions between reactant molecules per second, which decreases the rate of the reaction.
- For our reaction, if the concentration of carbon monoxide gas and/or gaseous water decreases, the rate of collisions between carbon monoxide molecules and water molecules will decrease, and the rate of the forward reaction will decrease.

Dynamic Equilibrium

$$CO(g) + H_2O(g) \overset{870\ °C}{\rightleftharpoons} CO_2(g) + H_2(g)$$

Initially, there is only CO and H_2O in the container, so R_f is at its peak, and R_r is zero.

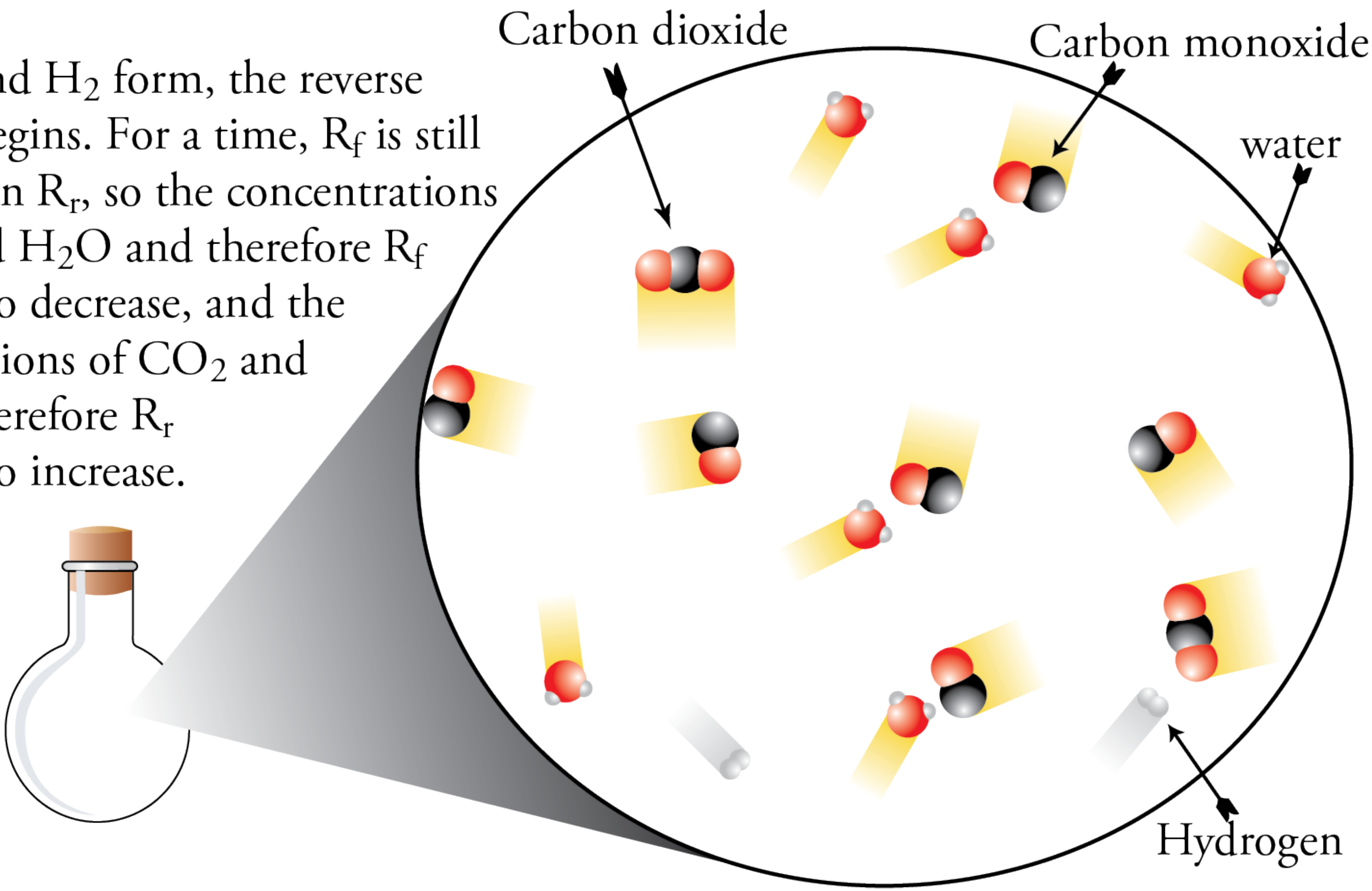




Dynamic Equilibrium

$$CO(g) + H_2O(g) \underset{}{\overset{870\ °C}{\rightleftharpoons}} CO_2(g) + H_2(g)$$

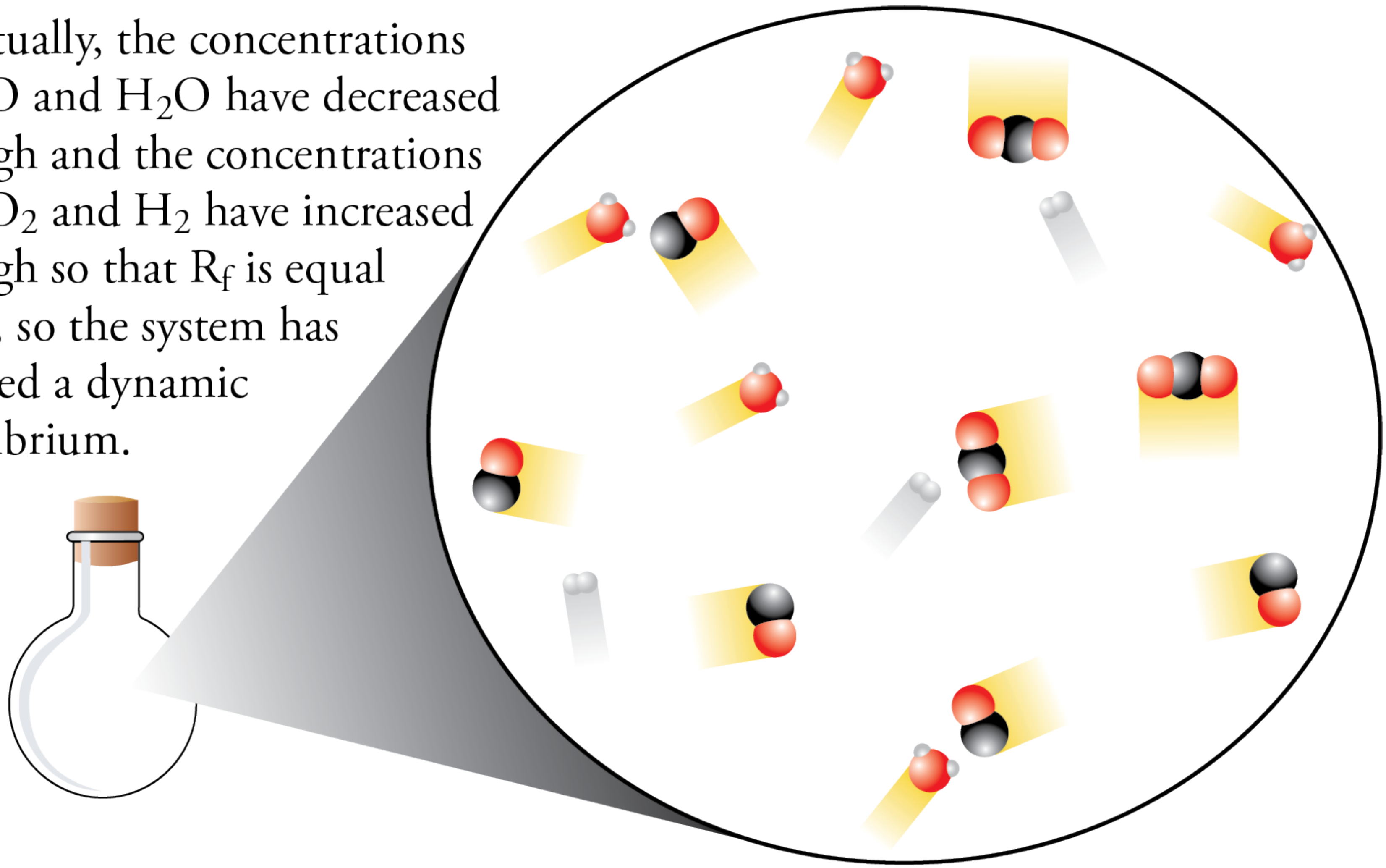
As CO_2 and H_2 form, the reverse reaction begins. For a time, R_f is still greater than R_r, so the concentrations of CO and H_2O and therefore R_f continue to decrease, and the concentrations of CO_2 and H_2 and therefore R_r continue to increase.



Dynamic Equilibrium

$$CO(g) + H_2O(g) \underset{}{\overset{870\ °C}{\rightleftharpoons}} CO_2(g) + H_2(g)$$

Eventually, the concentrations of CO and H_2O have decreased enough and the concentrations of CO_2 and H_2 have increased enough so that R_f is equal to R_r, so the system has reached a dynamic equilibrium.

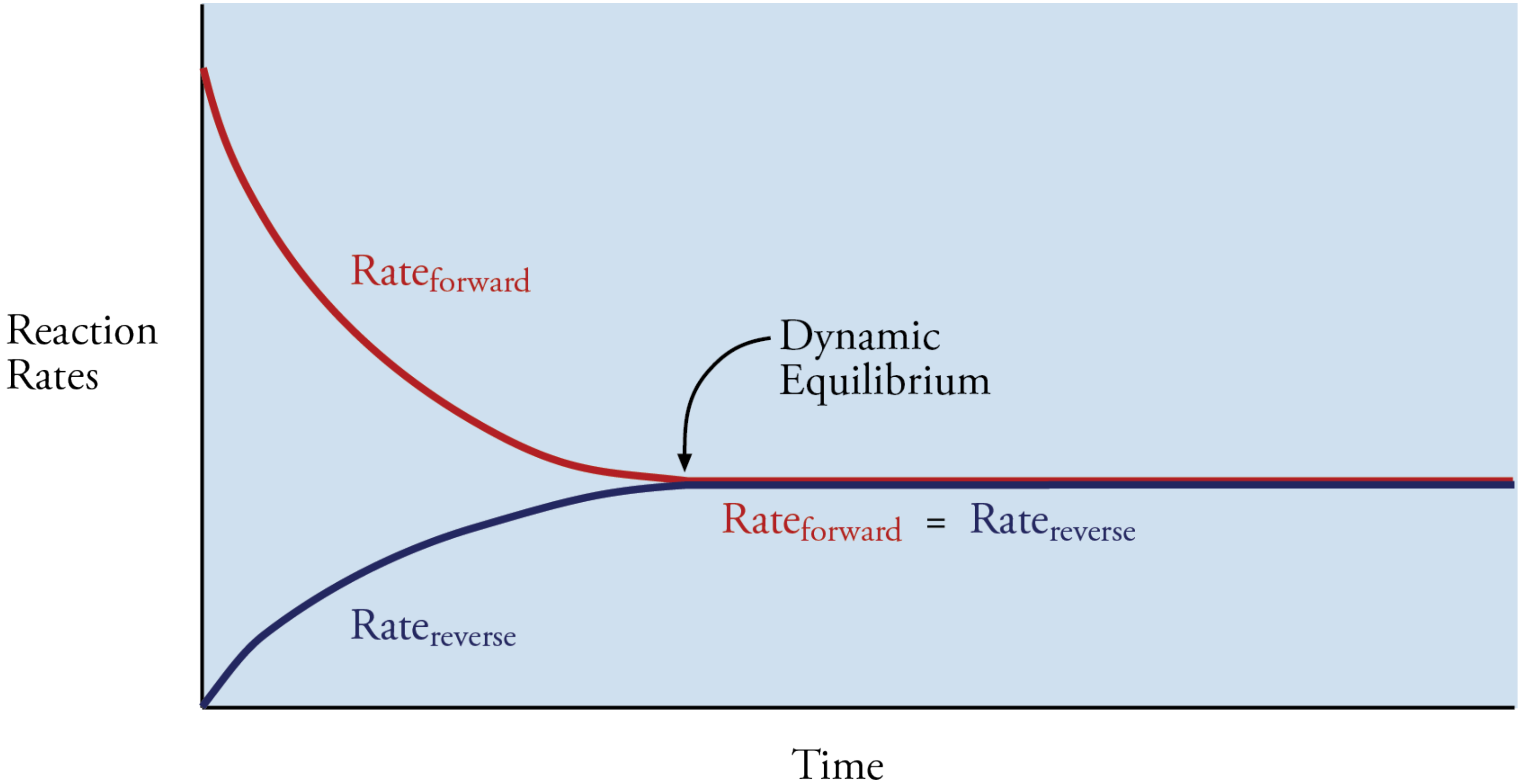


Summary

$$CO(g) + H_2O(g) \xrightleftharpoons{870\ °C} CO_2(g) + H_2(g)$$

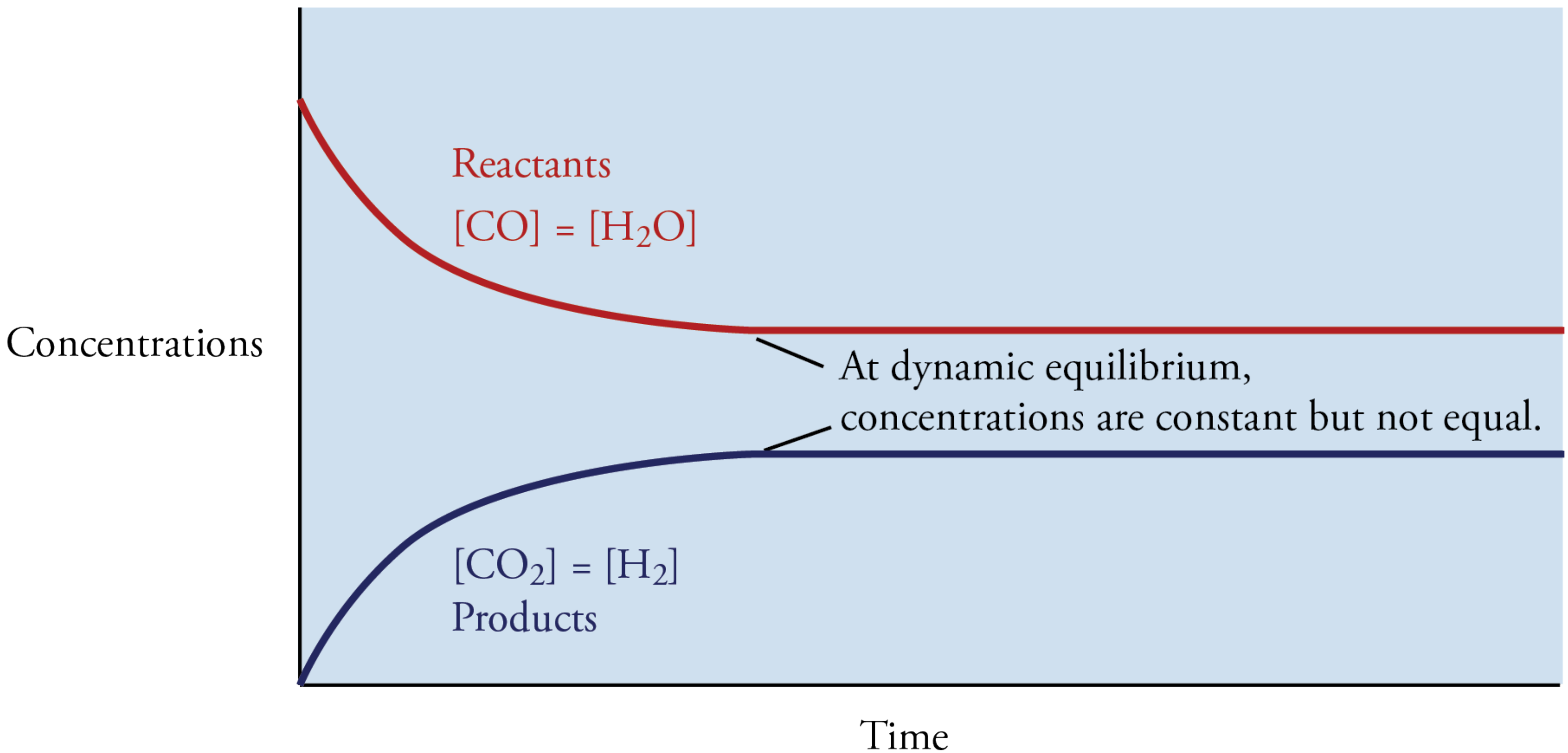
- Initially, R_f at its peak and $R_r = 0$.
- Decreased [CO] and [H_2O], so decreased R_f
- Increased [CO_2] and [H_2], so increased R_r
- At some point the rates become equal, so although the forward and reverse reactions continue, there is no net change in amounts of reactants and products…dynamic equilibrium.

Rates of Reaction for Reversible Reactions

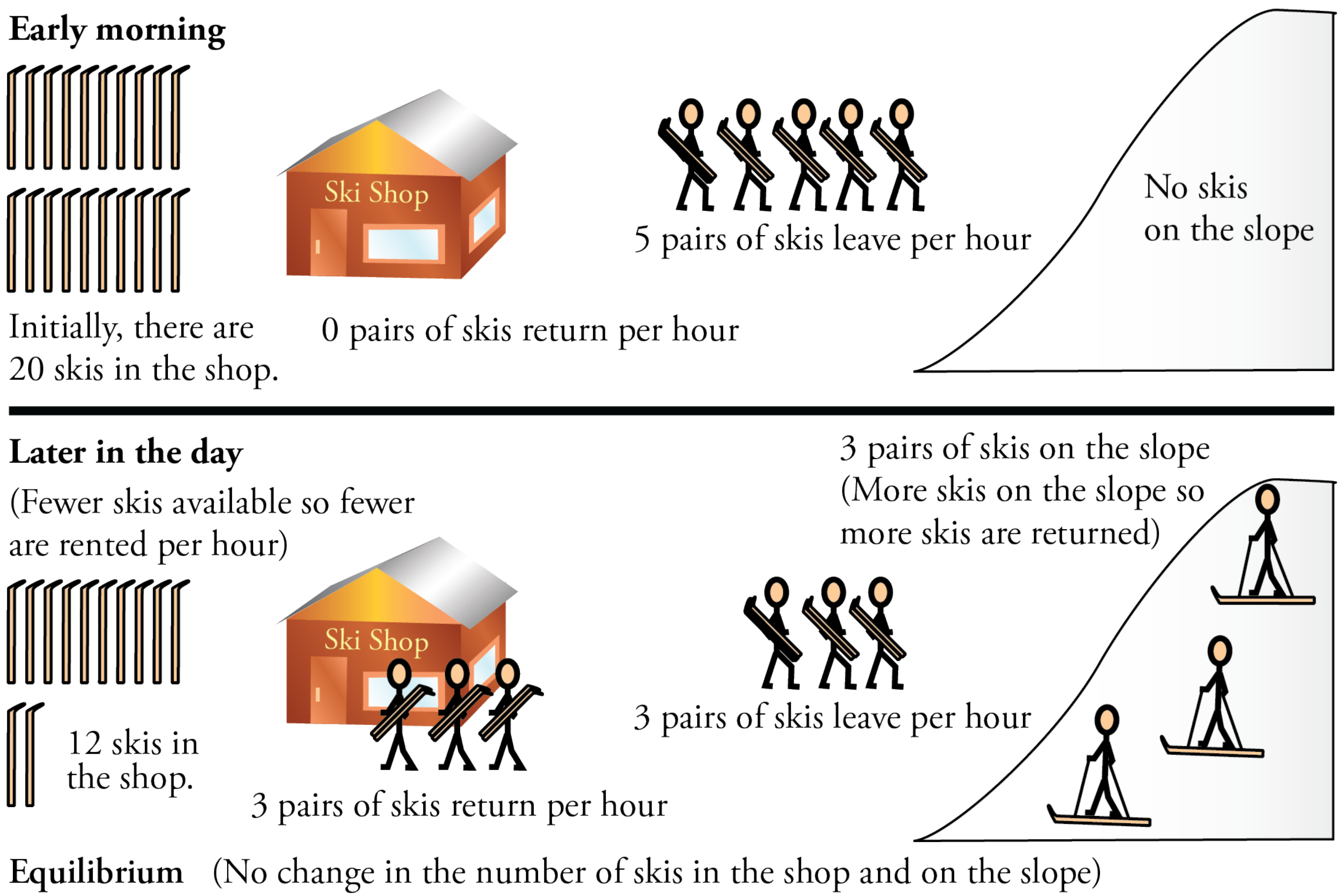


Changes in Concentrations for a Reversible Reaction

- Constant concentrations does not mean equal concentrations.



Ski Shop Analogy for Equilibrium



Equilibrium Constant

$$aA + bB + \ldots \rightleftharpoons eE + fF + \ldots$$

$$\text{Equilibrium constant} = K_C = \frac{[E]^e\,[F]^f \ldots}{[A]^a\,[B]^b \ldots}$$

$$\text{Equilibrium constant} = K_P = \frac{P_E{}^e\,P_F{}^f \ldots}{P_A{}^a\,P_B{}^b \ldots}$$

Equilibrium Constant Expression

$$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$$

The coefficient before H_2 is 3, so we raise the concentration or pressure to the third power.

$$K_C = \frac{[CO]\,[H_2]^3}{[CH_4]\,[H_2O]} \qquad K_P = \frac{P_{CO}\,P_{H_2}^3}{P_{CH_4}\,P_{H_2O}}$$

- Because all the reactants and products are gaseous, this system is an example of a **homogeneous equilibrium**, an equilibrium system in which all of the components are in the same state.

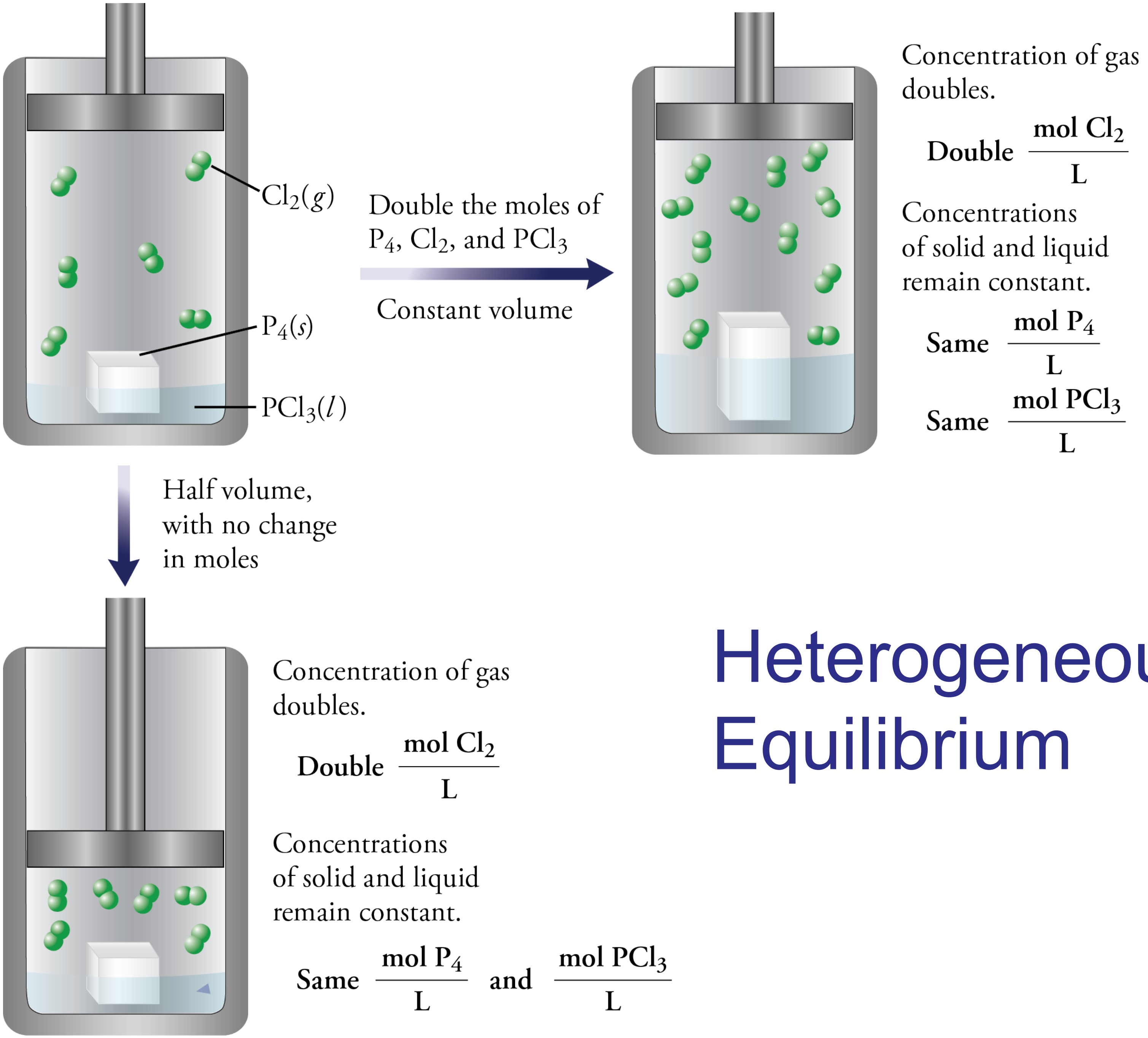
Heterogeneous Equilibrium

$$P_4(s) + 6Cl_2(g) \rightleftharpoons 4PCl_3(l)$$

$$K' = \frac{[PCl_3]^4}{[P_4]\,[Cl_2]^6} \qquad \frac{K'[P_4]}{[PCl_3]^4} = \frac{1}{[Cl_2]^6} = K_C$$

$$K_P = \frac{1}{P_{Cl_2}^6}$$

Heterogeneous Equilibrium



Equilibrium Constant Expressions for Heterogeneous Equilibria

$$SO_2(g) + 2H_2S(g) \rightleftharpoons 3S(s) + H_2O(g)$$

The solid does not appear in the K_C and K_P expressions.

$$K_C = \frac{[H_2O]}{[SO_2]\,[H_2S]^2} \qquad K_P = \frac{P_{H_2O}}{P_{SO_2}\,P_{H_2S}^2}$$

Extent of Reaction

$K > 10^2$	Products favored at equilibrium
$K < 10^{-2}$	Reactants favored at equilibrium
$10^{-2} < K < 10^2$	Neither reactants nor products favored

Example

Using the information on Table 14.1 of the atoms-first version of my text or 16.1 of the chemistry-first version, predict whether each of the following reversible reactions favors reactants, products, or neither at 25 ºC.

a. This reaction is partially responsible for the release of pollutants from automobiles.

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

b. The $NO_2(g)$ molecules formed in the reaction in part (a) can combine to form N_2O_4.

$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$