N15 - Equilibrium

Reaction Quotient – Q and ICE Tables

Link to YouTube Presentation: https://youtu.be/_658doDe_xQ

N15 - Equilibrium

Reaction Quotient – Q and ICE Tables

Target: I can compare the values of Q and K to determine which direction a reaction will proceed, and can use ICE Tables to perform calculations related to equilibrium.

The Reaction Quotient

If a reaction mixture containing both reactants and products is not at equilibrium, how can we determine in which direction it will proceed?

The answer is to compare the current concentration ratios to the equilibrium constant.

The concentration ratio of the products (raised to the power of their coefficients) to the reactants (raised to the power of their coefficients) is called the **reaction quotient, *Q*.**

Significance of the Reaction Quotient

$aA + bB \Leftrightarrow cC + dD$

the **reaction quotient** is:

$$Q_c = \frac{[C]^c[D]^d}{[A]^a[B]^b} \qquad Q_p = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

Set up the same way as K, but the [ ] values you plug in are not necessarily the same as the ones at equilibrium.

The Reaction Quotient

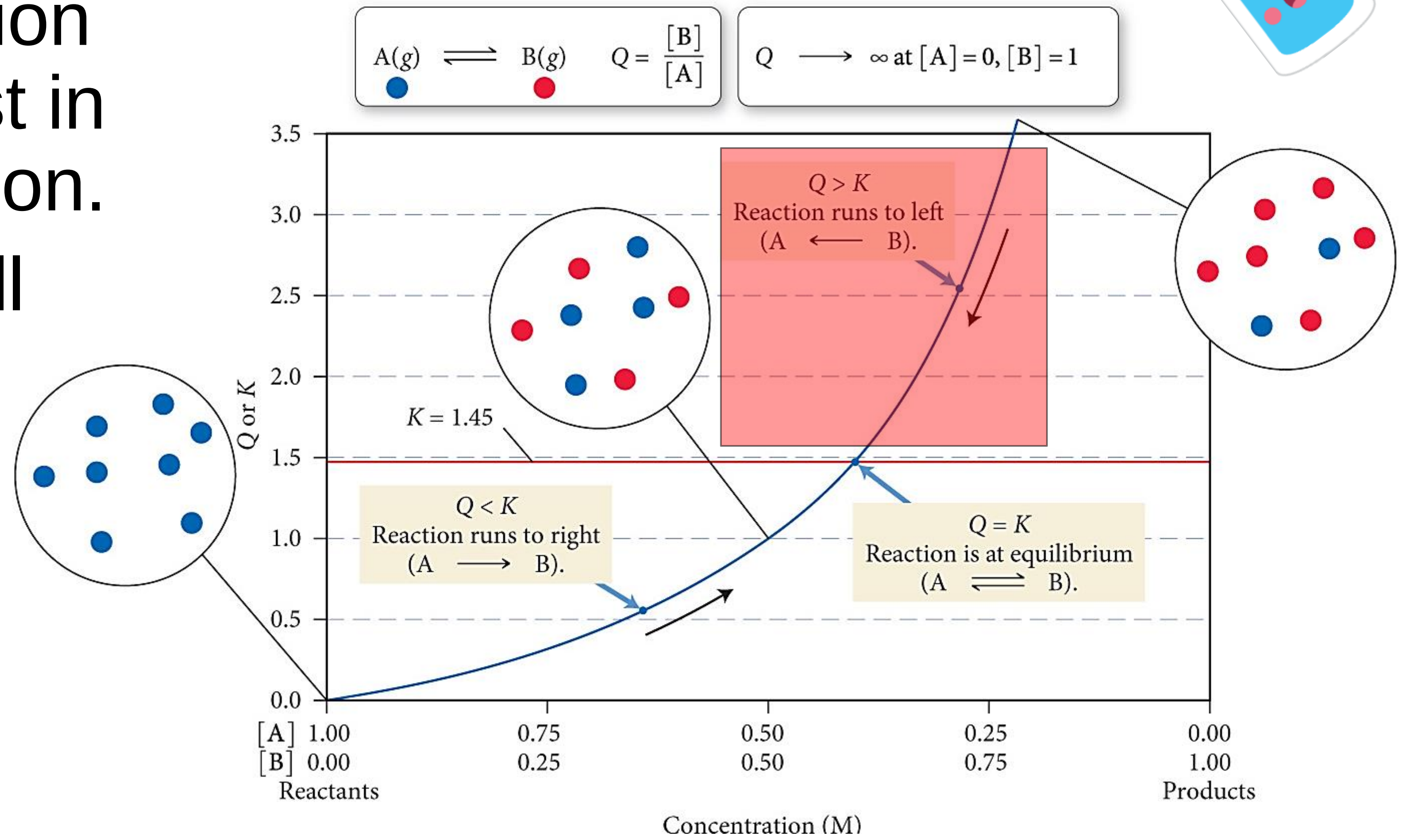
If $K = Q$, the system is **at equilibrium**

If $Q > K$, there are more products than when at equilibrium, **the system proceeds to the left,** consuming products and forming reactants until equilibrium is achieved

If $Q < K$, there are more reactants than when at equilibrium, **the system proceeds to the right,** consuming reactants and forming products until equilibrium is achieved

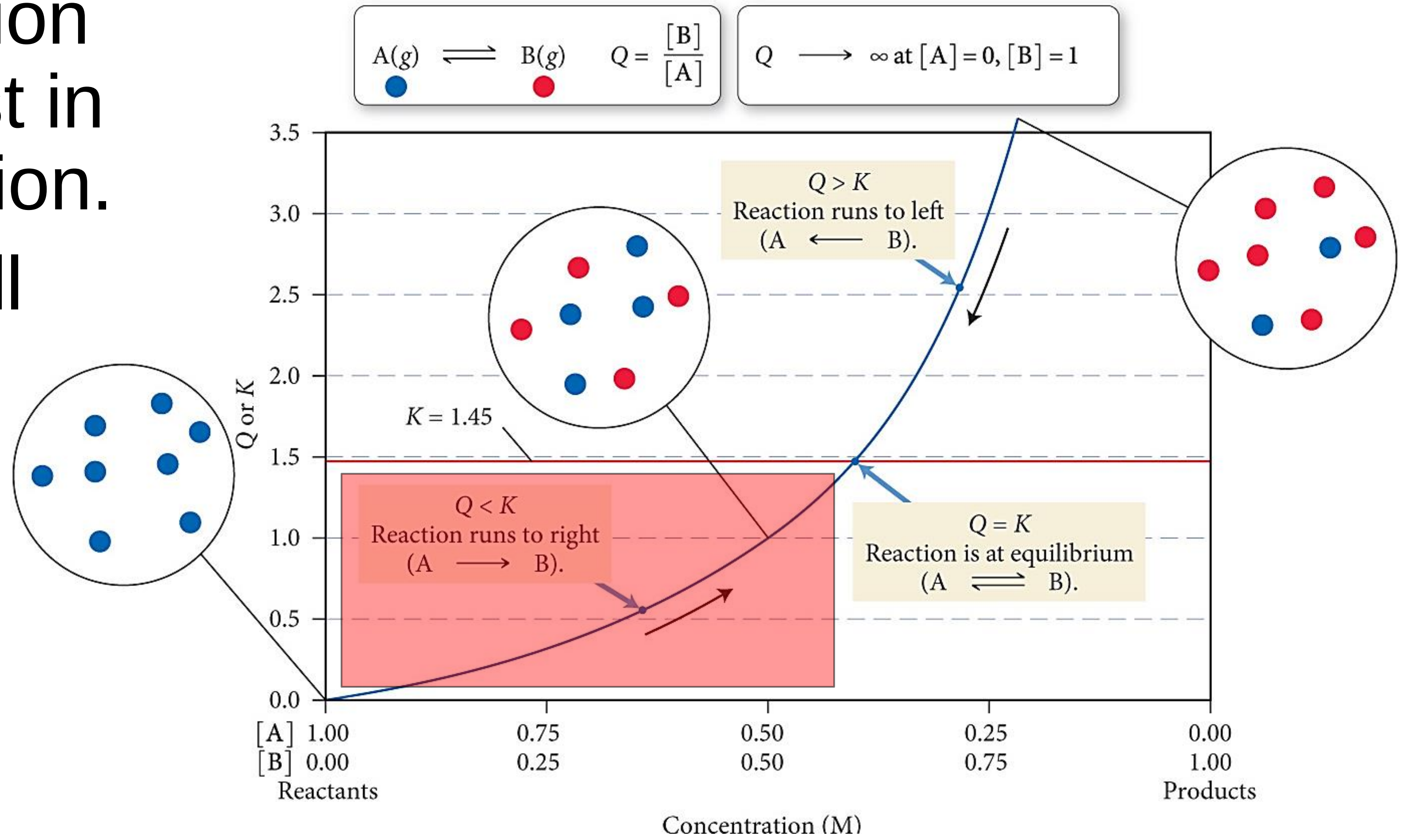
The Reaction Quotient – Predicating the Direction of Change

- If ***Q* > *K***, the reaction will proceed fastest in the **reverse** direction.
  - The products will decrease and reactants will increase.



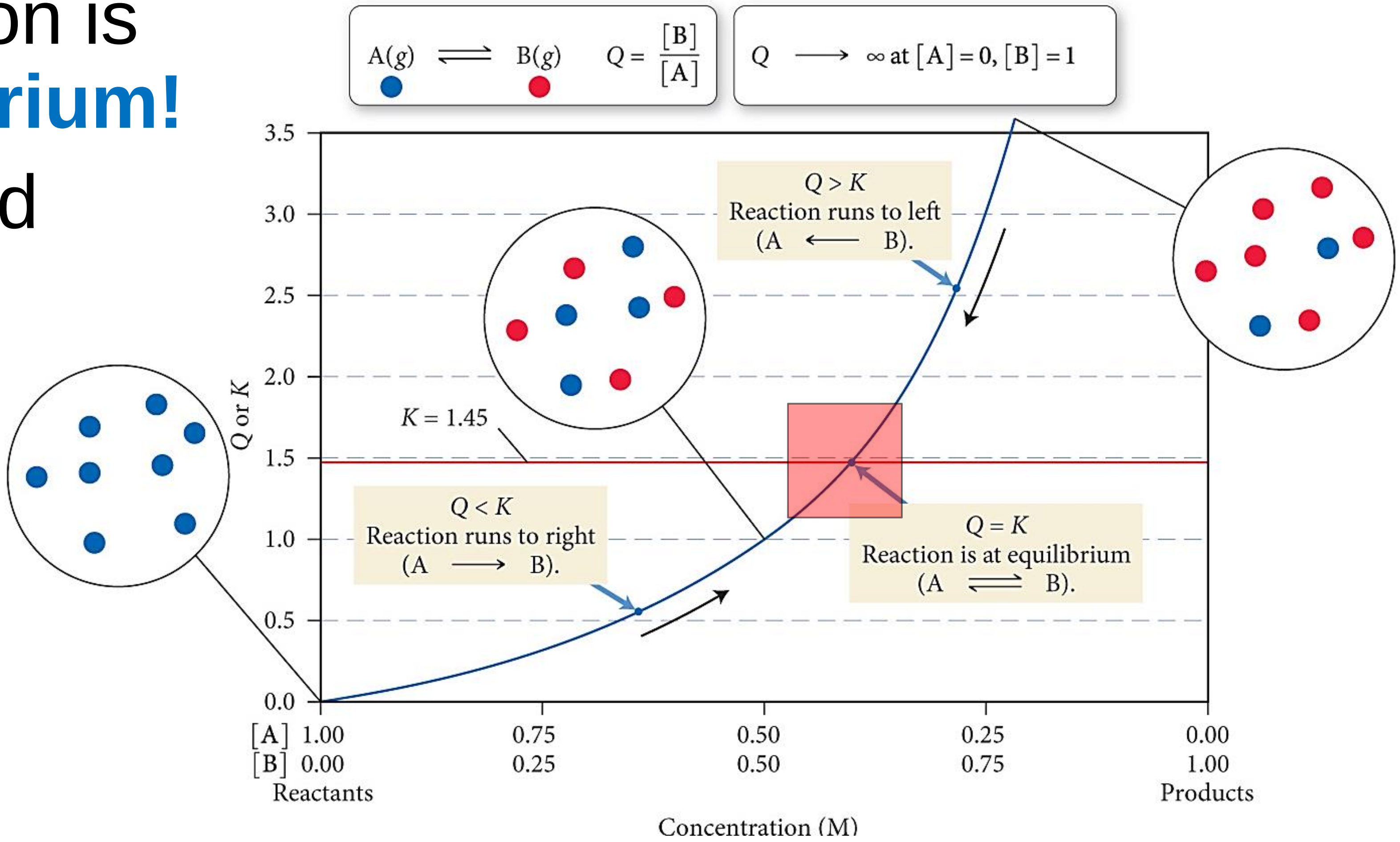
The Reaction Quotient – Predicating the Direction of Change

- If ***Q* < *K***, the reaction will proceed fastest in the **forward** direction.
  - The products will increase and reactants will decrease.



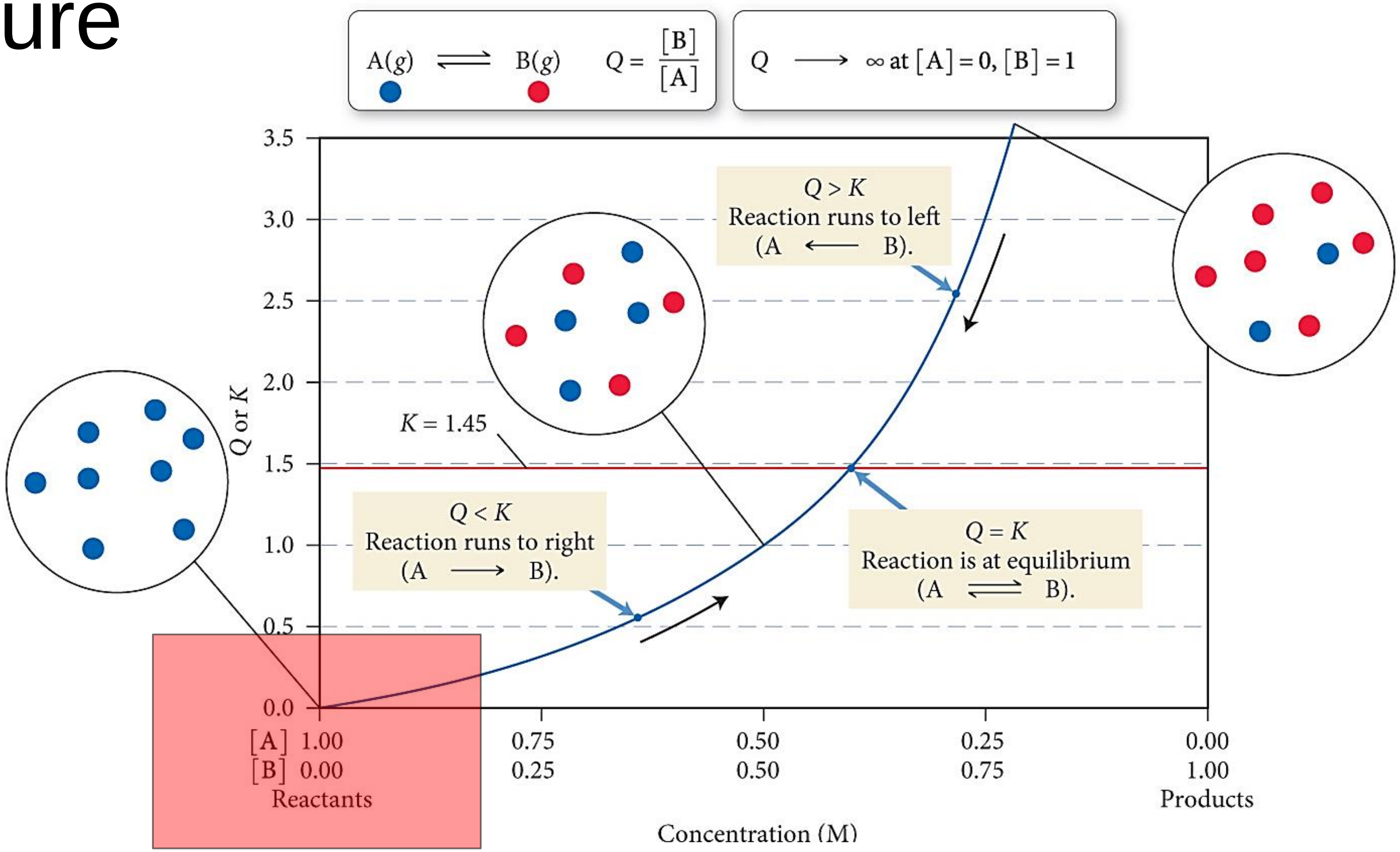
The Reaction Quotient – Predicating the Direction of Change

- If ***Q* =*K***, the reaction is **already at equilibrium!**
  - The products and reactants will not change.



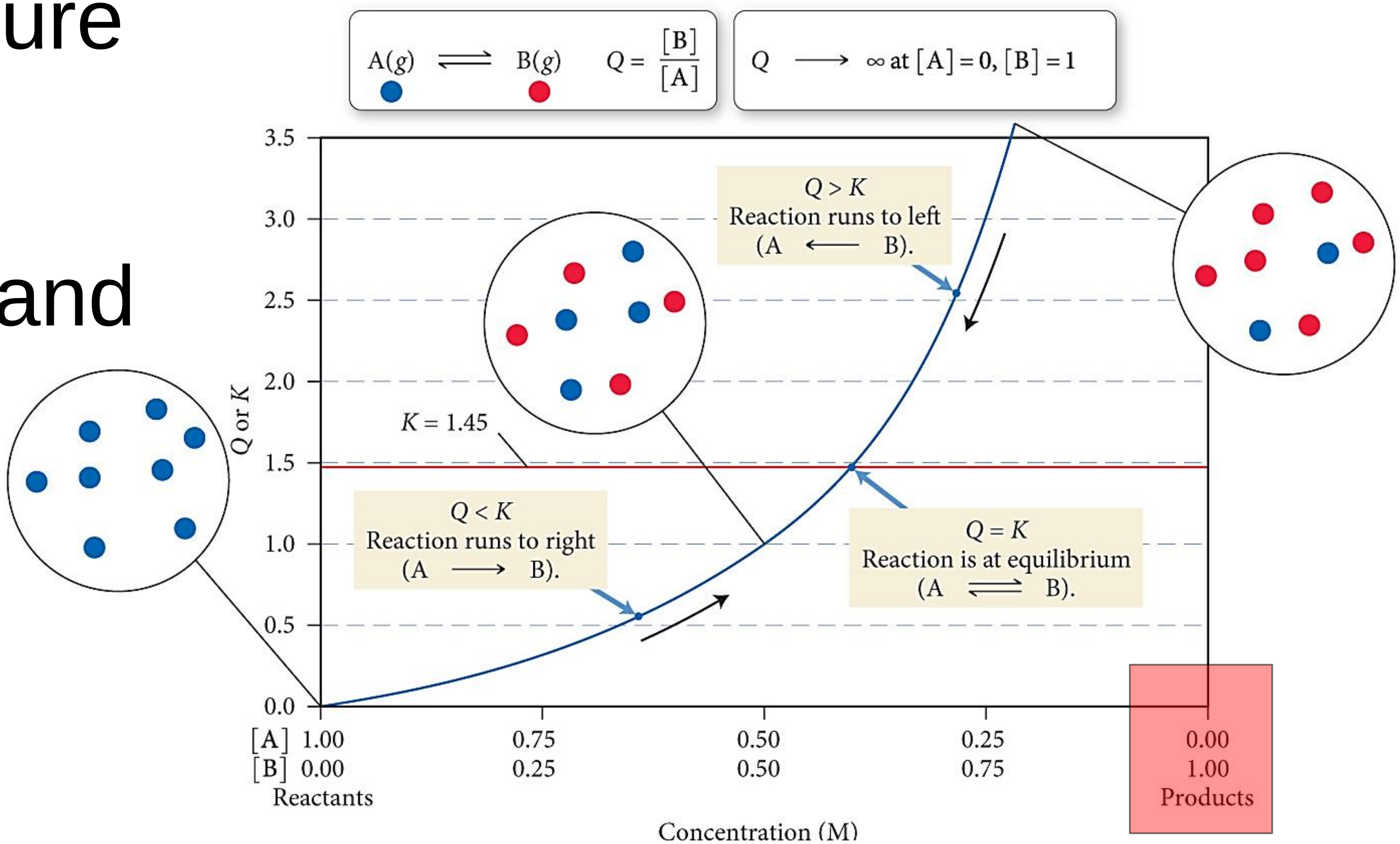
The Reaction Quotient – Predicating the Direction of Change

- If a reaction mixture contains just reactants, then **Q = 0**, and the reaction will proceed in the **forward direction**.



The Reaction Quotient – Predicating the Direction of Change

- If a reaction mixture contains just products, then **Q = undefined**, and the reaction will proceed in the **reverse direction**.



Solving for Equilibrium Concentrations

Consider this reaction at some temperature:

$H_2O(g) + CO(g) \leftrightarrows H_2(g) + CO_2(g)$ $K = 2.0$

Assume you start with 8 moles of H_2O and 6 moles of CO, in a **2 L container**. What is the concentration of H_2O, CO, H_2, and CO_2 present at equilibrium?

Time for "ICE Tables" – one the most important problem solving technique in the year. You will use it a lot!

Solving for Equilibrium Concentration

Step #1: We "ICE" the problem, beginning with the Initial concentrations – ***always check your volume!***

$$H_2O(g) + CO(g) \leftrightarrows H_2(g) + CO_2(g)$$

Initial:	4	3	0	0
Change:	-x	-x	+x	+x
Equilibrium:	4-x	3-x	x	x

***Don't forget to check your coefficients! Not always 1 x !!!**

Solving for Equilibrium Concentration

$H_2O(g) + CO(g) \leftrightarrows H_2(g) + CO_2(g)$ $K = 2.0$

Step #2: Write the law of mass action for the rxn

$$2.0 = \frac{[H_2][CO_2]}{[H_2O][CO]}$$

Solving for Equilibrium Concentration

Step #3: Plug equilibrium concentrations into the equilibrium expression, and solve for x

$$H_2O(g) + CO(g) \leftrightarrows H_2(g) + CO_2(g)$$

Equilibrium:	**4-x**	**3-x**	**x**	**x**

$$2.0 = \frac{(x)(x)}{(4-x)(3-x)}$$

Some of These Can Get Really Tricky…

- **Quadratic Equations** - *yuuuck*
- **Perfect squares** - *example:* $9 = (5 - x)^2$
  take square root of both sides… $3 = 5 - x$, *so* $x = 2$
- **Approximations** – Remember the 5% rule from Honors Chemistry ???? ($K<1$, $x < 5\%$ of initial concentrations)
- Crazy substitutions and rearrangements

But the concept is always the same!
You HAVE to practice to see crazy examples!

Quick Reminder...

Sometimes you have enough info to know what x is without having to use the 5% rule! Do not drop the x when you don’t have to! You don’t want a rounded/approximated answer when you don’t need one!

Solving for Equilibrium Concentration

Step #3: Plug equilibrium concentrations into the equilibrium expression, and solve for x

$$H_2O(g) + CO(g) \leftrightarrows H_2(g) + CO_2(g)$$

Equilibrium:	**4-x**	**3-x**	**x**	**X**

$$x = 2$$

Answer	**4-2 = 2 M**	**3-2 = 1 M**	**2 M**	**2 M**

****Note* If you had used 5% rule your answer would have been x = 4.9.***
That doesn't even make sense, does it?!

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