**Chapter 14 Equilibrium reading guide**

Be active while reading the text. Take notes, think about what you’ve read, and ask yourself questions while reading. Use this document as a guide for making your reading more meaningful.

**Section 14.1: Baby blood.**

Ignore the creepy baby picture, but read about how a fetus gets O₂ through blood. It’s all about equilibrium. One equilibrium constant is bigger than another one. How does that help a fetus get oxygen?

The two “reaction shifts...” pictures are a big deal in chemistry, see section 14.9 for further details.

**Section 14.2: What is dynamic equilibrium?**

What does the word ‘dynamic’ mean in reference to chemical equilibrium?

Do reactions at equilibrium ever stop?

Are the concentrations of the reactants necessarily equal at equilibrium? What can you say about the concentration of a reactant (or a product) when a system is at equilibrium?

How does the upper 2/3 of figure 14.2 show that the system is at equilibrium?

In the lower 1/3 of figure 14.2, what is described as being equal at equilibrium?

Figure 14.3 is nice, but remember that atoms don’t have feelings and won’t move from one place to another because they want to. Reactions reach equilibrium because that is the lowest potential energy state.

**Section 14.3: What is an equilibrium constant?**

Can you use the coefficients of an overall balanced chemical equation as the exponents in the equilibrium expression?

What does a large equilibrium constant tell a chemist about the ratio of products to reactants, or about how far towards products a reaction progressed before equilibrium was established?

If K is a small number, did the reaction go towards products to a great extent?

Check out figures 14.4. and 14.5 to see how the reaction progress is reflected in the value of K.

I don’t really find the ‘multiply an equilibrium constant by a number’ or ‘add equations and multiply Ks’ very useful to learn.
Section 14.4: How do $K_c$ and $K_p$ relate to one another?

There is a derivation of how $K_p$ can be calculated from $K_c$. The end result, equation 14.9, is useful.

The $\Delta n$ in 14.9 results from the differences in total moles of gas particles, taken as total moles of gas particles in the products minus those in the reactants. Work through Example 14.3 and For Practice 14.3 to make sure you can calculate $\Delta n$ and $K_p$.

Units for $K$: We will be a bit sloppy with units of $K$, but it turns out that there are no units for $K$. The very last portion of section 14.4 describes why. We will come up against this concept in experiment 8 also.

Section 14.5: What do I include in an equilibrium expression?

Figure 14.6 is a great illustration for why solids and liquids are not included in equilibrium expressions. The bottom line is that the concentration of a solid is not changing, as long as some solid is leftover. Absolutely critical to remember this when you answer Le Chatelier questions in section 14.9.

Avoid mistakes: always write out the equilibrium expression!

Section 14.6: How are equilibrium constants calculated?

This section shows how to calculate an equilibrium constant, and introduces the ICE table. We will use ICE tables for the next three chapters. Using an ICE table helps with the stoichiometry prior to doing any calculations of equilibrium constants.

Section 14.7: $K$ and $Q$, can the direction of a reaction be predicted?

Read through this section to determine the difference between $K$ and $Q$ (hint: the actual calculation is the same, but the numbers that go in to the equation are different).

Why is $Q$ useful?

Figure 14.7 helps determine the direction of a made up reaction where $K = 1.5$, and combined with the small table of $Q$ and $K$ should allow for $Q$ and $K$ to be compared.

The summarization of change predictions is very useful too.

Section 14.8: Can concentrations at equilibrium be calculated?

Yes. This section shows you how, using ICE tables again. Thrown in is the use of a variable ($x$) to describe the unknown concentrations.

This section is why you had the quadratic formula on the math review sheet. Aren't you glad you did that already?
Examine Example 14.11. Notice that the first 3 things written in the solution are:

a. The balanced chemical equation
b. An ICE table
c. The equilibrium expression (K = prod/react)

You should always do these three steps, for every equilibrium problem.

There is a simplifying approximation at the end of this section. It's really useful if you don't like the quadratic formula. It will be used extensively in the next two chapters.

Section 14.9: Predicting the effect of a change in conditions

We're back to the overcrowded country analogy again in figure 14.8. Chemicals don’t leave because of over-crowding, but if the rate is controlled by concentration, then the rate reverse (B to A) would increase.

Figure 14.9 is great in showing the effect of adding some product and having more reactant form. Keep in mind it's all about rates. Why is the change in equilibrium after the addition of product (NO₂) attributable to rates?

The summary of effects of concentration on p 679 is very useful.

Volume changes don’t effect product and reactant pressures equally if the numbers of gas particles aren't the same on both sides of the arrow. We will prove this with the ideal gas law in class.

Figure 14.11 is great for illustrating the changes of volume on equilibrium.

What happens to the equilibrium if an inert gas is added to a system at equilibrium, increasing the overall pressure?

Temperature changes on systems at equilibrium are challenging for students. Pay close attention to the first two equations in this portion of section 14.9, the ones that show endo- or exo- thermic with heat. Adding or removing heat to a system at equilibrium is just like adding or subtracting molecules of product or reactant, in terms of shifting equilibrium.

However, for systems not at equilibrium, adding heat always makes the reactions go faster!