

# Chemistry Grade 12

## Chapter 7 Review

Elston Almeida

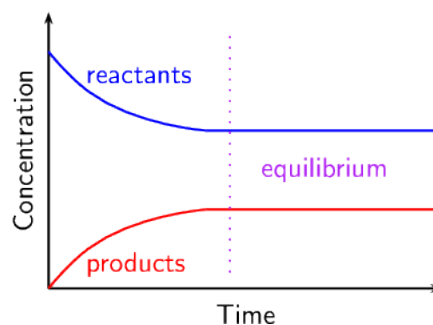
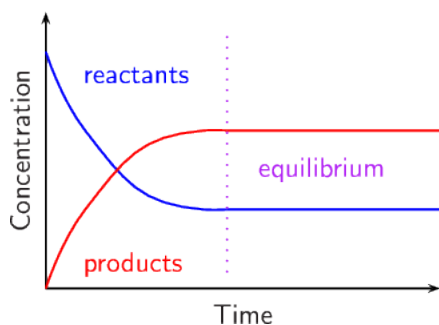
November 28, 2016

### 7.1 Chemical Equilibrium

The state where the rate of forward reaction is equal to the rate of reverse reaction is called chemical equilibrium. The left right arrows denote a reaction at equilibrium ( $\rightleftharpoons$ ). Equilibrium can only truly occur in a closed system or the reverse reaction cannot take place.

When at equilibrium, all macroscopic quantities such as color, conductivity, pH, density, etc. Remain the same. However the solution may seem static, it is rather dynamic. On the molecular level the molecules are constantly reacting in a forwards and reverse reaction  $eq^m$

This does not mean that the concentrations of the reactants and products are equal, instead it means that their concentrations stay constant, As that would mean the reaction is in the dynamic state of constantly producing and consuming the same amount (rate of forward and reverse reactions are equal). On a graph, the concentrations of the products and reactants have to be a straight line.  $m = 0, d(M(x))/dx = 0$



## 7.2 Equilibrium Law and the Equilibrium Constant



$$K_c = \frac{[A][B]}{[C][D]} \quad (1)$$

$K_c$  is the equilibrium constant and is the simple ratio of products to reactants.

$K_c$  must only be taken at the equilibrium. It is seen as the constant for the forward reaction, similarly, the reciprocal of it is the value of the reverse reaction.

$$K_{cf} = \frac{1}{K_{cr}} \quad K_{cr} = \frac{1}{K_{cf}} \quad (2)$$

$K_{cf}$  and  $K_{cr}$  defines the forward and reverse reactions respectively.  $K_c$  does not include solids or liquids, only aqueous solutions and gasses. This is because solids and liquids have a fixed density and therefore are omitted from the expression.

The very important idea to note is the equilibrium constant is only dependent on the temperature.

## 7.4 Qualitative Changes to Equilibrium

\*If the rates of temperature, pressure, or concentration change under equilibrium conditions, this would change the rates of the forward and reverse reactions.

In Accordance to **Le Chatelier's Principal**, if one of these conditions change the chemical will undergo an equilibrium shift, chemical inertia. It will resist the change and start trying to reverse the effect. Think of it as a see saw trying to constantly be perfectly level. If the concentration increases the reaction shifts to the side to consume more reactant, and vice versa.

When pressure is increased the reaction will favour the side where there are less amount of reactants, and vice versa. This can also be a factor when volume is changed due to Boyle's law, as pressure is inversely related to volume.

When temperature is increased, the reaction will favour the side to consume all the added thermal energy and the opposite effect will occur if the opposite was to occur. Only if temperature is a product or reactant in the reaction.

However, note that the  $K_c$  does not change for pressure, or concentration, only temperature. Thus why  $K_c$  values are always reported with the temperature. \* There is no effect on the system when introduced to an inert gas or a

catalyst (It is to be noted that when a catalyst is added the rate of forward and reverse reactions occur equally faster).

## 7.5 Quantitative Changes to Equilibrium

The ratio of the product over the reactants at any time other than equilibrium is considered as the Q value. The Q value helps determine the type of shift required for the system to reach equilibrium.

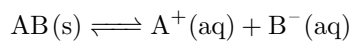
$Q = K$	system is at Equilibrium
$Q > K$	system needs to shift left
$Q < K$	system needs to shift right

## 7.6 Solubility Equilibria and the $K_{sp}$ Value

3 Different types of Equilibria:

- Phase Equilibrium: The difference physical states of a pure substance in an enclosed system.
- Chemical Reaction Equilibrium: Between reactants and products of a chemical reaction in a closed reaction.
- Solubility Equilibrium: Between a solvent and a solute in a saturated closed system.

At  $eq^m$  water molecules and ions still collide with the surface of the crystal, but the **rate of dissolution equals the rate of crystalization**.



Solubility product constant  $K_{sp}$  can help determine how soluble is an item. Measured by saturating solutions of item or precipitate it out of solution. Generally, ionic compounds have a low  $K_{sp}$  value, and compounds with high solubility do not precipitate and do not form equilibria.

In this we can use the Q value (now called **Trial Ion Product**) to determine if a product will precipitate, you take the ions forming the precipitate and find the Q of the ion(s) specific reaction, then compare the Q with the value of  $K_{sp}$ .

$Q = K_{sp}$	Precipitate will not form, saturated solution
$Q > K_{sp}$	Precipitate will form, supersaturated solution
$Q < K_{sp}$	Precipitate will not form, undersaturated solution

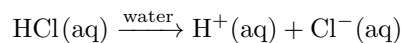
The common ion effect is to reduce the solubility of an ionic compound by adding a common ion, this can be described by Le Chatelier's Principal as the reaction tries to balance itself.

## 8 The nature of Acids and Bases — Elston Almeida

### 8.1 Arrhenius

Acid  $\text{H}^+$  ions

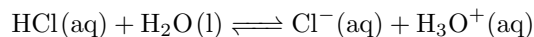
Bases  $\text{OH}^-$  ions



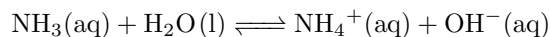
#### Brønsted Theory

Acid is a hydrogen donor. Base is a hydrogen acceptor.

### 8.2 Brønsted lowry acids



Water can be both acids and bases depending on the reaction. In the reaction above, it acts as a base to accept the hydrogen ion. (HCl donates the H ion to the water)



$\text{NH}_3$  is an base,  $\text{H}_2\text{O}$  acts as an acid,  $\text{NH}_4$  is the conjugate acid,  $\text{OH}$  is the conjugate base

Conjugate acid is when the substance that forms then a base accepts a hydrogen ion. Conjugate base is when the substance that forms when an acid loses a hydrogen ion.

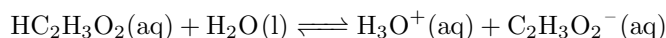
Amphiprotic is a substance that can donate and accept a hydrogen ion. Water is a Amphiprotic as you can get Hydronium, and Hydroxide.

$K_a$  value is the eq constant for the ionization of an acid (called the acid dissociation constant)

General Equation:

$$K_a = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{A}^-(\text{aq})]}{[\text{HA}(\text{aq})]} \quad (3)$$

Example:

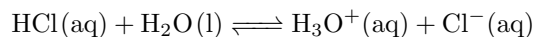


$$K_a = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})]}{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})]}$$

HW492 #1, 493 #1.

### 8.3 Strong and Weak Acids

A strong acid ionizes almost completely in water  
A weak acid is one that only partially ionizes in water



Property	Strong Acid	Weak Acid
Value of acid ionization constant, $K_a$	$K_a$ is large	$K_a$ is small
Position of the ionization equilibrium	far to the right	far to the left
Equilibrium concentration of $\text{H}^+(\text{aq})$ compared with the original concentration	$[\text{H}^+(\text{aq})]_{\text{eq}} \approx [\text{HA}(\text{aq})]_i$ Equal. pH $\approx$ Initial pH	$[\text{H}^+(\text{aq})]_{\text{eq}} \ll [\text{HA}(\text{aq})]_i$ Equal. pH $\ll$ Initial pH

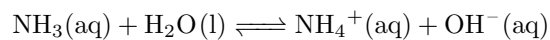
### 8.4 Strong and Weak Bases

Strong base dissociates completely in water  
Weak base partially dissociates in water

The base ionization constant( $K_b$ ) is the base equilibrium constant for the ionization of a base(it is also called the base dissociation constant)

$$K_b = \frac{[\text{BH}^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{B}(\text{aq})]} \quad (4)$$

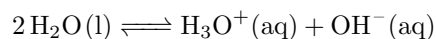
Example:



$$K_b = \frac{[\text{OH}^-(\text{aq})][\text{NH}_4^+(\text{aq})]}{[\text{NH}_3(\text{aq})]}$$

Chart for  $K_b$  for weak acids: pg 727

The autoionization of water is the transfer of a hydrogen ion from one water molecule to another.



$$K_w = [\text{H}_3\text{O}^+(\text{aq})][\text{OH}^-(\text{aq})]$$

$$K_w = [1.0 \cdot 10^{-7}][1.0 \cdot 10^{-7}]$$

$$K_w = 1.0 \cdot 10^{-14}$$

$K_w$  is always  $1.0 \cdot 10^{-14}$  at SATP

$[H] = [OH]$	<b>Neutral</b> solution
$[H] > [OH]$	<b>Acidic</b> solution
$[H] < [OH]$	<b>Basic</b> solution

Example to find the  $[\text{H}_3\text{O}^+(\text{aq})]$

$$K_w = [\text{H}_3\text{O}^+(\text{aq})][\text{OH}^-(\text{aq})]$$

$$[\text{H}_3\text{O}^+(\text{aq})] = \frac{K_w}{[\text{OH}^-(\text{aq})]}$$

**Other Equations:**

$$K_w = K_a \cdot K_b$$

$$\text{pH} = -\log[H]$$

$$[H] = 10^{-\text{pH}}$$

$$\text{pOH} = -\log[OH]$$

$$[OH] = 10^{-\text{pOH}}$$

$$14 = \text{pH} + \text{pOH}$$

$$\text{p}K_w = \text{pH} + \text{pOH}$$

- pH meter is an electronic device that measures the acidity of a solution and displays the result as a pH value.
- An acid base indicator is a substance that changes color specific to the pH range

pg 495-509

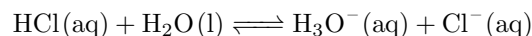
pg 502 #1,2; pg 505 #1-3; pg 508 #1-4; pg 509 #1-10

## 8.5 Calculations Involving Acidic Solutions

Since strong acids almost completely ionize in water, we can assume that the concentration of hydrogen ions is equal to the concentration of the acid.

Ex: A solution of hydrochloric acid has a concentration of 0.1M.

Calculate:



$$[\text{H}^+] = 0.1M$$

$$[\text{OH}^-] = 1 \cdot 10^{-13}M$$

$$\text{pH} = 1$$

$$\text{pOH} = 13$$

Percentage ionization is the percentage of a solute that ionizes when it dissolves in a solvent.

$$\% \text{ ionization} = \frac{[\text{ Ionized Acid }]}{[\text{ Initial Acid }]} \cdot 100\%$$

## 8.6 Monoprotic and Polyprotic acids

Monoprotic acid is an acid that possess only one ionizable hydrogen acid.

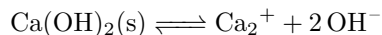
Polyprotic acid is an acid that possess more than one ionizable.

$K_{a1}$  is larger than the other  $K_{an}$  values:

$$K_{a1} > K_{a2} > K_{a3}$$

If the 5 percent rule is does not work the pH is zero

## 8.7 Calculations involving basic solutions



Metal oxides dissolve in water to produce a basic solution

Non-metallic oxides dissolve in water to produce acidic

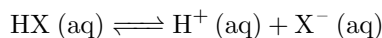
## 8.8 Acid-Base Titration

- Titration is to determine the pH of a solution by neutralization.
- Titrations are used to determine the concentration of an acid or base.
- The equivalence point is the point of the titration when the acid and base completely react with each other.
- If you know the volumes of both solution at the equivalence point, and the concentrations of one of them, you can calculate the unknown concentration.

## 8.9 Buffers and the Common Ion Effect

The Solubility of a partially soluble salt is decreased with the addition of a common ion.

A buffered solution consists of a mixture of a weak acid and its conjugate base  $\text{X}^-$ :



$$K_a = \frac{[\text{H}^+][\text{X}^-]}{[\text{HX}]}$$

### Buffer Capacity and pH

- Buffer capacity is the amount of acid or base neutralized by the buffer before there is a significant change in pH.
- Buffer capacity depends on the composition of the buffer.
- The greater the amounts of the conjugate acid-base pair, the greater the buffer capacity.
- The pH of the buffer depends on  $K_a$ .