# Bronsted-Lowry Acid Base Chemistry

Just a few reminders

## What makes an acid an acid?

- A Bronsted-Lowry Acid is a compound that donates a proton (a hydrogen ion with a positive charge, $H^+$)
  - Think of the standard acids that are commonly mentioned in Gen Chem: sulfuric acid, $\mathbf{H_2SO_4}$, hydrochloric acid, **HCl**, phosphoric acid, $\mathbf{H_3PO_4}$, nitric acid, $\mathbf{HNO_3}$
  - The term "proton" is applied to this hydrogen ion because it's a hydrogen atom whose only "exciting" piece is the one proton in its nucleus. No electron…

## And a base a base?

- A Bronsted-Lowry Base is a compound that accepts a proton
  - Think of some standard bases that are commonly mentioned in Gen Chem: Sodium hydroxide, **NaOH** or sodium bicarbonate, $\mathbf{NaHCO_3}$

## Acids react with bases

- The general reaction for acids and bases:

**Acid + Base ←→ Conj Base + Conj Acid**

## Bronsted-Lowry Acids

- When an acid donates a proton ($H^+$), it forms the conjugate base of the acid
- The following are acid/conj base pairs:
  - $H_2SO_4$, $HSO_4^-$
  - $HCl$, $Cl^-$
  - $H_3PO_4$, $H_2PO_4^-$
- Notice how all the formulas lose an $H^+$ when forming the conjugate base

## Bronsted-Lowry Acids

- When an acid donates a proton ($H^+$), it forms the conjugate base of the acid. If the acid is neutral, it forms a negatively charged conjugate base:

$$H_2SO_4 \rightarrow H^+ + HSO_4^-$$

- The charges on both sides must still balance. What would be the charge now?

$$H_2PO_4^- \rightarrow H^+ + HPO_4^{??}$$

# Bronsted-Lowry Acids

- If the left side has a (-1) charge overall, then the right side must also have a (-1) charge:

  $H_2PO_4^- \rightarrow H^+ + HPO_4^{??}$

- ?? must equal -2 in order to balance the charges

  $H_2PO_4^- \rightarrow H^+ + HPO_4^{-2}$

# Write the conjugate base

- For each of the following acids, write the conjugate base
- (the most acidic proton is highlighted to help you)
  - $CH_3CO_2H$
  - $CH_3CCH$
  - $HNO_3$

# Write the conjugate base

- The conjugate bases would be (shown in blue):
  - $CH_3CO_2H \rightarrow H^+ + CH_3CO_2^-$
  - $CH_3CCH \rightarrow H^+ + CH_3CC^-$
  - $HNO_3 \rightarrow H^+ + {}^-NO_3$

# Bronsted-Lowry Bases

- When an base accepts a proton ($H^+$), it forms the conjugate acid of the base
- The following are base/conj acid pairs:
  - $NaOH$, $H_2O$
  - $NaHCO_3$, $H_2CO_3$
  - $NaOCH_3$, $HOCH_3$
- Notice how all the formulas gain an $H^+$ when forming the conjugate acid. Anions become neutral species.

# Write the conjugate acid

- For each of the following bases, write the conjugate acid (ignore the $M^+$ metal spectator ions when present):
  - $LiOH$ (${}^-OH$)
  - $H_2O$
  - $NaOCH_3$ (${}^-OCH_3$)

# Write the conjugate acid

- The conjugate acids would be:
  - ${}^-OH + H^+ \rightarrow HOH$
  - $H_2O + H^+ \rightarrow H_3O^+$
  - ${}^-OCH_3 + H^+ \rightarrow HOCH_3$
- Anions become neutral species and neutral bases becomes positively charged conjugate acids.
- Charges MUST balance!

## Acid-Base Chemistry

- When the acid donates a proton, it becomes a conjugate base
- When the base accepts a proton, it becomes a conjugate acid
- **YOU MUST BE ABLE TO IDENTIFY ACIDS VERSUS BASES**
- **And remember – you never two acids on the same side, nor two bases!**

## Acid-Base Chemistry

- In Bronsted-Lowry Acid-Base theory, the strength of an acid is determine by the acid's ability to dissociate when in water.
- The Base, on the left side, in this process is therefore always WATER, $H_2O$

**Acid + $H_2O$ ←→ Conj Base + $H_3O^+$**

**Or**

**$HA + H_2O \leftrightarrow A^- + H_3O^+$**

## Acid-Base Chemistry

- The general equation:

**$HA + H_2O \leftrightarrow A^- + H_3O^+$**

- Write the equation for ammonium ion, $^+NH_4$, dissociating in water.

## Acid-Base Chemistry

- The general equation:

**$HA + H_2O \leftrightarrow A^- + H_3O^+$**

- For Nitric acid:

$^+NH_4 + H_2O \rightarrow NH_3 + H_3O^+$

## Acid-Base Chemistry

- The generic equation:

**$HA + H_2O \leftrightarrow A^- + H_3O^+$**

- Strong acids willingly donate protons.
- The acid, HA, gives up its proton to form the conjugate base, $A^-$
- Why?

## What makes a proton acidic?

- Strong acids, HA, form a conjugate base, $A^-$, that is highly stable.
- When a conjugate base is highly stable, its lower in Energy!
- All reactions want to drive towards a lower energy system! And this is an equilibrium process… lower energy favored…

## Bronsted-Lowry Acid/ Base theory

- The acidity constant, Ka, is determined to be the concentration of the products divided by the concentration of the reactants, at equilibrium (taking out water, which is a constant):

$$Ka = \frac{[Products]}{[Reactants]}$$

## Bronsted-Lowry Acid/ Base theory

- For the general acid-base reaction equation, the acidity constant, Ka, is determined to be:

$$Ka = \frac{[A^-][H_3O^+]}{[HA]}$$

Note that for a strong acid, with a lot of $A^-$ formation, Ka is a LARGE value.

## Bronsted-Lowry Acid/ Base theory

- Strong Acids – want to donate a proton, want to form stable conjugate base ($A^-$) – Large Ka value
- Weak Acids – do not want to donate a proton, do not want to form unstable conjugate base ($A^-$) – Small Ka value

## Bronsted-Lowry Acid/ Base theory

- Ka values cover a broad range of values – $10^{15}$ to $10^{-60}$
- With such a huge range of values, it is easier to work in logarithmic form, where

$$pKa = -\log Ka$$

## Bronsted-Lowry Acid/ Base theory

- When working in logarithms, remember there's an inverse relationship.
- Compounds with large Ka values (strong acids, want to dissociate) will have small pKa values
- Compounds with small Ka values (weak acids, won't dissociate) will have large pKa values.

## Bronsted-Lowry Acid/ Base theory

- You should be able to convert between Ka and pKa.

$$pKa = -\log Ka$$

- Calculate the pKa of $7.95 \times 10^{-10}$
- Calculate the pKa of $3.16 \times 10^{-3}$

# Bronsted-Lowry Acid/ Base theory

pKa = -log Ka

- Calculate the pKa of $7.95 \times 10^{-10}$

  pKa = - log ($7.95 \times 10^{-10}$)

  pKa = 9.1

- Calculate the pKa of $3.16 \times 10^{-3}$

  pKa = - log ($3.16 \times 10^{-3}$)

  pKa = 2.5

# Bronsted-Lowry Acid/ Base theory

- Now in the other direction:

  Ka = inv log (-pKa)

- Calculate the Ka of 6.7
- Calculate the Ka of 35

[Note: inv log will be the $10^x$ key on your calculator]

# Bronsted-Lowry Acid/ Base theory

**Ka = inv log (-pKa)**

- Calculate the Ka of 6.7

  Ka = inv log (-6.7)

  Ka = $1.99 \times 10^{-7}$

- Calculate the Ka of 12.2

  Ka = inv log (-12.2)

  Ka = $6.31 \times 10^{-13}$

[Note: if you do not have an 'inv" key, inv log may be $10^x$ on some calculators]

# Understanding Relationships

- Stronger acids have larger Ka's and smaller pKa's.

- Which acid is stronger?

  Acid 1: pKa 6.9 or Acid 2: pKa 12.6

# Understanding Relationships

- Stronger acids have larger Ka's and smaller pKa's.

  Acid 1: pKa 6.9 or Acid 2: pKa 12.6

- Since Acid 1 has the lower pKa value, it is the stronger acid.

# Understanding Relationships

- Which acid is stronger?

  Acid 1: pKa 35.4 or Acid 2: pKa 15.8

## Understanding Relationships

- Which acid is stronger?
  Acid 1: pKa 35.4 or Acid 2: pKa 15.8
- Stronger acids have lower pKa values, so Acid 2 would be the stronger acid.

## Understanding Relationships

- Which acid is stronger?
  Acid 1: Ka of $1.48 \times 10^4$ or
  Acid 2: Ka of $1.98 \times 10^{-5}$

## Understanding Relationships

Acid 1: Ka of $1.48 \times 10^4$ or
Acid 2: Ka of $1.98 \times 10^{-5}$

- Stronger acids have larger Ka values, so Acid 1 would be the stronger acid.
  $10^4 > 10^{-5}$

## Understanding Relationships

- Strong Acids have weak conjugate bases and vice versa
- Which acid has the stronger conjugate base?
  Acid 1: pKa 25 or
  Acid 2: pKa 9

## Understanding Relationships

- Which acid has the stronger conjugate base?
  Acid 1: pKa 25 or
  Acid 2: pKa 9
- The stronger conjugate base is associated with the weaker acid. The weaker acid has the larger pKa value.
  Acid 1

## Equilibrium Shift?

- In which direction will the reaction be favored?
- In an Acid-Base reaction, the reaction always proceeds to form the weaker, more stable acid/base pair

Strong Acid/Base → Weak Acid/Base

## Always towards lower energy

Strong Acid/Base → Weak Acid/Base

- Note that this means that the stronger acid is **ALWAYS** on the same side as the stronger base, and the same for the weaker acid/base pair.

## Find the strong acid…

Acid + Base → Conj Base + Conj Acid

- If you know the pKa values for the two acids shown above, you know which one is the strong one. The reaction always moves AWAY FROM the strong acid towards the weak acid.

## Which way will it go?

Acid + Base → Conj Base + Conj Acid

$^{+}NH_4 + H_2O \rightarrow NH_3 + H_3O^+$

- The pKa of $^{+}NH_4$ is 9.2. The pKa of $H_3O^+$ is -1.7. Which direction is favored?

## Find the strong versus weak

- The pKa of $^{+}NH_4$ is 9.2. The pKa of $H_3O^+$ is -1.7. Which direction is favored?
- The stronger acid has the lower pKa value. -1.7 is lower than 9.2. $H_3O^+$ is the stronger acid, making $^{+}NH_4$ the weaker acid.

$^{+}NH_4 + H_2O \rightarrow NH_3 + H_3O^+$

WA WCB SCB SA

- This reaction will move backwards.

## Which way will it go?

Acid + Base → Conj Base + Conj Acid

$CH_3CO_2H + NaOH \rightarrow CH_3CO_2Na + H_2O$

- The pKa of $CH_3CO_2H$ is 4.75. The pKa of $H_2O$ is 15.7. Which direction is favored?

## Find the strong versus the weak

- The pKa of $CH_3CO_2H$ is 4.75. The pKa of $H_2O$ is 15.7. Which direction is favored?
- The stronger acid has the lower pKa value. 4.75 is lower than 15.5. $CH_3CO_2H$ is the stronger acid, making $H_2O$ the weaker acid.

$CH_3CO_2H + NaOH \rightarrow CH_3CO_2Na + H_2O$

SA SCB WCB WA

- This reaction will move forwards.

- That pretty much covers all of the basic types of questions you should be able to answer about acid-base chemistry… for now…
- Thanks for reading!