 qwe means ⇌ (equil. rxn)

Electrolytes, Acids, and Bases

(Which compounds produce ions when dissolved in water?)

Suggested Demonstration: Electrolytes

ssd means → 100% rxn

Model 1: Electrolytes

Only separate, charged particles (such as ions) can carry electrical currents.

Electrolytes can carry an electrical current when dissolved in water.

Critical Thinking Question:

1. What happens to electrolytes when they dissolve in water?

Electrolytes separate into cations & anions* that can carry electrical current.

* via ion-dipole IMF's

Model 2: Types of electrolytes

In water solution, *strong electrolytes* dissociate completely into ions, *weak electrolytes* dissociate only slightly, and *nonelectrolytes* dissociate undetectably or not at all.

Critical Thinking Question:

2. Describe a method by which you could tell if a particular solution contains a strong electrolyte, weak electrolyte, or nonelectrolyte.

We could use a conductivity meter to indicate the presence of ions. If the light bulb lights, then ions are present.

Model 3: Some common acids and bases

Type of electrolyte	Acids	Bases
Strong (any acid or base not listed here is weak)	HCl HBr HI H_2SO_4 HNO_3 $HClO_4$	LiOH NaOH KOH $Ca(OH)_2$ $Sr(OH)_2$ $Ba(OH)_2$
Weak	$HC_2H_3O_2$ HCN	$Mg(OH)_2$

Acids dissociate in water to give hydrogen (H^+) ions and an anion. Bases dissociate in water to give hydroxide (OH^-) ions and a cation. *Strong* acids and bases dissociate *completely*. Note that the acids are molecules, while the bases are ionic compounds.

Critical Thinking Questions:

3. Consider Model 3. What do all the molecular formulas of the acids have in common?

The molecular formulas begin with H (Hydrogen).

4. How can you recognize an acid from the molecular formula?

~~I~~ An acid's molecular formula begins w/ ~~a~~ Hydrogen.

5. How can you recognize a base from its formula?

The formula of a base can contain hydroxide (OH^{-}). Note: Not ALL bases have hydroxide in their formula.

6. What happens to strong acids when dissolved in water?

They completely dissociate

$HCl \xrightarrow{H_2O} H^{+}_{(aq)}* + Cl^{-}_{(aq)}$

* H_3O^{+} is also correct.

7. What happens to weak acids when dissolved in water?

They partially dissociate.

$HC_2H_3O_2 \rightleftharpoons H^{+} + C_2H_3O_2^{-}$

8. Are all acids strong electrolytes? Explain.

No, weak acids only partially dissociate, so they are weak electrolytes.

9. Describe what happens to the ions in solid sodium hydroxide (NaOH) during the process of dissolving in water.

The Na^{+} & OH^{-} ions become solvated by H_2O & separate.

$NaOH_{(s)} \xrightarrow{H_2O} Na^{+}_{(aq)} + OH^{-}_{(aq)}$

Model 4: Solubilities of some ionic compounds

Ionic compound	Solubility in water	Type of electrolyte	Major species present when dissolved in water
$MgCl_2$	soluble	strong	$Mg^{2+}(aq)$, $Cl^{-}(aq)$
MgO	insoluble	nonelectrolyte	MgO(s)
K_2S	soluble	strong	$K^{+}(aq)$, $S^{2-}(aq)$
CuS	insoluble	nonelectrolyte	CuS(s)
$Ca(NO_3)_2$	soluble	strong	$Ca^{2+}(aq)$, $NO_3^{-}(aq)$
$Ca_3(PO_4)_2$	insoluble	nonelectrolyte	$Ca_3(PO_4)_2(s)$

Molecular compounds other than acids and bases, and *insoluble ionic compounds* do not dissociate in water and so are nonelectrolytes.

The phase label *(aq)*, meaning "aqueous," can be used to show that a species is dissolved in water. **We will not consider how to predict if an ionic compound is soluble in water until later in the course.**

Critical Thinking Questions:

10. A solution of $MgCl_2$ in water could be written as $MgCl_2(aq)$. Besides water, what species would actually be present in the solution? $Mg^{2+}_{(aq)} + 2\,Cl^{-}_{(aq)}$

11. MgO is an ionic compound that does not dissolve in water, it would just collect as a solid—*i. e.,* MgO(s)—at the bottom of the container. Would there be any ions dissolved in the water? Explain.
There may be a few Mg^{2+} & O^{2-} ions dissolved in the H_2O, but not enough to carry an electrical current.

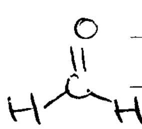
12. Describe a method by which you could tell if an ionic compound is soluble in water.
Add a small amount of the ionic cpd to H_2O & see if it dissolves.

13. Consider Model 4. How is the solubility of an ionic compound related to its classification as an electrolyte?
The greater solubility - the stronger the electrolyte

14. The following compounds from Model 4 are placed in water. Add the phase labels (s) or (aq) to represent whether the compounds are soluble or not:

K_2S (aq) CuS (s) oops :) $Ca(NO_3)_2$ (aq) $Ca_3(PO_4)_2$ (s)

15. Suppose that each compound in the table below is placed in water. The phase labels given describe whether the compound dissolves or not. Fill in the table with the properties for each compound.



Compound	**Acid**, other **molecule**, **base**, or other **ionic compound**	**Strong**, **weak**, or **non**electrolyte
HBr(aq)	acid	strong
KOH(aq)	base	strong
CH_2O(aq)	molecule	non
$Ca_3(PO_4)_2$(s)	ionic	non
$Al(OH)_3$(s)	base	non
HOCl(aq)	acid	weak ~~strong~~
H_2S(aq)	acid	weak
Fe_2O_3(s)	ionic	non
Na_2S(aq)	ionic	strong

Exercises:

1. Write formulas for the major species present in the solutions from CTQ 15.

Compound	Major species present when dissolved in water (or "need more information")
HBr(aq)	(H^+ or H_3O^+) aq $Br^{\ominus}$ (aq)
KOH(aq)	K^+ (aq) $OH^{\ominus}$ (aq)
CH_2O(aq)	CH_2O (aq)
$Ca_3(PO_4)_2$(s)	$Ca_3(PO_4)_2$ (s)
$Al(OH)_3$(s)	$Al(OH)_3$ (s)
HOCl(aq)	(H^+ or H_3O^+) aq $OCl^{\ominus}$ (aq)
H_2S(aq)	(H^+ or H_3O^+) (aq) S^{2-} (aq)
Fe_2O_3(s)	Fe_2O_3 (s)
Na_2S(aq)	Na^+ (aq) S^{2-} (aq)

2. Read the assigned pages in your text, and work the assigned problems.

Acids and Bases

(What happens when hydrogen ions are transferred between species?)

Information: Two definitions of acids and bases

Arrhenius definitions

An acid is a species that dissociates into H^+ (hydrogen) ions and anions when dissolved in water.

A base is a species that dissociates into OH^- (hydroxide) ions and cations when dissolved in water.

Brønsted-Lowry definitions

An acid donates a proton (H^+ ion) to another species.

A base accepts a proton (H^+ ion) from another species.

The Brønsted-Lowry definition explains why the hydrogen ions (H^+) in water are actually hydronium ions—a water molecule (H_2O) has accepted the proton to become hydronium (H_3O^+).

Table 1: Some common acids and bases

Type of electrolyte	Acids	Bases
Strong (all not listed here are weak)	HCl HBr HI H_2SO_4 HNO_3 $HClO_4$	LiOH NaOH KOH $Ca(OH)_2$ $Sr(OH)_2$ $Ba(OH)_2$
Weak	$HC_2H_3O_2$ HCN	$Mg(OH)_2$

Recall that when strong acids and bases dissolve in water, they dissociate completely into ions, while weak acids and bases dissociate only slightly.

Critical Thinking Questions:

1. Write the three species that actually exist in significant amounts in a one-tenth molar aqueous solution of HCl.

   H^+ or H_3O^+ ; Cl^- ; H_2O

2. Explain why hydrogen ion (H^+) is *not* one of the three species in CTQ 1.

   H^+ reacts w/ H_2O to form H_3O^+ (hydronium ions)

3. Write the three species that actually exist in significant amounts in a one-tenth molar aqueous solution of LiOH.

   Li^+ ; OH^- ; H_2O

Information: Hydronium-hydroxide balance

In pure water, a small amount of self ionization occurs, with one water molecule acting as an acid (donating a proton) and another as a base (accepting a proton):

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

In pure water, the concentrations of hydronium ion and hydroxide ion are each 1.0×10^{-7} M. Furthermore, the product of the concentrations of hydronium ion and hydroxide ion in aqueous solution at 25°C is always 1.0×10^{-14} M.

$$[H_3O^+][OH^-] = 1.0\times10^{-14}$$

This means that if the hydronium ion concentration $[H_3O^+]$ increases, the hydroxide ion concentration $[OH^-]$ decreases. This relationship allows us to calculate the amounts of hydronium ion and hydroxide ion in any solution of strong acid or base.

Critical Thinking Questions:

4. What is the hydronium ion concentration in a 1.0×10^{-5} M aqueous solution of HCl?

Since HCl fully ionizes, the $[H_3O^+]$ is 1.0×10^{-5} M
HCl is a strong acid.

5. What is the hydroxide ion concentration in an aqueous solution if the hydronium ion concentration is 1.0×10^{-5} M?

$$[H_3O^+][OH^-] = 1.0\times10^{-14}$$
$$[1.0\times10^{-5}][OH^-] = 1.0\times10^{-14}$$
$$\boxed{[OH^-] = 1.0\times10^{-9}\,M}$$

Information: pH

pH (the "power of hydrogen") is defined as the negative of the logarithm of the molar concentration of hydronium ions (without units):

$$pH = -\log[H_3O^+]$$

Therefore, in pure water, the pH is $-\log(1.0\times10^{-7})$, which equals 7. pH values below 7 are called *acidic*; those above 7 are termed *basic* or *alkaline*. pH values can actually go below 0 and above 14, though this is not commonly seen.

Table 2: The relationship between acidity, pH, and the hydronium and hydroxide ion concentrations of a solution.

Relative Concentrations	pH	Solution
$[H_3O^+] > [OH^-]$	< 7	acidic
$[H_3O^+] < [OH^-]$	> 7	basic
$[H_3O^+] = [OH^-]$	= 7	neutral

Critical Thinking Questions:

6. What is the pH of the 1.0×10^{-5} M aqueous solution of HCl from CTQ 3? (Be sure you can enter this into your calculator correctly, *e. g.*, 1 . 0 EE 5 ± log ±)

$$pH = -\log[H_3O^+] = -\log(1.0\times10^{-5})$$
$$\boxed{pH = 5}$$

$[H_3O^+] > [OH^-]$

7. According to Table 2, which ion does the solution in CTQ 5 contain more of:

hydronium or **hydroxide** (circle one)? $1.0\times10^{-5} > 1.0\times10^{-9}$

8. Does your answer to CTQ 7 agree with your answers to CTQs 4 and 5?

yes $pH = -\log[H_3O^+]$

9. a. If the pH of a cola drink is 3.2, what is the hydronium ion concentration? Be sure you can enter this into your calculator correctly, *e. g.*, 3 . 2 ± 10ˣ (the 10ˣ key is often an inverse or 2nd log).

$[H_3O^+] = 10^{-pH} = 10^{-3.2} = \boxed{6.3\times10^{-4} M}$

b. What is the hydroxide ion concentration in the cola?

$[H_3O^+][OH^-] = 1.0\times10^{-14}$

$(6.3\times10^{-4})[OH^-] = 1.0\times10^{-14}$ $\Rightarrow \boxed{1.6\times10^{-11} M}$

10. a. What is the hydroxide ion concentration in a 1.0×10^{-5} M aqueous solution of **NaOH**?

Since NaOH fully ionizes $\boxed{[OH^-] = 1.0\times10^{-5} M}$

b. What is the pH of this solution? (Careful!)

must solve for $[H_3O^+]$ 1st :)

$[H_3O^+] = \dfrac{1.0\times10^{-14}}{1.0\times10^{-5}} = 1.0\times10^{-9}$

$pH = -\log(1.0\times10^{-9}) = \boxed{9}$

Model: Conjugate acid-base pairs

According to the Brønsted-Lowry theory, a reaction of an acid and a base involves a proton (*i. e.*, hydrogen ion) transfer from the acid to the base. Two ions or molecules that *differ only by that one hydrogen ion* make up a **conjugate acid-base pair**. Three example pairs are shown below:

H_3O^+ and H_2O H_2O and OH^- NH_4^+ and NH_3

The **conjugate acids** have one more proton (H^+) than the **conjugate bases**. For example, consider the reaction below:

conjugate acid-base pair

$CH_3COOH + H_2O$ qwe $CH_3COO^- + H_3O^+$

acetic acid + water acetate + hydronium ion

conjugate acid-base pair

Sometimes we simplify the naming by saying an acid and a base react to give a conjugate acid and conjugate base.

CH_3COOH +	H_2O	qwe	CH_3COO^-	+	H_3O^+
acid	base		conjugate base		conjugate acid

- The conjugate base is what results after the acid gives up a hydrogen ion; so we say that *acetate is the conjugate base of acetic acid*.
- The conjugate acid is what results after the base picks up a hydrogen ion; so we say that *hydronium ion is the conjugate acid of water*.

Critical Thinking Questions:

11. For the equations below, identify the acid and base on the reactant side and the conjugate acid and conjugate base on the product side. Draw a line to connect conjugate acid-base pairs together.

acid base conj acid conj base

a. $HCN(aq) + OH^{-}(aq)$ qwe $H_2O(l) + CN^{-}(aq)$

base acid conj base conj acid

b. $F^{-}(aq) + H_3O^{+}(aq)$ qwe $H_2O(aq) + HF(aq)$

base acid conj acid conj base

c. $NH_3(aq) + HCl(aq)$ qwe $NH_4^{+}(aq) + Cl^{-}(aq)$

Exercises:

1. Which definition of an acid—Arrhenius or Brønsted—is more complete for aqueous solutions? Explain.

   The Brønsted definition is more complete because it describes the transfer of H^+.

2. Calculate the pH of each of the following aqueous solutions.
   a. 1.0×10^{-4} M nitric acid $pH = -\log(1.0\times10^{-4}) =$ a) $\boxed{4}$
   b. 5.0×10^{-3} M hydrobromic acid $pH = -\log(5.0\times10^{-3}) =$ b) $\boxed{2.3}$
   c. a 1.0×10^{-6} M solution of the diprotic acid H_2SO_4 (diprotic means that each molecule of H_2SO_4 donates <u>two</u> hydrogen ions to water molecules) $pH = -\log(2\times1.0\times10^{-6})$ c) $\boxed{5.7}$
   d. 0.0012 M $Ca(OH)_2$ $[OH^-] = 2(.0012) = .0024M$
      $[H_3O^+] = 1.0\times10^{-14} / .0024 = 4.17\times10^{-12}$ d) $\boxed{11}$

3. Calculate the hydroxide ion concentrations of the solutions in Exercise 2.

   a) $[OH^-] = 1.0\times10^{-14} / 1.0\times10^{-4} = \boxed{1.0\times10^{-10} M}$
   b) $[OH^-] = 1.0\times10^{-14} / 5.0\times10^{-3} = \boxed{2.0\times10^{-12} M}$
   c) $[OH^-] = 1.0\times10^{-14} / 2.0\times10^{-6} = \boxed{5.0\times10^{-9} M}$
   d) $[OH^-] = \boxed{0.0024 M}$

4. Write formulas for the conjugate bases of the acids in Exercises 2a and 2b.

   The CB of HNO_3 is $NO_3^{\ominus}$. ← CB
   The CB of HBr is $Br^{\ominus}$.

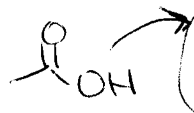
5. Read the assigned pages in your textbook, and work the assigned problems.

Buffers

(How do acids and bases react together?)

Model 1: A buffer system

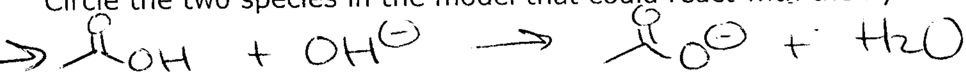
Consider a solution containing both the **weak** acid acetic acid and its **conjugate base**, sodium acetate. This sets up the equilibrium shown below. (The sodium ion is just a spectator, and does not show up in the equilibrium expression.)



CH_3COOH	+	H_2O	qwe	CH_3COO^-	+	H_3O^+
acetic acid (acid)		water (base)		acetate (conjugate base)		hydronium ion (conjugate acid)

Critical Thinking Questions:

1. Suppose some strong base (hydroxide ion) is added to the buffer system in Model 1. Circle the two species in the model that could react with the hydroxide.



see above

$H_3O^{(+)} + OH^{(-)} \rightarrow 2H_2O$

2. Write a chemical equation for the reaction of hydroxide ion being neutralized by reacting with acetic acid, producing acetate and water.

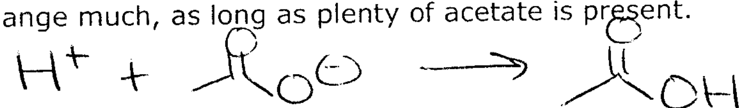
see above

3. Explain why addition of strong base (hydroxide ion) to the solution in Model 1 will not cause a great change in the pH of the solution, as long as the acetic acid is not used up.

There is no significant change in the H_3O^+ concentration.

4. Draw a box around the one species in Model 1 that could react with and neutralize any added hydronium ion.

5. Explain why addition of strong acid will not cause the pH of the solution in Model 1 to change much, as long as plenty of acetate is present.



There is no significant change in the H_3O^+ concentration.

Model 2: A solution of a strong acid and its conjugate base

Consider a solution containing both the **strong** acid hydrochloric acid and its **conjugate base**, sodium chloride. This sets up the system shown below.

$\cancel{HCl}$	+	H_2O	ssd	Cl^-	+	H_3O^+
~~hydrochloric acid (acid)~~		water (base)		chloride (conjugate base)		hydronium ion (conjugate acid)

Critical Thinking Questions:

6. Recalling that strong acids dissociate completely in water, draw a large 'X' through the species in Model 2 that is not present in any significant amount. see previous P3

7. Why is a forward arrow used in Model 2 (ssd) instead of an equilibrium arrow (qwe)?

   The → indicates 100% rxn.

8. Explain the following statement: There is no species in Model 2 that can neutralize added hydronium ion.

   The reverse rxn does not take place b/c Cl⊖ does NOT react w/ H_3O^+

Information: A summary

A solution containing both a **weak** acid and its **conjugate base** is resistant to changes in pH when small amounts of acid or base are added. This solution is called a **buffer**.

A solution of a **strong** acid and its conjugate base is **not** a buffer, since any added strong acid will not be neutralized and will just increase the H_3O^+ concentration. Similarly, a solution of a **strong base** and its conjugate acid is **not** a buffer.

Exercises:

1. State if each solutions would be useful as a buffer or not. Then explain the reason for your choice.

   a. A solution containing 0.08 M NaCN and 0.10 M HCN

      Yes, a buffer sol'n is created
      HCN reacts w/ $OH^{\ominus}$ & $CN^{\ominus}$ reacts w/ $H^{\oplus}$

   b. A solution containing 0.05 M NaOH in H_2O

      Not a buffer.
      There is no substance to neutralize $OH^{\ominus}$

   c. A solution containing 0.25 M HCl and 0.20 M NaCl

      Not a buffer.
      There is no substance to neutralize $H^{\oplus}$

   d. A solution containing 0.05 M NH_4Cl and 0.10 M NH_3

      Yes, a buffer sol'n is created.
      NH_4^+ reacts w/ $OH^{\ominus}$ & NH_3 reacts w/ $H^{\oplus}$

   e. A solution containing 0.20 M KF and 0.15 M HF

      Yes, a buffer sol'n is created.
      $F^{\ominus}$ reacts w/ H^+ & HF reacts w/ $OH^{\ominus}$

2. Considering the solution in Exercise 1b.

   a. Write the three species that would be present in significant amounts.

      $Na^{\oplus}$, $OH^{\ominus}$, H_2O

   b. Is there any species present that can neutralize added hydroxide? Explain.

      No $OH^{\ominus} + H_2O \rightarrow H_2O + OH^{\ominus}$

3. Complete ChemWorksheet 3, Stoichiometry Practice Worksheet 2.

4. Read the assigned pages in your textbook, and work the assigned problems.