I. Electrolytes - must be soluble in water

\* electrolyte - substance which conducts electricity when dissolved in water

- examples; strong: ***HCl(hydrochloric acid), NaCl(salt)***

- examples; weak: ***HC2H3O2(vinegar), NH3(ammonia)***

\* nonelectrolyte - substance which does not conduct electricity when dissolved in water

\* why do electrolytes do this? ***When an ionic compound dissolves in water, the ions dissociate.***

***As a result, you get free-flowing ions in the aqueous solution (see picture below)***

II. Ionization vs. Dissociation

A. Dissociation - process through which ionic solids separate into free flowing ions in water solution

\* remember pirhana-cow theory?

\* In order for ions to be by themselves, they must be in aqueous solution

 

Give the dissociation equation for:

**T**.

1. NaCl(s)  **Na+1(aq) + Cl-1(aq)**

2. Na2SO4(s) **2 Na+1(aq) + SO4-2(aq) \*Note the sulfate does not split**

**S**.

1. (NH4)3PO4(s) **3 NH4+1(aq) + PO4-3(aq)**

2. Pb(NO3)2(s) **Pb+2(aq) + 2 NO3-1(aq)**

B. Ionization - formation of ions caused by the reaction of a molecular compound with water

\* example: HC2H3O2(aq) + H2O(l)  C2H3O2-1(aq) + H3O+1(aq)

- hydronium ion **H3O+1**



III. Arrhenius Acid - molecular substance which produces H+ (or H3O+) ions when it reacts with water

Give the ionization equation for:

**T**.

1. HCl(g) **H+ + Cl-1** 2. H2CO3(aq) **H+ + HCO3-1**

\****Note only one hydrogen (H+) comes off***

***in this process***

3. HCO3-1(aq) **H+ + CO3-2 *\*Note that no (aq) is required; all ions by definition can only***

***exist in aqueous solution***

**S**.

1. HNO3(aq) **H+ + NO3-1** 2. H2S(aq) **H+ + HS-1**

3. HS-1(aq) **H+ + S-2**

IV. Naming Acids Review

**hydrochloric acid** 1) HCl **H2S**  4) hydrosulfuric acid

**sulfuric acid** 2) H2SO4 **HNO3**  5) nitric acid

**sulfurous acid** 3) H2SO3 **HNO2**  6) nitrous acid

V. Properties of Acids

A. Molecular substances which ionize when added to water to form H3O+1 ions

\* therefore all are electrolytes

B. React with active metals to form H2(g)

**T**.

1.  **2**  Na(s) +  **2**  HCl(aq) 🡪 **2 NaCl + H2** *\*****Note the single-replacement pattern (see unit VI)***

**S**.

1.  **1**  Zn(s) +  **1**  H2SO4(aq) 🡪 **ZnSO4 + H2**

2.  **2**  Al(s) +  **6**  HNO3(aq) 🡪 **2 Al(NO3)3 + 3 H2**

C. Acids affect the colors of indicators

D. Acids neutralize bases

E. Dilute acids taste sour – ***the sour taste of a lemon is due to the citric acid, for example***

**\*\*SAFETY TIP: Acids release tremendous amounts of heat when you dilute them (esp. H2SO4)**

**🡪 ALWAYS ADD ACID TO WATER**

VI. Strength of Acids and Ka

HC2H3O2(aq)  H+1(aq) + C2H3O2-1(aq) 

\* strong acid - make lots of H+ ions, Ka >> 1

\* weak acid - make a few ions, Ka<< 1

\* *diprotic* acid has 2 H’s, examples? ***H2CO3, H2SO4***

\* *triprotic* acid has 3 H’s, examples? ***H3PO4***

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Acid** | **Conjugate Base** | **Ka** | **Acid** | **Conjugate Base** | **Ka** |
| HI 🡪 | H+1 + I-1 | very large | H2CO3 🡪 | H+1 + HCO3-1 | 4.3 x 10-7 |
| HBr 🡪 | H+1 + Br-1 | very large | HSO3-1🡪 | H+1 + SO3-2 | 1.1 x 10-7 |
| HCl 🡪 | H+1 + Cl-1 | very large | H2S 🡪 | H+1 + HS-1 | 9.5 x 10-8 |
| HNO3🡪 | H+1  + NO3-1 | very large | H2PO4-1 🡪 | H+1 + HPO4-2 | 6.2 x 10-8 |
| H2SO4🡪 | H+1 + HSO4-1 | large | NH4+1 🡪 | H+1 + NH3 | 5.7 x 10-10 |
| H2SO3 🡪 | H+1 + HSO3-1 | 1.5 x 10-2 | HCO3-1 🡪 | H+1 + CO3-2 | 5.6 x 10-11 |
| HSO4-1 🡪 | H+1 + SO4-2 | 1.2 x 10-2 | HPO4-2 🡪 | H+1 + PO4-3 | 2.2 x 10-13 |
| H3PO4 🡪 | H+1 + H2PO4-2 | 7.5 x 10-3 | HS-1 🡪 | H+1 + S-2 | 1.3 x 10-14 |
| HF 🡪 | H+1 + F-1 | 6.3 x 10-4 | H2O🡪 | H+1 + OH-1 | 1.0 x 10-14 |
| HNO2 🡪 | H+1 + NO-1 | 5.6 x 10-4 | OH-1 🡪 | H+1 + O-2 | < 10-36 |
| HC2H3O2 🡪 | H+1 + C2H3O2-1 | 1.8 x 10-5 |  |  |  |

***\* Note that this table has the acids ranked from strongest to weakest***

Calculate the [H+1] in a:

**T**.

1) 1.00 M HCl solution – ***because HCl is a strong acid, it will, for all practical purposes dissociate 100%, as a result the concentration of hydrogen ions (H+) will equal the concentration of the acid***

**(H+) = 1.00 M**

2) 2.00 M H2SO4 solution – ***strong acid***

**(H+) = 2.00 M**

3) 1.35 M HNO2 solution – ***weak acid; when a weak acid ionizes, we can calculate the concentration of hydrogen***

***ions by using the Ka (see chart on p.3); since the ionization equation is: HNO2🡪 H+ + NO2-***

***we know the concentration of hydrogen ions will equal the concentration of nitrites (NO2-), so both will be assigned the value of x.. Also, because the acid is a weak acid, we know that the number of HNO2’s that ionize is negligible, so at equilibrium, the concentration of HNO2’s will still be roughly 1.35 M.***

   



**S**.

1) 2.50 M HNO3 solution – ***strong acid***

**(H+) = 2.50 M**

2) 0.400 M HF solution

  

3) 0.250 M HC2H3O2

  

4) 0.0441 M HNO2

  

VII. Base - ionic substance which dissociates to form OH- ions in water

\* examples ***NaOH(lye), Ca(OH)2 (limewater)***

Naming Review: name the following Arrhenius bases

**S**.

1. NaOH – ***sodium hydroxide***

2. Mg(OH)2 – ***magnesium hydroxide***

3. aluminum hydroxide – ***Al(OH)3***

4. ammonium hydroxide – ***NH4OH***

VIII. Properties of Bases - often referred to as *caustic* or *alkaline* substances

A. Bases are electrolytes - dissociate in water to form OH-1.

B. Bases affect the colors of indicators.

C. Neutralize acids.

D. Water solutions are bitter and slippery.

E. Emulsify fats and oils

IX. Salt – any ionic compound that does not contain OH-1.

\* all are good electrolytes

\* formed by a neutralization reaction – ***this reaction follows the same pattern as double replacement***

***(see Unit VI)***

**T**.

1.  **1** HCl(aq) +  **1**  NaOH(aq) 🡪 **NaCl(salt) + HOH (water)**

2.  **1**  H2SO4(aq) +  **2** KOH(aq) 🡪 **K2SO4(salt) + 2 HOH**

**S**.

1.  **2**  HBr(aq) +  **1**  Ca(OH)2(aq) 🡪 **CaBr2 + 2 HOH**

2.  **1**  HC2H3O2(aq) +  **1**  NaOH(aq) 🡪 **NaC2H3O2 + HOH**

Acid, Base, Salt, or Neither:

**T**.

1. NaCl - ***salt*** 2. KCl - ***salt*** 3. KOH - ***base*** 4. SO2 - ***neither*** 5. NH4C2H3O2 - ***salt***

**S**.

1. KBr – ***salt*** 2. H2SO4 - ***acid*** 3. HgCl2 - ***salt*** 4. Al(OH)3 – ***base*** 5. HCl - ***acid***

6. KOH - ***base*** 7. CaO - ***salt*** 8. K3PO4 - ***salt*** 9. CO2 - ***neither*** 10. NH4OH - ***base***

X. Another look at Acids and Bases



\* operational definition - based directly on observable

properties

\* conceptual definition - based on interpretation of observed

facts

XI. Lowry-Bronsted Conceptual Definition of Acids and Bases

\* acid - proton donor; proton is also an H+1 ion

\* bases - proton acceptor

Label the Lowry-Bronsted acids and bases:

**T**,

1. HCl(aq) + H2O(l)  H3O+1(aq) + Cl-1(aq)

**A B A B**

**\* Note that HCl is the acid because it donates a proton to become**

**Cl-1. Similarly, Cl-1 accepts a proton to become HCl, so it is the base.**

2. NH3(aq) + H2O(l)  NH4+1(aq) + OH-1(aq)

**B A A B**

**S**.

1. HNO3(aq) + H2O(l)  NO3-1(aq) + H3O+1

**A B B A**

2. HC2H3O2(aq) + OH-1(aq)  C2H3O2-1(aq) + H2O(l)

**A B B A**

XII. Conjugate Acid-Base Pairs - pairs within a Lowry-Bronsted acid/base reaction that differ by one proton

Name the conjugate pairs from the above reactions:

**T**.

1. ***HCl, Cl-1 H3O+1, H2O*** 2. ***NH4+1, NH3 H2O, OH-1***

**S**.

1. ***HNO3, NO3-1 H3O+1, H2O*** 2. ***HC2H3O2, C2H3O2-1 H2O, OH-1***

\* amphoteric substance - acts as both an acid and a base in different situations; examples from

above?

***\* Water (H2O) is an amphoteric substance because it acts as a base in T#1, and***

***an acid in T#2***

XIII. Self-Ionization of Water - water does ionize a little

H2O(l)  H+1(aq) + OH-1(aq) Kw = [H+1][OH-1] = 1.00 x 10-14 at 25oC

**T**. What is [H+1] if:

1. [OH-1] = 1.00 x 10-3 M - ***Note that because (H+)<(OH-), we would classify this solution as basic***

 

2. [OH-1] = 3.61 x 10-10 M - ***(H+)>(OH-), so we would call this solution acidic***

 

**S**. What is [OH-1] if:

1. [H+1] = 1.00 x 10-7 M – ***this solution will be neutral because (H+) = (OH-)***

 

2. [H+1] = 5.09 x 10-2 M

 

\* How does this relate to LeChatlier’s Principle? ***As [H+] increases, [OH-1] must decrease in order***

***For Kw to remain constant***

XIV. pH - an easier way of expressing hydrogen concentration

|  |  |
| --- | --- |
| **pH** | **Acidity** |
| 7 | neutral |
| less than 7 | acidic |
| greater than 7 | basic |

pH = -log[H+1]



What is the pH if:

**T**.

1) [H+1] = 1.31 x 10-5 M 2) [H+1] = 1.31 x 10-6 M

pH = -log (H+) = -log (1.31 x 10-5) = **4.88** pH = -log (H+) = -log (1.31 x 10-6) = **5.88**

***\*Note this shows that the lower the pH number, the higher the concentration of hydrogens, therefore, the more acidic the solution is. The solution in T#1 is more acidic than the solution in T#2***

3) you dissolve 2.61 L of HCl gas in 5.00 L of water?

***\*( ) means concentration in M, so we need to find that first before doing the pH***

HCl is a strong acid, so (H+) = (HCl) = 0.0233M

pH = -log(H+) = -log (0.0233) = **1.63**

4) you dissolve 25.0 g of HC2H3O2 in 25θ mL of water?

HC2H3O2 is a weak acid, so need to do weak acid calculation (see p. 3)

 

pH = -log (0.0055) = **2.3**

**S**.

1) [H+1] = 6.02 x 10-7 M 2) [H+1] = 6.31 x 10-11 M

pH = -log(6.02x10-7) = **6.22** pH = -log (6.31 x 10-11) = **10.2**

3) you dissolve 5.00g of HNO3 in 2.50 liters of water? – ***strong acid***



pH = -log (0.03716) = **1.50**

4) you dissolve 16.0 g of HF in 15θ mL of water? – ***weak acid***



  pH = -log(0.0580) = **1.2**

XV. pOH

|  |  |
| --- | --- |
| **pOH** | **Acidity** |
| 7 | neutral |
| less than 7 | basic |
| greater than 7 | acidic |

pOH = -log[OH-1] pH + pOH = 14

**T**.

1) What is the pOH if [OH-1] = 3.71 x 10-4 M?

pOH = -log(3.71 x 10-4) = **3.43**

2) What is the pOH if you dissolve 3.61g of NaOH(s) in 2.00 L of water?



pOH = -log (.0451) = **1.35**

3) What is the pH of (T)#1? pH = 14 – pOH = 14 – 3.43 = **10.57**

4) What is the pH of (T)#2? pH = 14 – 1.35 = **12.65**

**S**.

1) What is the pOH if [OH-1] = 6.42 x 10-8 M?

pOH = -log(6.42 x 10-8) = **7.19**

2) What is the pOH if you dissolve 1.50g of KOH(s) in 573 mL of water?



pOH = -log (.0467) = **1.33**

3) What is the pH of (S)#1? pH = 14 – 7.19 = **6.81**

4) What is the pH of (S)#2? pH – 14 – 1.33 = **12.67**

XVI. Buffer - a solution which is able to resist major changes in pH

*weak acid*

\* example: HC2H3O2(aq)  H+1(aq) + C2H3O2-1(aq)

- *common-ion effect* - by adding a salt with the negative ion (NaC2H3O2 or KC2H3O2) we increase

the concentration of that ion, therefore: ***we have a solution containing a weak acid***

***(HC2H3O2) and its conjugate base (C2H3O2-1)***

\* *add H+:* ***the hydrogens in the acid will react with the conjugate base (C2H3O2-1);***

***the base neutralizes the added acid:***

***H+1 + C2H3O2-1 🡪 HC2H3O2***

\* *add OH-1:* ***the hydroxides will react with the weak acid (HC2H3O2);***

***the acid neutralizes the added base***

***OH-1 + HC2H3O2 🡪 C2H3O2-1 + HOH***

\* biological example: carbonic acid/bicarbonate in blood

***There is a balance between the ions which acts as a buffer, keeping the pH of the blood***

***right around 7.4. The hemoglobin molecule in red blood cells can only withstand pH***

***extremes of 7.2-7.6***

XVII. Acid-Base Indicators - chemicals specifically designed to show specific colors in acids and different colors

in bases

\* most are weak acids:

HIn  H+ + In-1

acid color base color

\* How does it work? ***Much like a buffer, in acidic solution, the base (In-) reacts with the***

***acid to show the acid color, whereas in basic solution, the acid(HIn)***

***will react with the base to show the base color***

|  |  |  |  |
| --- | --- | --- | --- |
| **Indicator** | **pH Range** | **below pH color** | **above pH color** |
| methyl violet | 0.0 – 1.6 | yellow | blue |
| methyl yellow | 2.9 – 4.0 | red | yellow |
| bromophenol blue | 3.0 – 4.6 | yellow | blue |
| methyl orange | 3.2 – 4.4 | red | yellow |
| methyl red | 4.8 – 6.0 | red | yellow |
| litmus | 5.5 – 8.0 | red | blue |
| bromothymol blue | 6.0 – 7.6 | yellow | blue |
| phenol red | 6.6 – 8.0 | yellow | red |
| phenolphthalein | 8.2 – 10.6 | colorless | red |
| thymolphthalein | 9.4 – 10.6 | colorless | blue |
| alizarin yellow | 10.0 – 12.0 | yellow | red |

XVIII. Acid-Base Neutralization

H+1 + OH-1  H2O

\* if you have 35 molecules of acid, 35 molecules of base will neutralize it

\* equivalence point - when an equivalent amount of OH-1 ions has been added to H+1 ions; it’s

“neutralized”

- strong acid - strong base – ***equivalence point = 7***

\* good indicators? ***litmus, bromothymol blue, phenol red (pH range overlaps 7)***

- strong acid - weak base – ***equivalence point < 7***

\* good indicators? ***methyl yellow, bromophenol blue, methyl orange***

- weak acid - strong base – ***equivalence > 7***

\* good indicators? ***phenolphthalein, thymolphthalein, alizarin yellow***

XIX. Acid-Base Titration - lab procedure used to determine the concentration of an unknown acid or base solution.

\* standard solution – solution whose concentration is known

\* unknown solution – solution whose concentration you are trying to determine

MaVa = MbVb

\* The above formula also works for dilution of acids (MiVi = MfVf) - see Unit XI

**T**.

1) If you begin a titration with 20.0 mL of unknown HCl and titrate it to the equivalence point using 35.6 mL of

0.600 M standard NaOH, what is the concentration of HCl?

Ma(20.0 mL) = (0.600M)(35.6 mL) **Ma = 1.07 M**

2) If you titrate 65.0 mL of an unknown NH3 solution to the equivalence point with 31.2 mL of a 1.50 M HCl solution, what is the concentration of the ammonia?

(1.50M)(31.2mL) = Mb(65.0mL) **Mb = 0.720 M**

**S**.

1) Ma = ??? Va = 50.0 mL Mb = 1.50 M Vb = 71.3 mL

Ma(50.0mL) = (1.50M)(71.3mL) **Ma = 2.14 M**

2) What is the concentration of an unknown NaOH solution if you titrate 100.0 mL of it to the equivalence point with 43.5 mL of 6.0 M HCl?

(6.0M)(43.5mL) = Mb(100.0mL) **Mb = 2.6 M**

3) What is the concentration of a vinegar (HC2H3O2) solution if you titrate exactly 20 drops of it to the equivalence

point with 26 drops of 0.600M NaOH?

Ma(2θdr) = (0.600M)(26dr) **Ma = 0.78 M**

4) How much water must be added to make 500.0 mL of a 0.100 M HCl solution from concentrated HCl if its con-

centration is 12 M?

(0.100M)(500.0mL) = (12M)V2 \* **Note this is the dilution formula**

**(Unit XI)**

V2 = 4.2 mL 🡪 amount of 12M HCl needed

Volume water added = 500.0 mL – 4.2 mL = **495.8 mL**