Worksheet 18 - Acids and Bases

The **Brønsted-Lowry** definition of an acid is a substance capable of **donating** a **proton** (H^+), and **a base** is a substance capable of **accepting** a **proton**. For example, the weak acid, **HF**, can be dissolved in water, giving the reaction:

$$\underset{\text{acid}}{HF\ (aq)} + \underset{\text{base}}{H_2O\ (l)} \leftrightarrows \underset{\text{conjugate acid}}{H_3O^+\ (aq)} + \underset{\text{conjugate base}}{F^-\ (aq)}$$

In this reaction, HF is the species losing the proton (H^+), making it the **acid**. Water is the species accepting the proton, to form the **hydronium ion**, H_3O^+, making it the **base.** The F^- (aq) is called the **conjugate base** of HF. It can gain a proton in the reverse reaction. H_3O^+ is the conjugate acid of H_2O, since it can lose a proton in the reverse reaction. The stronger an acid, the weaker its conjugate base will be and the stronger the base, the weaker its conjugate acid.

The equilibrium concentrations of these species will be determined by the relative strengths of the acids and bases. The strongest acid will dissociate to the greatest extent. H_3O^+ (H^+) is the strongest acid that can exist in an aqueous system. So, the equilibrium in this system will favor the reactants, HF and H_2O, the weaker acid and base. The equilibrium state is described by an equilibrium constant, $\boldsymbol{K_a}$, in the case of acids, and $\boldsymbol{K_b}$ in the case of bases. These are related by the expression $\boldsymbol{K_w} = \boldsymbol{K_a} \times \boldsymbol{K_b} = 1 \times 10^{-14}$.

1. Classify the following as Brønsted acids, bases or both.

If the compound has a hydrogen, it can be an acid; if it has any lone pairs, it can be a base

a) H_2O	b) OH^-	c) NH_3
both	**both**	**both**
d) NH_4^+	e) NH_2^-	f) CO_3^{2-}
acid	**both**	**base**

2. What is the conjugate base of the following acids?

a) $HClO_4$	b) NH_4^+	c) H_2O	d) HCO_3^-
ClO_4^-	**NH_3**	**OH^-**	**CO_3^{2-}**

(Remove one hydrogen and reduce the charge by 1)

3. What is the conjugate acid of the following bases?

a) CN^-	b) SO_4^{2-}	c) H_2O	d) HCO_3^-
HCN	**HSO_4^-**	**H_3O^+**	**H_2CO_3**

(Add one hydrogen and increase the charge by 1)

4. HSO_3^- is **amphoteric**; it can behave as either an acid or a base. In the examples above, you may have noticed that H_2O and HCO_3^- can also exist as acids or bases.

a) Write the equation for the reaction of HSO_3^- with H_2O in which it acts like an acid and identify the acid-base pairs. Circle the strongest acid. Then write an expression for K_a. The value of K_a at 25° C is 1.23 x 10^{-7}. Use an arrow to indicate which side is favored by the equilibrium

$$\mathbf{HSO_3^- + H_2O \leftrightharpoons SO_3^{2-} + H_3O^+}$$

$$K_a = \frac{[SO_3^{2-}][H_3O^+]}{[HSO_3^-]} = 1.23\times10^{-7}$$

$$\longleftarrow$$

b) Write the equation for the reaction of HSO_3^- with water in which it acts like a base and identify the acid-base pairs. Circle the strongest base. Then write an expression for K_b. What is the value of K_b at 25° C, given that the K_a for H_2SO_3 = 1.58 x 10^{-2}? Use an arrow to indicate which side is favored by the equilibrium.

$$\mathbf{HSO_3^- + H_2O \leftrightharpoons H_2SO_3 + OH^-}$$

$$K_b = \frac{[H_2SO_3][OH^-]}{[HSO_3^-]}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{1.58\times10^{-2}} = 6.33\times10^{-13}$$

$$\longleftarrow$$

c) If HSO_3^- is placed in water, is the resulting solution acidic or basic?
$K_a > K_b$, so HSO_3^- will produce more H_3O^+ than OH^- and the solution overall will be acidic

d) Compute a value for the equilibrium constant for the reaction shown below:

$SO_3^{2-} + H_3O^+ \leftrightharpoons HSO_3^- + H_2O$ $\quad$ ***K*** = <u>**8.13x10^6**</u>

$$K = \frac{[HSO_3^-]}{[SO_3^{2-}][H_3O^+]} = \frac{1}{K_a \text{ from part a}} = \frac{1}{1.23\times10^{-7}} = 8.13\times10^6$$

Circle the strongest acid and indicate which side is favored by the equilibrium.

$$\longrightarrow$$

5. Identify the conjugate acid-base pairs in the following reactions. Circle the strongest acid and indicate which side is favored at equilibrium:

a) $NH_4^+ + H_2O \leftrightarrows NH_3 + H_3O^+$

NH_4^+ / NH_3 **(acid/base)**
H_2O / H_3O^+ **(base/acid)**

b) $HCOOH + CN^- \leftrightarrows HCOO^- + HCN$

K_a HCOOH = 1.8×10^{-5}
K_a HCN = 6.2×10^{-10}

HCOOH / $HCOO^-$ **(acid/base)**
CN^- / HCN **(base/acid)**

c) $NH_4^+ + CO_3^{2-} \leftrightarrows NH_3 + HCO_3^-$

K_b NH_3 = 1.8×10^{-5}
K_b CO_3^{2-} = 1.8×10^{-4}

NH_4^+ / NH_3 **(acid/base)**
CO_3^{-2} / $HCOO^-$ **(base/acid)**

d) $H_2PO_4^- + OH^- \leftrightarrows HPO_4^{2-} + H_2O$

$H_2PO_4^-$ / HPO_4^{2-} **(acid/base)**
OH^- / H_2O **(base/acid)**

6. Name 6 strong acids. Name 6 strong bases.

Strong acids: HCl, HBr, HI, HNO_3, $HClO_4$, $HClO_3$, H_2SO_4

Strong bases: LiOH, NaOH, KOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

7. Complete the following table, which deals with pH, $[H^+]$, acidity, neutrality and basicity.

pH	**$[H^+]$**	**solution is:**
< 7	> 1.0×10^{-7}	acidic
> 7	< 1.0×10^{-7}	basic
7	= 1.0×10^{-7}	neutral

8. Calculate the pH of the solution formed when 100 mL of 1.00 M HCl is added to1.00 L of water.

$$0.100\ \mathrm{L} \times \frac{1.00\ \mathrm{mol\ HCl}}{1\ \mathrm{L}} \times \frac{1\ \mathrm{mol\ H^+}}{1\ \mathrm{mol\ HCl}} = 0.100\ \mathrm{mol\ H^+}$$

$$\left[H^+\right] = \frac{0.100\ \mathrm{mol\ H^+}}{1.00 + 0.100\ \mathrm{L}} = 0.0909\ \mathrm{M}$$

$$\mathrm{pH} = \text{-}\log(0.0909) = 1.04$$

9. Calculate the pH of the following solutions of strong acids and bases:

a) 0.001 M HCl

$$\left[H^+\right] = 0.001\ \mathrm{M}$$

$$\mathrm{pH} = \text{-}\log(0.001) = 3$$

b) 0.76 M KOH

$$\left[OH^-\right] = 0.76\ \mathrm{M}$$

$$\mathrm{pOH} = \text{-}\log(0.76) = 0.119$$

$$pH = 14 - 0.119 = 13.9$$

c) 2.8×10^{-4} M $Ba(OH)_2$ (Think carefully about this one!)

When $Ba(OH)_2$ dissociates, two OH^- ions are produced, so the total concentration of OH^- is double the concentration of $Ba(OH)_2$

$$\left[OH^-\right] = 2\left(2.8 \times 10^{-4}\ \mathrm{M}\right) = 5.6 \times 10^{-4}\ \mathrm{M}$$

$$\mathrm{pOH} = \text{-}\log\left(5.6 \times 10^{-4}\right) = 3.25$$

$$pH = 14 - 3.25 = 10.75$$