**Acid-Base Theories Packet** Name____________________________________ Date______ Period______

**Arrhenius Acid-Base Theory**

The Arrhenius definition of acids and bases is one of the oldest. An **Arrhenius acid** is a substance that when added to water increases the concentration of $H^+$ ions present. The chemical formulas of Arrhenius acids are written with the acidic hydrogen first (HCl, $H_2SO_4$, $H_3PO_4$, etc.). An **Arrhenius base** is a substance that when added to water increases the concentration of $OH^-$ ions present. (NaOH, KOH, LiOH are examples of an Arrhenius bases.)

if the reaction happens in water you only form the hydrogen ion.

Examples of Arrhenius Acid ($H^+$): $HCl_{(g)} \xrightarrow{H_2O_{(l)}} H^+_{(aq)} + Cl^-_{(aq)}$

Examples of Arrhenius Acid ($H_3O^+$): $HCl_{(g)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

if water is a reactant, you will form hydronium as a product.

Example of Arrhenius Base: $NaOH_{(s)} \xrightarrow{H_2O_{(l)}} Na^+_{(aq)} + OH^-_{(aq)}$

The $H^+$ ion produced by an Arrhenius acid is always associated with a water molecule to form the hydronium ion, $H_3O^+_{(aq)}$. Arrhenius acids are frequently referred to as proton donors, hydrogen ion ($H^+$) donors, or hydronium ion ($H_3O^+$) donors, depending on whether we are trying to emphasize the species liberated by the acid (proton or hydrogen ion) or the species present in solution (hydronium ion). To represent the transfer of the $H^+$ ion to water to form the hydronium ion, we must include $H_2O$ in the chemical equation for acid ionization.

**Instructions: Tell if the following are Arrhenius acids or bases and predict the products of the reaction:**

1. ________________ $CH_3COOH_{(s)} \xrightarrow{H_2O_{(l)}}$

2. ________________ $KOH_{(aq)} \xrightarrow{H_2O_{(l)}}$

3. ________________ $HClO_{4\,(aq)} \xrightarrow{H_2O_{(l)}}$

4. ________________ $H_2SO_{4\,(aq)} \xrightarrow{H_2O_{(l)}}$

5. ________________ $RbOH_{(aq)} \xrightarrow{H_2O_{(l)}}$

6. ________________ $NaOH_{(aq)} \xrightarrow{H_2O_{(l)}}$

**If the reactants (left side of the reaction) include water, then you will want to include the hydronium ion ($H_3O^+$) on the product side instead of the hydrogen ion ($H^+$).**

Look back at intro.

7. ________________ $HBr_{(s)} + H_2O_{(l)} \rightarrow$

8. ________________ $HF_{(g)} + H_2O_{(l)} \rightarrow$

9. ________________ $HCl_{(g)} + H_2O_{(l)} \rightarrow$

10. ________________ $HNO_{3\,(s)} + H_2O_{(l)} \rightarrow$

## Brønsted-Lowry Acids & Bases

### B-L Acids

With the Brønsted-Lowry concept we usually refer to a hydrogen ion as a **proton**. That is because a proton is all that is left when a **hydrogen atom** loses an electron to become an **ion**.

Brønsted and Lowry independently came up with the idea that an acid is an acid because it **provides or donates a proton** to something else. When an acid reacts, the proton is **transferred** from one chemical to another. As will be noted later, the chemical which accepts the proton is a base. When an acid dissolves and dissociates in water it gives a proton to the water.

Equations to represent this are shown here. The Brønsted-Lowry view is that the acid gives a proton to water to make two ions, one of which is $H_3O^+$. $H_3O^+$ is called **hydronium ion**. These equations show a different acid ($H_2SO_4$) giving a proton to water. In this case, the product $HSO_4^-$ still has a proton that can be donated to another water molecule.

donates a proton.

conj. Base conj. Acid

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$$

B-L Acid B-L Base (Accepts a proton.)

it will now give the second $H^+$

Conj. Base Conj Acid

$$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$$

B-L Acid B-L Base (Accepts a proton.)

This equation shows HCl giving a proton to a hydroxide ion ($OH^-$) rather than water. The first chemical in each of these equations is an acid because they are each giving a proton to something else.

donates proton

conj. Base conj. Acid

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

B-L Acid B-L Base

donates a proton

conj. acid conj. Base

$$HCl + OH^- \rightarrow H_2O + Cl^-$$

B-L Acid B-L Base

**B-L Bases**

Note that in order for an acid to act like an acid, there needs to be something for it to react with. *There needs to be something to take the proton.* There needs to be a **base**. In the B-L Acid-Base theory, a base is a **proton acceptor**. Compare this to the definition that an acid is a proton donor. Bases are the opposite of acids. Bases are basic because they take or **accept protons**. Hydroxide ion, ($OH^-$) for example can accept a proton to form water. Brønsted and Lowry realized that not all bases had to have a hydroxide ion. As long as something can accept a proton it is a base. *So anything, hydroxide or not, that can accept a proton is a base under the Brønsted-Lowry definition.* The water molecules that accept protons when HCl dissolves in water are acting as bases.

$$OH^- + H^+ \rightarrow H_2O \qquad H_2O + HCl \rightarrow H_3O^+ + Cl^-$$

Some additional examples of Brønsted-Lowry bases are shown accepting protons in these equations. These examples do not show the acids which are providing the protons. Ammonia can accept or react with hydrogen ion to give ammonium ion $NH_4^+$. Ammonia can accept or react with hydrogen ion to give ammonium ion $NH_4^+$. Carbonate ion can accept a hydrogen ion, or accept a proton, to become bicarbonate ion. Also, water molecules, as mentioned before, can act as a base by accepting protons.

$$OH^- + H^+ \rightarrow H_2O \qquad NH_3 + H^+ \rightarrow NH_4^+$$

$$CO_3^{2-} + H^+ \rightarrow HCO_3^- \qquad H_2O + H^+ \rightarrow H_3O^+$$

Hydroxide, ammonia, carbonate and water are all Brønsted-Lowry bases. Be sure to note the distinction between **ammonia** and **ammonium**. $NH_3$ is ammonia and $NH_4^+$ is ammonium. They sound very much the same and their formulas are very similar, but their chemical properties are quite different. They are different because one has one more proton than the other. Ammonia is a base and ammonium is an acid. We'll take up another aspect of their relationship when we consider conjugate pairs. Some phenomena that are readily explained using the Brønsted-Lowry concept are acid-base reactions (explained as **proton transfer reactions**), **conjugate pair** relationships, and **amphoterism**. *(Remember, a substance that is amphoteric can act as an acid or a base.)*

**Brønsted-Lowry Acids & Bases**

Identify the acid, base, conjugate acid and conjugate base for each of the following:

a) $HClO_{4(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + ClO_4^-{}_{(aq)}$

b) $H_2SO_{3(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + HSO_3^-{}_{(aq)}$

c) $HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$

d) $H_2S_{(g)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + HS^-_{(aq)}$

e) $HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + SO_3^{2-}{}_{(aq)}$

## BRONSTED-LOWRY ACIDS AND BASES

Name ____________________

According to Bronsted-Lowry theory, an acid is a proton ($H^+$) donor, and a base is a proton acceptor.

$H^+$

**Example:** $HCl + OH^- \rightarrow Cl^- + H_2O$

**The HCl acts as an acid, the $OH^-$ as a base. This reaction is reversible in that the $H_2O$ can give back the proton to the $Cl^-$.**

Label the Bronsted-Lowry acids and bases in the following reactions and show the direction of proton transfer.

$H^+$ $H^+$

**Example:** $H_2O + Cl^- \leftrightarrow OH^- + HCl$

Acid Base Base Acid

1. $H_2O + H_2O \leftrightarrow H_3O^+ + OH^-$

2. $H_2SO_4 + OH^- \leftrightarrow HSO_4^- + H_2O$

3. $HSO_4^- + H_2O \leftrightarrow SO_4^{-2} + H_3O^+$

4. $OH^- + H_3O^+ \leftrightarrow H_2O + H_2O$

5. $NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

## CONJUGATE ACID-BASE PAIRS

Name ____________________

In the exercise, Bronsted-Lowry Acids and Bases, it was shown that after an acid has given up its proton, it is capable of getting back that proton and acting as a base. Conjugate base is what is left after an acid gives up a proton. The stronger the acid, the weaker the conjugate base. The weaker the acid, the stronger the conjugate base.

Fill in the blanks in the table below.

### Conjugate Pairs

| | ACID | BASE | EQUATION |
|---|---|---|---|
| 1. | $H_2SO_4$ | $HSO_4^-$ | $H_2SO_4 \leftrightarrow H^+ + HSO_4^-$ |
| 2. | $H_3PO_4$ | | |
| 3. | | $F^-$ | |
| 4. | | $NO_3^-$ | |
| 5. | $H_2PO_4^-$ | | |
| 6. | $H_2O$ | | |
| 7. | | $SO_4^{-2}$ | |
| 8. | $HPO_4^{-2}$ | | |
| 9. | $NH_4^+$ | | |
| 10. | | $H_2O$ | |

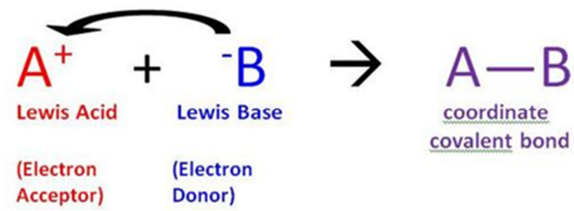
Which is a stronger base, $HSO_4^-$ or $H_2PO_4^-$? ____________________

Which is a weaker base, $Cl^-$ or $NO_2^-$? ____________________

## Lewis Acids & Bases

The way to define a Lewis acid and base is related to the donation and acceptance of an electron pair:

- Lewis Acid: an electron acceptor.
- Lewis Base: an electron donor.

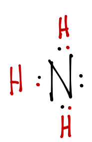



A Lewis acid is any atom, ion, or molecule which can accept electrons and a Lewis base is any atom, ion, or molecule capable of donating electrons. It turns out that it may be more accurate to say that Lewis acids are substances which are **electron-deficient** (or low electron density) and Lewis bases are substances which are **electron-rich** (or high electron density).

Several categories of substances can be considered Lewis acids: 1) positive ions, 2) having less than a full octet in the valence shell, 3) polar double bonds (one end), and 4) expandable valence shells. Several categories of substances can be considered Lewis bases: 1) negative ions, 2) one of more unshared pairs in the valence shell 3) polar double bonds (the other end), and 4) the presence of a double bond .

**Identify the nature of each of the following as either a Lewis Acid or a Lewis Base:**

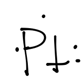
1. $NH_3$



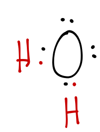
2. $Ag^+$

3. $Ni^{2+}$

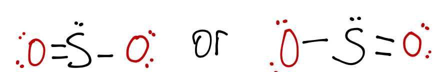
4. $Pt^{4+}$



5. $H_2O$

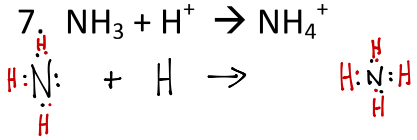


6. $SO_2$



**In the reactions below, which is the Lewis Acid and/or which is the Lewis Base?**

7. $NH_3 + H^+ \rightarrow NH_4^+$



8. $H_2O + H^+ \rightarrow H_3O^+$