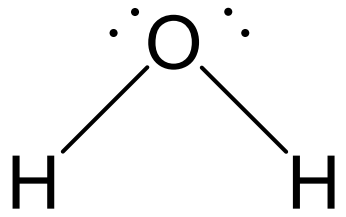
Aqueous Solution:

Aqueous(aq) systems refer to systems in which a substance is dissolved in the polar solvent water(H_2O).



Solution – Mixture of a solute in a solvent.

Solute – Component dissolved in a solvent.
Solvent – Component of mixture present in the greatest amount.

Concentration:

Relates the amount of solute to the amount of solvent in a solution.

$$\text{Molarity(M)} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

$$\text{molality(m)} = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

$$\text{mole fraction(X)} = \frac{\text{moles of solute}}{\text{total number of moles}}$$

Molarity:

$$\text{Molarity(M)} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

A 1.00 M HCl solution contains 1.00 mol of HCl in 1.00 L(1000 mL) of solution.
Likewise 0.500 mol HCl in 0.500 L.
Likewise 2.00 mol HCl in 2.00 L.

Molarity relates amount of solute in mol to volume of solution in L.

Preparation of Solutions:

1. Dilution of a stock solution.
2. Dissolving a solute in a given amount of solvent.

Dilution of Stock Solutions:

$$M_1V_1 = M_2V_2$$

M_1: concentration of stock solution
V_1: volume of stock solution used or required.
M_2: concentration of the final diluted solution.
V_2: volume of the final diluted solution.

Ex: What volume of a 12.0 M HCl stock solution should be used to prepare 500. mL of a 3.00 M HCl solution.

Ex: 1

Calculate the concentration of a KCl solution prepared by dissolving 0.500 moles KCl in 400. mL of water.

Ex: 2

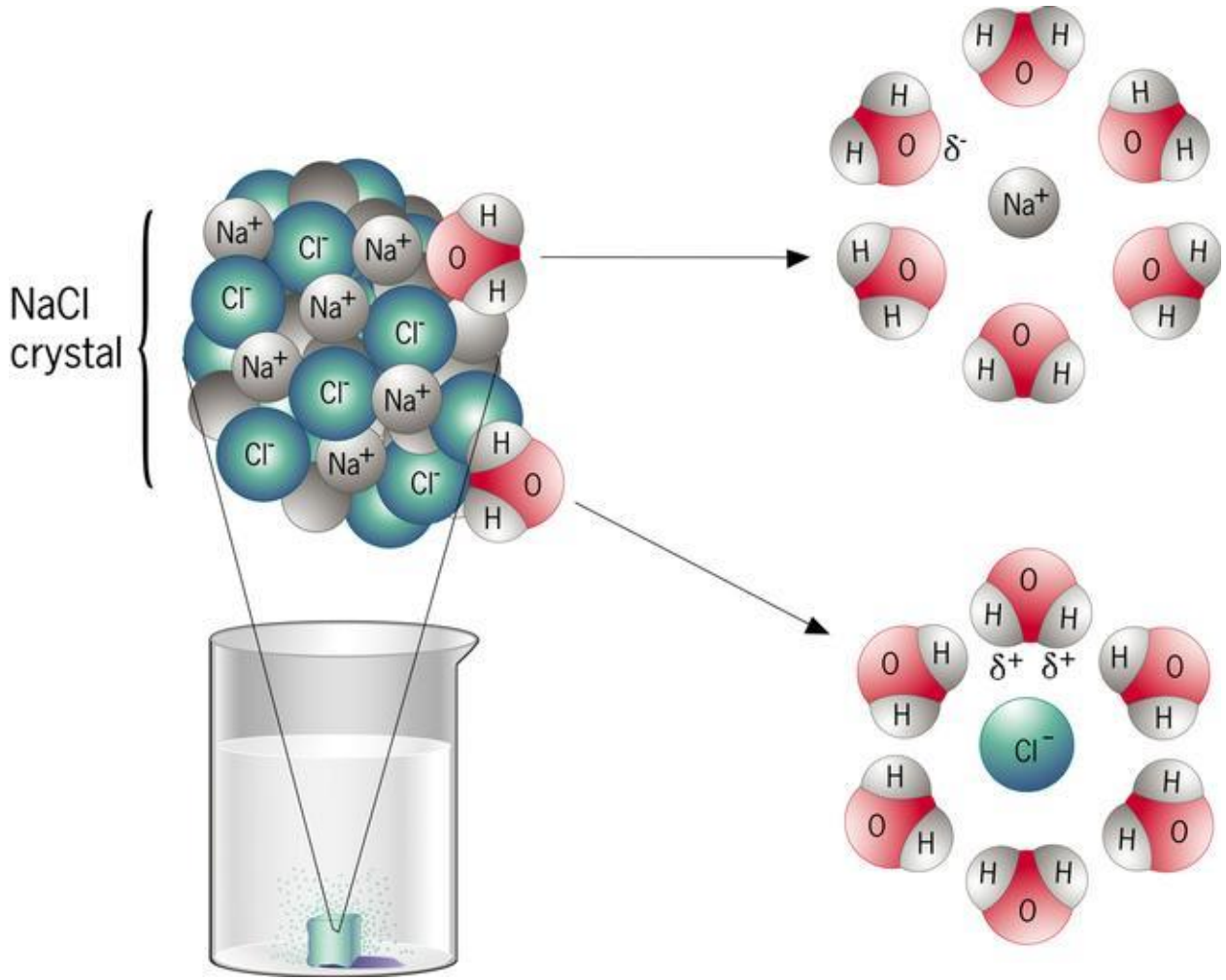
Calculate the concentration of a solution, in molarity, prepared by dissolving 51.1 g NaCl in 250. mL water.

Ex: 3

How many grams of NaOH should be dissolved in 500. mL of water to prepare a 0.750 M solution.

Chemical Reactions in Aqueous Solution:

Many chemical reactions occur in water.



Reference: http://www.bio.miami.edu/tom/courses/bil255/bil255goods/02_bonds.html

Ionic vs. Molecular Equations:

In a molecular equation reactants and products are written using their whole chemical formula.

Molecular vs. Net Ionic Equations.

Molecular equations indicate the chemical compounds interacting while Net Ionic Equations indicate the ion interactions.

Ex:

Molecular equation

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Net Ionic Equation

$$H^+(aq) + OH^-(aq) \rightarrow H_2O$$

Stoichiometry Involving Molarity:

Ex: What volume of 0.750 M NaOH is required to react with 50.0 mL of 0.159 M H_2SO_4?

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Acids and Bases:

Arrhenius Definition:

acid - A substance that produces H^+ ions(protons) in water.

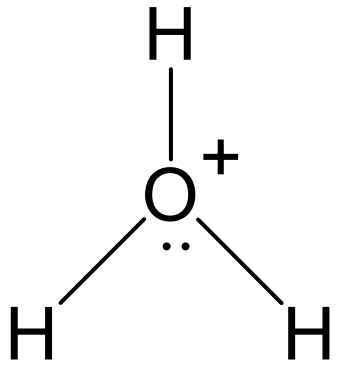
base - A substance that produces OH^- ions in water.

Bronsted Definition:

acid- Proton donor.

base- Proton acceptor.

H^+ same as



Common Acids:

Monoprotic Acids

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Hydrochloric Acid

$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

Nitric Acid

$HC_2H_3O_2(aq) \leftrightharpoons H^+(aq) + C_2H_3O_2^-(aq)$

Acetic Acid

Diprotic Acid

$$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) \leftrightharpoons H^+(aq) + SO_4^{2-}(aq)$$

Sulfuric Acid

Capable of donating two $H^+(aq)$ ions.

Triprotic Acid

$$H_3PO_4(aq) \leftrightharpoons H^+(aq) + H_2PO_4^-(aq)$$

$$H_2PO_4^-(aq) \leftrightharpoons H^+(aq) + HPO_4^{2-}(aq)$$

$$HPO_4^{2-}(aq) \leftrightharpoons H^+(aq) + PO_4^{3-}(aq)$$

Phosphoric Acid

Electrolytes and Nonelectrolytes:

Water decomposes slightly to produce ions.

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

Solutions that contain ions can conduct electricity.

Electrolyte - Solute of an aqueous solution that is a better electric conductor than pure water. Ex: HCl, NaCl.

Nonelectrolytes - Solutes that do not enhance the conductivity of the solution. Ex: sugar.

Reduction-Oxidation Reactions (REDOX):

A redox-oxidation(REDOX) reaction is a reaction in which electrons are transferred.

Oxidation- Process in which oxidation state of an element increases. Species loses electrons.

Reduction- Process in which oxidation state of an element decreases. Species gains electrons.

Assigning Oxidation Numbers:

1. Oxidation number of a free element or diatomic molecule is zero.
Ex: Na(s), Cu(s), H_2(g), F_2(g)

2. In most cases the oxidation number of hydrogen is +1, oxygen is -2, and fluorine is -1 when combined with another element.

3. The sum of the oxidation numbers of each of the elements in a molecule or ion must equal the charge.

Using Oxidation Numbers:

Ex:

$$Zn(s) + Cu^{2+} \rightarrow Zn^{2+} + Cu(s)$$

Zn(s): oxidized(lost electrons).

Cu^{2+}(aq): reduced(gained electrons).

Ex:

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

H_2(g): oxidized(lost electrons).

O_2(g): reduced(gained electrons).

REDOX cont…:

OXIDATION $Zn(s) \rightarrow Zn^{2+} + 2e^-$

REDUCTION $Cu^{2+} + 2e^- \rightarrow Cu(s)$

REDOX $Zn(s) + Cu^{2+} \rightarrow Zn^{2+} + Cu(s)$

Zn(s): oxidized/reducing agent.

Cu^{2+}(aq): reduced/oxidizing agent.

Writing Balanced Redox Reactions:

Oxidation and reduction reactions occur together. Occur in acidic or basic medium.

Ex: (acidic)

$$SO_3^{2-} + MnO_4^- \rightarrow SO_4^{2-} + Mn^{2+}$$

STEP 1: Identify the oxidized and reduced species and write the corresponding half reactions.

Writing Balanced Redox Reactions cont…:

STEP 2: Balance each of the half reactions. First atoms other than H and O. Balance O atoms by adding H_2O molecules and then balance H atoms by adding H^+ ions.

STEP 3: Balance the number of electrons.

STEP 4: Add both half reactions and simplify.

Writing Balanced Redox Reactions cont…

Balance the following redox reaction which occurs in a basic medium.

$$CrO_4^{2-} + S^{2-} \rightarrow Cr(OH)_3(s) + S(s)$$

NOTE: In basic medium add an equal number of OH^- ions to both sides to neutralize H^+ ions.

$$OH^- + H^+ \rightarrow H_2O$$