

Chapter 4 – Reactions in Solutions

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October 8, 2020

- Solute gets dissolved, Solvent does the dissolving
- Molarity is (1)

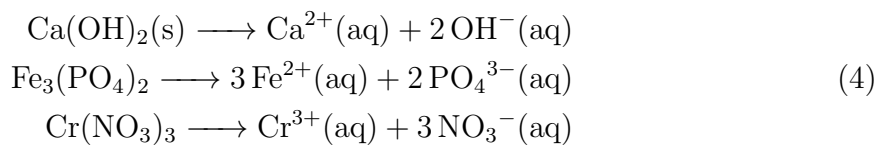
$$M = \frac{\text{mol}}{\text{L}} \quad (1)$$

- Electrolytes are ionic compounds that breakup in a solution (Ex. (2))



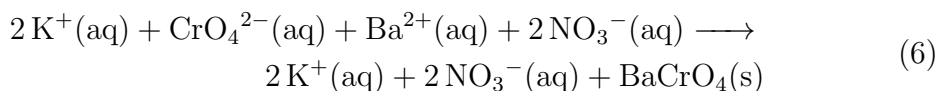
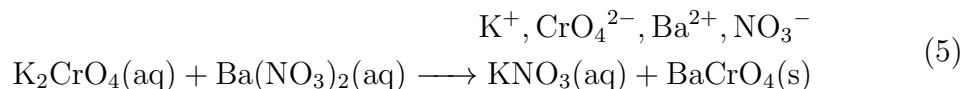
1. Strong Electrolyte vs. Weak Electrolyte – The more are broken up, the stronger the electrolyte
2. The Dilution Formula (3)

$$M_1V_1 = M_2V_2 \quad (3)$$



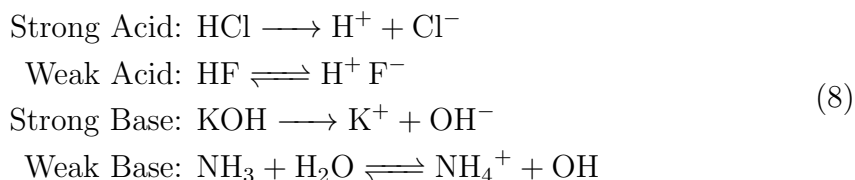
- Precipitation Reaction – Ionic compounds will either separate (soluble) or stay together (insoluble).
 1. Solubility Rules – If a compound contains any of the following three, it is soluble: Sodium (Na^+), Potassium (K^+), Nitrate (NO_3^-)

2. Ex. Potassium Chromate + Barium Nitrate (5). This is an example of a molecular equation with a double replacement. (6) is named a complete ionic equation. (7) is named a net equation, and is the only one that will be on the AP exam.



- Acid-Base Reaction

1. Acid – Produces H^+
2. Base – Produces OH^-
3. “Arrhenius” Way of thinking
4. Strong Acid – Completely Dissociates
 - (a) Examples: Hydrochloric (HCl), Sulfuric (H_2SO_4), Nitric (HNO_3), Perchloric (HClO_4), Hydrobromic (HBr), Hydroiodic (HI)
5. Weak Acid – Does not completely dissociate. Sets up an equilibrium.
 - (a) Not a strong acid.
6. Strong Base – Completely dissociates.
 - (a) Hydroxides of column I and II.
7. Weak Base – Produce OH^- with reaction with water.
 - (a) It will always be explicitly stated if something is a weak base
8. Dissociation (8)



9. Strong Acid – H^+
10. Strong Base – OH^-
11. Weaks are represented as is

12. HCl and NaOH (9)
13. HB and KOH (10)
14. H₂SO₄ and NH₃ (11)
15. HF and CH₂NH₂ (12)



16. Titration – Adding acid to base or other way around (13)

$$M_a V_a = M_b V_b \quad (13)$$

- Oxidation/Reduction (Redox Reactions) – Transfer of electrons
- Oxidation – Loss of electrons
 1. Examples of half reactions: (14)



- Reduction – Gain of electrons
 1. Examples of half reactions: (15)



- Oxidation Numbers – Used to track electrons
 1. Group 1 = +1
 2. Group 2 = +2
 3. F=-1
 - (a) H is a +1; O is -2
 - (b) Only one atom = to charge
 - (c) Ionic Compounds = charge on atoms
 - (d) Sum must be equal to overall charge

- Oxidizing Agent – Does the oxidizing, gets reduced
- Reduction Agent – Does the reducing, gets oxidized
- Balancing Redox Reactions Steps:
 1. Break into half reactions
 2. Balance atoms: Oxygen by water, Hydrogen by H^+
 3. Balance Charge (using electrons)
 4. Add OH^- to H^+ in basic solutions
 5. Add half reactions back together, and multiply by necessary coefficients