KEY
SCAN

Unit 3 – Solution Equilibriums

PART 1 GOALS:

1. How to tell whether a substance will dissolve to form an ***ionic*** solution or a ***molecular*** solution.
2. Why ***ionic*** solutions conduct an electrical current while ***molecular*** solutions do not.

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Rules for Ionic and Molecular Solutions:

1. Compounds made up of a ***metal*** (left side of staircase) and a ***non-metal*** (right side of staircase) form ***Ionic*** Solutions.

   NaCl forms an ***ionic*** solution. $NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

   AlCl3 forms an ***ionic*** solution: $AlCl_{3(s)} \rightarrow Al^{3+}_{(aq)} + 3Cl^{-}_{(aq)}$

2. Compounds containing ***polyatomic ions*** form ***Ionic*** solutions.

   KMnO4 forms an ionic solution. $KMnO_{4(s)} \rightarrow K^{+}_{(aq)} + MnO_{4}^{-}{}_{(aq)}$

3. ***Covalent*** Compounds (made up of a Non-metal and a Non-metal) generally form ***Molecular*** solutions. An example is: ***SCl2***.

   Another example is the element ***iodine*** (formula is I_2). When iodine dissolves in water it does ***NOT*** break up into ions. $I_{2(s)} \rightarrow I_{2(aq)}$

4. Most ***organic substances*** (those with C's, H's and O's in the same formula) form ***molecular*** solutions *with the exception of organic acids.*

   Table sugar $C_{12}H_{22}O_{11}$ forms a ***molecular solution*** when dissolved in water.

   $$C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(aq)}$$

5. ***Organic Acids*** (Carboxylic acids -COOH) consist of neutral molecules as a pure substance. When they dissolve in water, they dissociate partially to form some ions, thus they become ***Ionic*** Solutions.

   CH_3COOH (acetic acid) is molecular as a pure liquid (and doesn't conduct) BUT, when it is dissolved in water, some of the molecules break apart and form ions. This is called "***ionization***". Only a small fraction of the molecules will do that, so you get a limited number of free ions. This is what makes the solution a "Weak" conductor or ***Weak Electrolyte***.

   $$CH_3COOH_{(l)} \rightleftharpoons H^{+}_{(aq)} + CH_3COO^{-}_{(aq)}$$

   The subscript (l) stands for a liquid. Notice that the arrow is double but the longer one pointing left tells us that the ***reactants are favoured***. In other words, not many H^+ and CH_3COO^- ions are formed, but it is still called an ***ionic*** solution.

An important note here: Only compounds ending in the *entire group*, "***COOH***" are organic acids!

eg.) CH_3CH_2COOH when dissolved in water forms an ***ionic*** solution. (It ends in "COOH", so it is an ***organic acid***.)

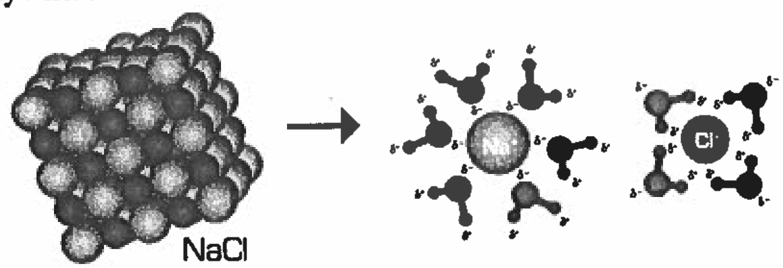
eg.) CH_3CH_2OH forms a molecular solution in water. Compounds with C's and H's and ending in "OH" are called *alcohols*. These are all ***molecular***.

One more little note: Although "OH" is a polyatomic ion in ionic compounds, when you see it with *organic compounds* (containing C's, & H's mostly), it does NOT act as an ion.

What Causes Conductivity?

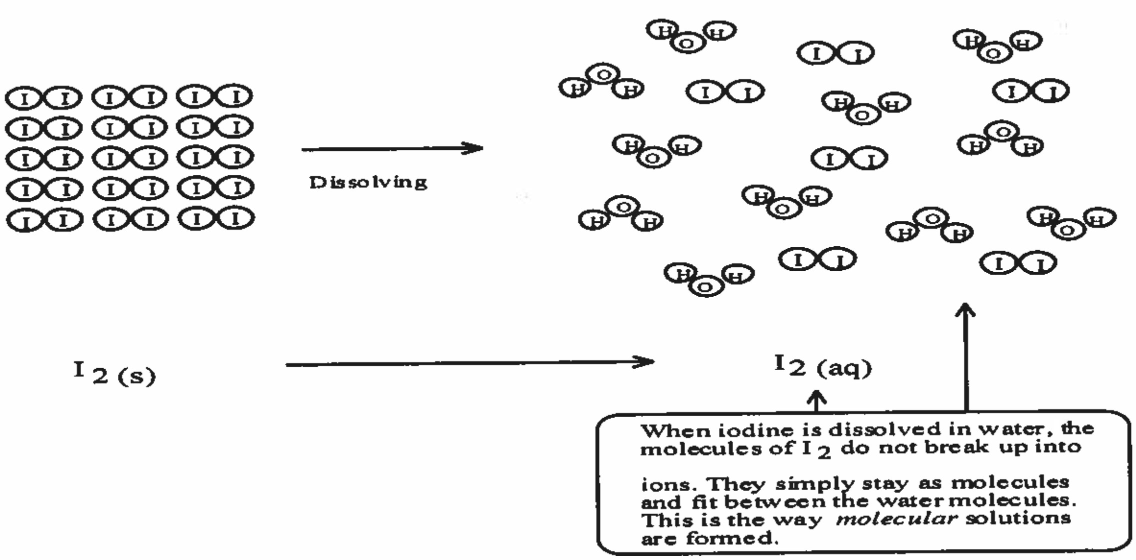
If a liquid conducts an electrical current, it means the liquid must have ***IONS*** in it. (The only exception to this is liquid mercury, which is a metal.) Even though a *solid* ionic compound is made up of ions, they are all "stuck together" or "immobile" in the solid form. Therefore, ionic compounds do not conduct in the solid form.

When they are dissolved in water, the ions are removed from the solid and move *independently* anywhere in the solution



The "+" and "-" ions are now free to move around. The "+" ions would be attracted to a negative electrode and the "-" ions would be attracted to a positive electrode. In this way, the ***ionic solution*** conducts a current.

A molecular solid (I_2) dissolving:



1. Decide whether each of the following compounds will form an Ionic (I) solution or a Molecular (M) solution in water. Assume that all substances dissolve at least partially.

a) $NiCl_2$ .... ionic

b) CH_3OH .... molecular

c) CH_3CH_2COOH .... ionic (weak)

d) $Fe(NO_3)_3$ .... ionic

e) $K_2Cr_2O_7$ .... ionic

f) $C_6H_{12}O_6$ .... molecular

g) PCl_3 .... molecular

h) CsBr .... ionic

i) HNO_3 .... ionic

j) HCOOH .... ionic (weak)

2. Write an equation showing what happens when each of the following are dissolved in water: (*"a" and "b" are done as an examples*)

a) $Na_2SO_{4(s)}$; (*ionic*) $Na_2SO_{4(s)} \rightarrow 2Na^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

b) $CH_3OH_{(l)}$; (*molecular*) $CH_3OH_{(l)} \rightarrow CH_3OH_{(aq)}$

c) $KCl_{(s)}$ .... $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$

d) $NH_4NO_{3(s)}$............. $NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$

e) $Ca_3(PO_4)_{2(s)}$ .... $Ca_3(PO_4)_2(s) \rightarrow 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$

f) $CH_3CH_2CH_2OH_{(l)}$ .... $CH_3CH_2CH_2OH(l) \rightarrow CH_3CH_2CH_2OH(aq)$

g) $CH_3CH_2CH_2COOH_{(l)}$ $CH_3CH_2CH_2COOH(l) \rightarrow CH_3CH_2CH_2COO^-(aq) + H^+(a$

3. Describe what happens when ionic solutions are formed? Do they conduct electricity?
Ionic solutions are formed when ionic crystals are brocken into free moving + and - ions by the polarity of water. These charges allow the solution to conduct electricity.

4. Describe what happens when molecular solutions are formed? Do they conduct electricity? When molecular solutions are formed, the molecules stay together and mix evenly with H_2O. No charges are present and so electricity cannot be conduct

Read over page 73-74 in Hebden and do #1a-j, #2a-d

↑ ↑ more practice same as above!

PART 2 – Solubility and Saturated solutions

Goals:

1. What is meant by the term "solubility".
2. What conditions are necessary to form a saturated solution.
3. What is happening at equilibrium in a saturated aqueous solution.
4. How to write the net ionic equation which represents a saturated solution.

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Imagine what happens to the ions of a soluble solid ionic substance as soon as you put it in water:

- At first, only the *forward reaction* is taking place: Ionic Solid → Aqueous Solution

- As dissolving continues, more and more free ions are found in solution. The chances that a free ion will collide with the crystal and stick onto it (precipitation) get greater and greater.

- As you have probably guessed by now, as more free ions are found in the solution, the rate of precipitation will continue to increase.

- Also, because the solution is getting "full" of ions (saturated with them), the rate of dissolving will decrease.

- Sooner or later, the rate of the precipitation become equal to the rate of dissolving. This is the situation we call ***Solubility Equilibrium***

Solubility Equilibrium exists when: ***The rate of dissolving = The rate of precipitation***

A saturated solution is a solution in which there exists a dissolved substance in equilibrium with the undissolved substance.

The equilibrium is shown with a double arrow: Ionic Solid ⇄ Aqueous Solution

For example: $CaCO_{3(s)} \rightleftarrows Ca^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$

How *much* solid dissolves at a certain temperature is called the ***Solubility***.

Remember that once *equilibrium* is established, ***dissolving*** will continue, but so will ***precipitation*** (at the same rate). So *the concentration of the solution will stay constant* as long as equilibrium is maintained. This concentration is called the ***equilibrium concentration***.

Solubility is the ***equilibrium concentration*** of a substance in a solution at a given temperature.

↑ changes with T°C (↑ as T°C↑ etc.)

Net Ionic Equations for Solubility Equilibrium

As was mentioned earlier, when a dissolving substance is at equilibrium:

The rate of dissolving = The rate of precipitation

We show this is at equilibrium by writing a double arrow. Chemists have chosen to show the **solid on the left** and the **dissolved ions on the right**. (Even though we know both the forward and reverse reaction are happening.)

The Net-Ionic Equation which represents the equilibrium reached when $MgSO_{4(s)}$ is dissolving is as follows:

$$MgSO_{4(s)} \rightleftarrows Mg^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Since a *saturated solution* exists at equilibrium, this can also be referred to as the equation which represents what is going on in a ***saturated solution*** of $MgSO_4$ or an equation describing the ***equilibrium present*** in saturated $MgSO_4$.

Here's another example: Write the Net-Ionic Equation for a saturated solution of $AlCl_3$.

Answer: $AlCl_{3(s)} \rightleftarrows Al^{3+}_{(aq)} + 3Cl^-_{(aq)}$

Notice that the "3" subscript after the "Cl" in $AlCl_3$ comes up *in front of* the Cl^- ion in the products. This is because there is no such ion as Cl_3. Check the ion sheet for this. There is only Cl^-.

Ions in your Net-Ionic Equation must have the same formulas and charges as the ones shown on the Table of Common Ions!

Self Test:

1. Define the ***solubility*** of a substance (use the word *equilibrium* in your definition.)

Solubility is the equilibrium concentration of a substance in a solution at a given temperature.

2. What two conditions are necessary to have a ***saturated solution*** of a substance?

1. → rate of dissolving = rate of precipitation
2. → maximum amount of solute is dissolved at given temperature.

→ some undissolved material is present.

3. When a substance is first mixed with water, the rate of dissolving is high and the rate of precipitation (or crystallization) is low. As time goes on, the rate of precipitations gets higher - and the rate of dissolving gets lower. When the rate of dissolving equals the rate of precipitations equilibrium has been reached.

4. Give the ***Net-Ionic Equation*** which represents a ***saturated solution*** of each of the following ionic substances in water:(*Hint: These are just like dissociation equations but they have a double arrow, indicating equilibrium.*)

a) $Ag_2SO_{4(s)}$ $Ag_2SO_4(s) \rightleftharpoons 2Ag^+(aq) + SO_4^{2-}(aq)$

b) $FeS_{(s)}$ $FeS(s) \rightleftharpoons Fe^{2+}(aq) + S^{2-}(aq)$

c) $Mg(OH)_{2(s)}$ $Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$

d) $Ca_3(PO_4)_{2(s)}$ $Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$

e) $BaSO_{3(s)}$ $BaSO_3(s) \rightleftharpoons Ba^{2+}(aq) + SO_3^{2-}(aq)$

f) $(NH_4)_2CrO_{4(s)}$ $(NH_4)_2CrO_4(s) \rightleftharpoons 2NH_4^+(aq) + CrO_4^{2-}(aq)$

g) $Fe(OH)_{3(s)}$ $Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$

h) $Al_2(SO_4)_{3(s)}$ $Al_2(SO_4)_3(s) \rightleftharpoons 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

Read pages 75-76 in Hebden and do pg. 76 #3-7 ← same as above #6+7 thinking Q.

Read pages 77 in Hebden (review unit conversion) and do pg. 77 #8-11

Do This!