Solubility:

Solubility is the measure of how much of a solute will dissolve in a solvent. In general chemistry, we usually talk about water as the solvent, so we are talking about what compounds will dissolve in water. We also talk a lot about ionic compounds, so the goal is to determine which ionic compounds are soluble in water and which aren't. There are many factors to consider in the solubility of ionic compounds, none of which you need to worry about. Instead, you will be given solubility rules which you must apply. If it turns out that something is soluble (it does dissolve in water), then you need to put a subscript (aq) next to it (i.e. $NaCl_{(aq)}$). The (aq) stands for aqueous and implies that you have a solution. If the compound is insoluble (does not dissolve in water), then you need to put a subscript (s) next to it. You will be given a list of the rules for the exam and the final, but I recommend memorizing the first rule, because it has no exceptions and it takes precedence over all of the other rules. If rule #1 does not apply, look up the anion in the remaining rules. **Solubility rules apply ONLY to ionic compounds, NOT to acids or molecular compounds!!!!!!!**

Example: calcium chromate – Calcium is not mentioned in the first rule, so you need to find chromate in the remaining rules. Rule #7 says that chromates are insoluble so it is $CaCrO_{4(s)}$

Example: rubidium oxide – rubidium is a group 1 metal, so you need look no further; all group 1 metal containing compounds are soluble, so it is $Rb_2O_{(aq)}$

Example: lead (II) bromide – according to rule #3, most bromide salts are soluble, but there are exceptions (silver, mercury (I) and lead (II). This means that lead (II) bromide is going to be insoluble: $PbBr_{2(s)}$

You try:
1) $(NH_4)_2S$
2) BaO
3) $NaCO_3$
4) Hg_2SO_4
5) $Pb(ClO_4)_2$
6) SrF_2
7) MgC_2O_4
8) Li_3PO_4
9) AgI
10) $Mn(OH)_4$

Electrolytes:

Electrolytes are broken down into 3 different types; non-electrolytes, weak electrolytes, and strong electrolytes. You will have to be able to differentiate between them. It is NOT that hard, it just takes a second to memorize. Seriously, there are no tricks here, simple memorization and you have it.

Non-electrolytes: Any molecular compound (compounds without metals)
Weak electrolytes: Any insoluble salt or any weak acid. (It is a weak acid if it is NOT one of the 7 strong acids)
Strong electrolytes: Any strong acid or any soluble salt

That's it. You can already tell between insoluble salts and soluble salts. You learned was a molecular compound was in the second week of class and as long as you memorize the 6 strong acids, you can tell the difference between strong and weak acids (the 7 strong acids are $HCl_{(aq)}$, $HBr_{(aq)}$, $HI_{(aq)}$, $HNO_{3(aq)}$, $HClO_{3(aq)}$, $HClO_{4(aq)}$, and $H_2SO_{4(aq)}$).

You try:
1) $(NH_4)_2S$
2) $HC_7H_6O_{2(aq)}$
3) $N_2O_{5(l)}$
4) Hg_2SO_4
5) $HClO_{4(aq)}$
6) SrF_2
7) $Rb_2C_2O_4$
8) $H_3PO_{4(aq)}$
9) AgI
10) $Sr(OH)_2$

Net ionic equations:

Remember, in order to write a net ionic equation, you must write out a BALANCED molecular equation.

Example: Sodium carbonate is mixed with bismuth (III) nitrate

Translate: $Na_2CO_3 + Bi(NO_3)_3$

Write symbols for products: $Na_2CO_3 + Bi(NO_3)_3 \rightarrow Na^+ \ NO_3^- + Bi^{+3} \ CO_3^{-2}$

Balance CHARGES of products: $Na_2CO_3 + Bi(NO_3)_3 \rightarrow NaNO_3 + Bi_2(CO_3)_3$

NO MORE TOUCHING THE SUBSCRIPTS!!!!!!!!

Balance the equation: (**ONLY CHANGE COEFICIENTS TO BALANCE EQUATION!!!**)

Initial: $Na_2CO_3 + Bi(NO_3)_3 \rightarrow NaNO_3 + Bi_2(CO_3)_3$
Multiply Na_2CO_3 by 3 so that the carbonates are equal on both sides.

Trial 1: $\mathbf{3}Na_2CO_3 + Bi(NO_3)_3 \rightarrow NaNO_3 + Bi_2(CO_3)_3$
Multiply $NaNO_3$ by 6 so that the nitrates are equal on both sides.

Trial 2: $\mathbf{3}Na_2CO_3 + Bi(NO_3)_3 \rightarrow \mathbf{6}NaNO_3 + Bi_2(CO_3)_3$
Multiply $Bi(NO_3)_3$ by 2 so that the bismuths are equal on both sides.

Trial 3: $\mathbf{3}Na_2CO_3 + \mathbf{2}Bi(NO_3)_3 \rightarrow \mathbf{6}NaNO_3 + Bi_2(CO_3)_3$

		Reactants					Products		
Initial		Trial 1	Trial 2	Trial 3	Initial		Trial 1	Trial 2	Trial 3
Na	2	6	6	**6**	Na	1	1	6	**6**
CO_3	1	3	3	**3**	CO_3	3	3	3	**3**
Bi	1	1	1	**2**	Bi	2	2	2	**2**
NO_3	3	3	3	**6**	NO_3	1	1	6	**6**

BALANCED!!!!

Last step in writing the balanced molecular equation is to add in to the state subscripts (aq, l, g, s). **Look at the solubility rules to determine if a salt is soluble**. All acids are written as (aq)

Add in state subscripts: $\mathbf{3}Na_2CO_{3(aq)} + \mathbf{2}Bi(NO_3)_{3(aq)} \rightarrow \mathbf{6}NaNO_{3(aq)} + Bi_2(CO_3)_{3(s)}$

So the balances molecular equation is: $\mathbf{3}Na_2CO_{3(aq)} + \mathbf{2}Bi(NO_3)_{3(aq)} \rightarrow \mathbf{6}NaNO_{3(aq)} + Bi_2(CO_3)_{3(s)}$

Step one is done, on to step two; writing the ionic equation:

Once the BALANCED molecular equation is written, you have to write the ionic equation. This is done by breaking up ALL strong electrolytes into ions. You break up ONLY the strong electrolytes; weak electrolytes and non-electrolytes are left alone. Make sure not to break up polyatomic ions; i.e. leave $NO_3^-{}_{(aq)}$ alone instead of splitting it into nitrogen and oxygen. Also, make sure you have the correct number of each ion present.

So, the ionic equation looks like this:

$$6Na^+_{(aq)} + 3CO_3^{2-}{}_{(aq)} + 2Bi^{3+}_{(aq)} + 6NO_3^-{}_{(aq)} \rightarrow 6Na^+_{(aq)} + 6NO_3^-{}_{(aq)} + Bi_2(CO_3)_{3(s)}$$

That is it for the IONIC EQUTION. The final step is to write the NET IONIC EQUATION. To do this, cross out anything that is IDENTICAL on both sides of the ionic equation, like so:

$$\cancel{6Na}^+_{(aq)} + 3CO_3^{2-}{}_{(aq)} + 2Bi^{3+}_{(aq)} + \cancel{6NO_3}^-{}_{(aq)} \rightarrow \cancel{6Na}^+_{(aq)} + \cancel{6NO_3}^-{}_{(aq)} + Bi_2(CO_3)_{3(s)}$$

The ions that are crossed out are called spectator ions. Rewrite anything that is not crossed out. This is the NET IONIC EQUATION!!

$3CO_3^{2-}{}_{(aq)} + 2Bi^{3+}{}_{(aq)} \rightarrow Bi_2(CO_3)_{3(s)}$

Examples:

Write the net ionic equation of the reaction between silver sulfate and calcium chloride.

Balance the equation: $Ag_2SO_4 + CaCl_2 \rightarrow 2AgCl + CaSO_4$

Add in state subscripts: $Ag_2SO_{4(aq)} + CaCl_{2(aq)} \rightarrow 2AgCl_{(s)} + CaSO_{4(aq)}$

Molecular equation: $\mathbf{Ag_2SO_{4(aq)} + CaCl_{2(aq)} \rightarrow 2AgCl_{(s)} + CaSO_{4(aq)}}$

Ionic equation: $\mathbf{2Ag^+{}_{(aq)} + SO_4^{-2}{}_{(aq)} + Ca^{+2}{}_{(aq)} + 2Cl^-{}_{(aq)} \rightarrow 2AgCl_{(s)} + Ca^{+2}{}_{(aq)} + SO_4^{-2}{}_{(aq)}}$

Cross out spectators: $2Ag^+{}_{(aq)} + \cancel{SO_4}^{-2}{}_{(aq)} + \cancel{Ca}^{+2}{}_{(aq)} + 2Cl^-{}_{(aq)} \rightarrow 2AgCl_{(s)} + \cancel{Ca}^{+2}{}_{(aq)} + \cancel{SO_4}^{-2}{}_{(aq)}$

Net ionic equation: $\mathbf{Ag^+{}_{(aq)} + Cl^-{}_{(aq)} \rightarrow AgCl_{(s)}}$
(note the remaining coefficients reduced to the lowest common denominator)

Okay, so what about the states on molecular compounds? For most of these, I will have to tell you what the state is. If it is a molecular compound, whether it is (aq), (g), (l), or (s), it does not change or break up EVER (molecular compounds are non-electrolytes)!!. **You should know that CO_2 is a gas and water is a liquid unless it is otherwise stated.**

Remember, strong acids are strong electrolytes and will therefore ionize (i.e. $HCl_{(aq)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)}$). If the acid that is present is NOT one of these 6 acids, then it is a WEAK acid (weak electrolyte) and **DOES NOT EVER BREAK UP IN THE N.I.E.**!!!!! (i.e. $HCO_2H_{(aq)}$ stays $HCO_2H_{(aq)}$)

Write the N.I.E. for the reaction between calcium hydroxide and hypochlorous acid.

Molecular equation: $\mathbf{Ca(OH)_{2\,(aq)} + 2HClO_{(aq)} \rightarrow Ca(ClO)_{2\,(aq)} + 2H_2O_{(l)}}$

Ionic equation: $\mathbf{Ca^{+2}{}_{(aq)} + 2OH^-{}_{(aq)} + 2HClO_{(aq)} \rightarrow Ca^{+2}{}_{(aq)} + ClO^-{}_{(aq)} + 2H_2O_{(l)}}$
(HClO is a weak electrolyte, so it does **NOT** break up)

Cross out spectators: $\cancel{Ca}^{+2}{}_{(aq)} + 2OH^-{}_{(aq)} + 2HClO_{(aq)} \rightarrow \cancel{Ca}^{+2}{}_{(aq)} + 2ClO^-{}_{(aq)} + 2H_2O_{(l)}$

Net ionic equation: $\mathbf{OH^-{}_{(aq)} + HClO_{(aq)} \rightarrow ClO^-{}_{(aq)} + H_2O_{(l)}}$
(since the coefficient on each of the remaining pieces was a 2, they were all reduced to 1)

Write the N.I.E. for the reaction between barium hydroxide and nitric acid.

Molecular equation: $Ba(OH)_{2\,(aq)} + 2HNO_{3\,(aq)} \rightarrow Ba(NO_3)_{2\,(aq)} + 2H_2O_{(l)}$

Ionic equation: $Ba^{+2}{}_{(aq)} + 2OH^-{}_{(aq)} + 2H^+ + 2NO_3^-{}_{(aq)} \rightarrow Ba^{+2}{}_{(aq)} + 2NO_3^-{}_{(aq)} + 2H_2O_{(l)}$
(HNO_3 is a strong acid)

Net ionic equation: $OH^-{}_{(aq)} + H^+{}_{(aq)} \rightarrow H_2O_{(l)}$

Write the N.I.E. for the reaction between ammonium sulfate and potassium iodide

Molecular equation: $(NH_4)_2SO_{4(aq)} + 2KI_{(aq)} \rightarrow K_2SO_{4(aq)} + 2NH_4I_{(aq)}$

Ionic equation: $2NH_4^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2K^+{}_{(aq)} + 2I^-{}_{(aq)} \rightarrow 2K^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2NH_4^+{}_{(aq)} + 2I^-{}_{(aq)}$

Net ionic equation: NO REACTION (when everything cancels on both sides, there is no reaction)

Practice:

Write M.E., I.E., and N.I.E. for each of the following reactions.

potassium phosphate and mercury (I) acetate

sodium sulfide with iron (III) nitrate

oxalic acid and cesium hydroxide

mercury (I) perchlorate and lead (II) nitrate

chromium (II) bromide with lithium oxalate

sulfuric acid with rubidium hydroxide

calcium hydroxide with periodic acid

cesium chromate with rubidium oxide

sodium fluoride with nickel (III) sulfate

antimony (V) carbonate reacts with sulfuric acid

lithium acetate with copper (II) nitrate

You should try to make up 2 of your own reactions. There should be a reaction (i.e. something must change from products to reactants) Do not look reactions up in a book, use only the solubility rules and the periodic table to come up with your own reactions.