Name ______________________ Section ________ Date __________

Chapter 8: Basic Review Worksheet

1. What is meant by the *driving force* for a reaction?

2. What is a *precipitation reaction*? Provide an example.

3. Define the term *strong electrolyte*. Provide three formulas of strong electrolytes and name them.

4. Summarize the simple solubility rules for ionic compounds.

5. In general terms, what are the *spectator ions*, in a precipitation reaction?

6. Describe some physical and chemical properties of *acids* and *bases*. What is meant by a *strong* acid or base?

7. What is a *salt*? Provide three formulas of salts and name them.

8. What is *oxidation*? What is *reduction*?

9. What is a *combustion* reaction? Write an equation that illustrates a combustion reaction.

10. Give an example of a *synthesis* reaction and of a *decomposition* reaction.

11. Classify the reaction represented by each of the following chemical equations in as *many* ways as possible based on what you have learned. Balance each equation.
    a. $NaOH(s) + CuSO_4(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
    b. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(l)$
    c. $FeO(s) + HNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + H_2O(l)$
    d. $C_{12}H_{22}O_{11}(s) \rightarrow C(s) + H_2O(g)$
    e. $B(s) + O_2(g) \rightarrow B_2O_3(s)$

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Chapter 8: Review Worksheet

1. Give some examples of driving forces that make reactants tend to form products. Write a balanced chemical equation illustrating each type of driving force you have named.

2. What would you see if a precipitation reaction were to take place in a beaker? Write a balanced chemical equation illustrating a precipitation reaction.

3. What types of substances tend to be strong electrolytes? What does a solution of a strong electrolyte contain? Give a way to determine if a substance is a strong electrolyte.

4. How do we use the solubility rules in determining the identity of the solid formed in a precipitation reaction? Give three examples including balanced complete and net ionic equations.

5. Why are the spectator ions not included in writing the net ionic equation for a precipitation reaction?

6. Write chemical equations showing the formation of three different salts. What other product is formed when an aqueous acid reacts with an aqueous base? Write the net ionic equation for the formation of this substance.

7. What is essential in an oxidation-reduction reaction? Write a balanced chemical equation illustrating an oxidation-reduction reaction between a metal and nonmetal. Indicate which species is oxidized and which is reduced.

8. Are combustion reactions a unique type of reaction, or are they a special case of a more general type of reaction?

9. Are synthesis and decomposition reactions always also oxidation-reduction reactions? Explain.

10. Classify the reaction represented by each of the following chemical equations in as *many* ways as possible based on what you have learned. Balance each equation.
    a. $Mg(s) + CO_2(g) + O_2(g) \rightarrow MgCO_3(s)$
    b. $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
    c. $Co(NH_3)_6Cl_2(s) \rightarrow CoCl_2(s) + NH_3(g)$
    d. $HCl(aq) + Pb(C_2H_3O_2)_2(aq) \rightarrow HC_2H_3O_2(aq) + PbCl_2(s)$
    e. $Al(s) + HNO_3(aq) \rightarrow Al(NO_3)_3(aq) + H_2(g)$

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Chapter 8: Challenge Review Worksheet

1. Spectator ions are not included in writing the net ionic equation for a precipitation reaction. Does this mean that the spectator ions do not have to be present in the solution?

2. Are strong acids and bases also strong electrolytes? Give several examples of strong acids and strong bases.

3. Three common strong acids are HCl, HNO_3, and H_2SO_4, whereas NaOH and KOH are two common strong bases. Write the neutralization reaction equations for each of these strong acids with each of these strong bases in aqueous solution.

4. The reagent shelf in a general chemistry lab contains aqueous solutions of the following substances; silver nitrate, sodium chloride, acetic acid, nitric acid, sulfuric acid, potassium chromate, barium nitrate, phosphoric acid, hydrochloric acid, lead nitrate, sodium hydroxide, and sodium carbonate. Suggest how you might prepare the following pure substances using these reagents and any normal laboratory equipment. If it is *not* possible to prepare a substance using these reagents, indicate why.
   a. $BaCrO_4(s)$
   b. $NaC_2H_3O_2(s)$
   c. AgCl(s)
   d. $PbSO_4(s)$
   e. $Na_2SO_4(s)$
   f. $BaCO_3(s)$

5. Can an oxidation reaction take place without a reduction also taking place? Why?

6. List and define all the ways of classifying chemical reaction that have been discussed in the text. Give a balanced chemical equation as an example of each type of reaction, and show clearly how your example fits the definition you have given.

c. $3Na (s) + P (s) \rightarrow Na_3P (s)$
$3K (s) + P (s) \rightarrow K_3P (s)$
d. $6Na (s) + N_2 (g) \rightarrow 2Na_3N (s)$
$6K (s) + N_2 (g) \rightarrow 2K_3N (s)$
e. $2Na (s) + H_2 (g) \rightarrow 2NaH (s)$
$2K (s) + H_2 (g) \rightarrow 2KH (s)$

4. $CO_2 (g) + H_2O(l) + NH_3 (aq) + NaCl (aq) \rightarrow NaHCO_3 (s) + NH_4Cl (aq)$

Chapter 8: Basic Review Worksheet

1. The concept of a "driving force" for chemical reactions, at this point, is a rather nebulous idea. Clearly there must be some reason why certain substances react when combined, and why other substances can be combined without anything happening. Because there are driving forces, we can use some generalizations about what sorts of events tend to make a reaction take place.

2. A precipitation reaction is one in which a *solid* forms when the reactants are combined: the solid is called a precipitate. An example is $Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2 (s) + 2NaNO_3(aq)$

3. A strong electrolyte is one that completely dissociates into ions when dissolved in water. That is, each unit of the substance that dissolves in water produces separated, free ions. For example, NaCl (sodium chloride), KNO_3 (potassium nitrate), and NaOH (sodium hydroxide) are strong electrolytes.

4. In summary, nearly all compounds containing the nitrate, sodium, potassium, and ammonium ions are soluble in water. Most salts containing the chloride and sulfate ions are soluble in water, with specific exceptions (see Table 8.1 for these exceptions). Most compounds containing the hydroxide, sulfide, carbonate, and phosphate ions are not soluble in water, unless the compound also contains one of the cations mentioned above (Na^+, K^+, NH_4^+).

5. The spectator ions in a precipitation reaction are the ions in the solution that do *not* precipitate.

6. Acids (such as the citric acid found in citrus fruits and the acetic acid found in vinegar) were first noted primarily because of their sour taste. The first bases noted were characterized by their bitter taste and slippery feel on the skin. Acids and bases chemically react with (neutralize) each other forming water. The net ionic equation is

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

The strong acids and bases are those that fully ionize when they dissolve in water: since these substances fully ionize, they are strong electrolytes.

7. A salt can be thought of as any ionic compound that contains ions other than H^+ and OH^- (compounds containing these ions are called acids and bases, respectively). In particular, a salt is formed in the neutralization reaction between an acid and a base. Examples will vary, but could include NaCl (sodium chloride), KNO_3 (potassium nitrate), and Na_2SO_4 (sodium sulfate).

8. Oxidation represents a loss of electrons by an atom, molecule, or ion, whereas reduction is the gain of electrons by such a species.

9. Combustion reactions represent processes involving oxygen gas that release energy rapidly enough that a flame is produced. An example is the burning of methane (natural gas):

$$CH_4(g) + 2O_2 \rightarrow CO_2(g) + 2H_2O(g) + \text{energy}$$

10. In general, a synthesis reaction represents the reaction of elements or simple compounds to produce more complex substances. There are many examples of synthesis reactions, for example

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Decomposition reactions represent the breakdown of a more complex substance into simpler substances. There are many examples of decomposition reactions, for example

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

11. a. $2NaOH(s) + CuSO_4(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
   precipitation, double-displacement
   b. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(l)$
   acid-base, double-displacement
   c. $FeO(s) + 2HNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + H_2O(l)$
   acid-base, double-displacement
   d. $C_{12}H_{22}O_{11}(s) \rightarrow 12C(s) + 11H_2O(g)$
   decomposition; oxidation-reduction
   e. $4B(s) + 3O_2(g) \rightarrow 2B_2O_3(s)$
   synthesis; oxidation-reduction

Chapter 8: Review Worksheet

1. A reaction is likely if any of the following things occur as a result of the reaction: formation of a solid, formation of water (or another non-ionized molecule), formation of a gas, or the transfer of electrons from one species to another. Examples of reactions illustrating each of these:

   formation of a solid:
   $BaCl_2(aq) + K_2CrO_4(aq) \rightarrow 2KCl(aq) + BaCrO_4(s)$
   $Pb(NO_3)_2(aq) + 2NaCl(aq) \rightarrow PbCl_2(s) + 2NaNO_3(aq)$

   formation of water:
   $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
   $HC_2H_3O_2(aq) + KOH(aq) \rightarrow KC_2H_3O_2(aq) + H_2O(l)$

   formation of a gas:
   $2NI_3(s) \rightarrow N_2(g) + 3I_2(s)$
   $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

   transfer of electrons:
   $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
   $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

2. If you were to perform such a reaction, the mixture would turn cloudy as the reactants are combined, and a solid would eventually settle from the mixture on standing. One example would be to combine barium nitrate and sodium carbonate solutions: a precipitate of barium carbonate would form.

$$Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaNO_3(aq)$$

3. Ionic compounds, since they already consist of ions in the solid state, are strong electrolytes if they are soluble in water. Certain acids and bases also behave as strong electrolytes. A solution of a strong electrolyte consists of free, separated ions moving through the solvent independently of one another (there are no molecules or clusters of combined positive and negative ions). An apparatus for experimentally determining whether or not a substance is an electrolyte is shown in Figure 8.2 in the text.

4. Examples will vary. The solubility rules are phrased as if you had a sample of a given solute and wanted to see if you could dissolve it in water. These rules can also be applied, however, to predict the identity of the solid produced in a precipitation reaction: a given combination of ions will not be soluble in water whether you take a pure compound out of a reagent bottle or if you generate the insoluble combination of ions during a chemical reaction. For example, the solubility rules say that $BaSO_4$ is not soluble in water. This means not only that a pure sample of $BaSO_4$ taken from a reagent bottle will not dissolve in water, but also that if Ba^{2+} ion and SO_4^{2-} ion end up together in the same solution, they will precipitate as $BaSO_4$. For example, solid barium sulfate forms when we combine barium chloride and sulfuric acid solutions:

$$BaCl_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$$

Also, if a barium nitrate solution were combined with a sodium sulfate solution, a precipitate of $BaSO_4$ would form. Barium sulfate is insoluble in water regardless of its source.

5. Since we take the actual chemical reaction in a precipitation process to be the formation of the solid, and since the spectator ions are not found in and do not participate in the formation of the solid, we leave them out of the net ionic equation for the reaction.

6. The textbook describes acid-base neutralization reactions as reactions that result in the formation of water: the water results form the combination of the $H^+(aq)$ ion from the acid with the OH^- (aq) ion from the base:

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

However, the H^+ ion must have been paired with some negative ion in the original acid solution, and the OH^- ion must have been paired with some positive ion in the original base solution. These counter-ions to the acid-base ions are what constitute the salt that is formed during the neutralization. Below are three acid-base neutralization reactions, with the salts that are formed indicated:

$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{NaOH(aq)} \rightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{a salt}}{NaCl(aq)}$$

$$\underset{\text{acid}}{HNO_3(aq)} + \underset{\text{base}}{KOH(aq)} \rightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{a salt}}{KNO_3(aq)}$$

$HC_2H_3O_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$
acid base water a salt

7. Oxidation-reduction reactions are electron-transfer reactions. An example of a simple oxidation reduction reaction between a metal and a nonmetal could be the following

$$Mg(s) + F_2(g) \rightarrow MgF_2(s)$$

In this process, Mg atoms lose two electrons each to become Mg^{2+} ions in MgF_2: Mg is oxidized. Each F atom of F_2 gains one electron to become an F^- ion, for a total of two electrons gained for each F_2 molecule: F_2 is reduced.

$$Mg \rightarrow Mg^{2+} + 2e^- \qquad 2(F + e^- \rightarrow F^-)$$

8. Combustion reactions are a special sub-class of oxidation-reduction reactions (the fact that elemental oxygen gas is a reactant but combined oxygen is a product shows this). The most common combustion reactions are those we make use of through the burning of petroleum products as sources of heat, light, or other forms of energy.

9. Synthesis and decomposition reactions are very often also oxidation-reduction reactions, especially if an elemental substance reacts or is generated. It is not necessary, however, for synthesis and decomposition reactions to always involve oxidation-reduction. For example, the reaction between NaOH and CO_2 as shown below does not represent oxidation-reduction.

$$NaOH(aq) + CO_2(g) \rightarrow NaHCO_3(s)$$

10. a. $2Mg(s) + 2CO_2(g) + O_2(g) \rightarrow 2MgCO_3(s)$
synthesis; oxidation-reduction
b. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
combustion, oxidation-reduction
c. $Co(NH_3)_6Cl_2(s) \rightarrow CoCl_2(s) + 6NH_3(g)$
decomposition
d. $2HCl(aq) + Pb(C_2H_3O_2)_2(aq) \rightarrow 2HC_2H_3O_2(aq) + PbCl_2(s)$
precipitation, double-displacement
e. $2Al(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3(aq) + 3H_2(g)$
oxidation-reduction; single-displacement

Chapter 8: Challenge Review Worksheet

1. Just because we leave the spectator ions out when writing a net ionic equation for a reaction does not mean that the spectator ions do not have to be present: the spectator ions are needed to provide a balance of charge in the reactant compounds for the ions which combine to form the precipitate. For a reaction in which silver chloride is formed, it would not be possible to have a reagent bottle containing just silver ions (there would have to be some negative ions present) or just chloride ions (there would have to be some positive ions present).

2. Yes, strong acids and bases are also strong electrolytes. The common strong acids are HCl (hydrochloric), HNO_3 (nitric), H_2SO_4 (sulfuric), and $HClO_4$ (perchloric). The most common strong bases are the alkali metal hydroxides, particularly NaOH (sodium hydroxide) and KOH (potassium hydroxide).

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

   $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

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

   $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$

   $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

   $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

4. a. $Ba(NO_3)_2(aq) + K_2CrO_4(aq) \rightarrow BaCrO_4(s) + 2KNO_3(aq)$
   b. $NaOH(aq) + HC_2H_3O_2(aq) \rightarrow H_2O(l) + NaC_2H_3O_2(aq)$
   (then evaporate the water from the solution)
   c. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   d. $Pb(NO_3)_2(aq) + H_2SO_4(aq) \rightarrow PbSO_4(s) + 2HNO_3(aq)$
   e. $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   (then evaporate the water from the solution)
   f. $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaNO_3(aq)$

5. No. Since an oxidation-reduction process represents the transfer of electrons between species, you can't have one without the other also taking place: the electrons lost by one species must be gained my some other species.

6. The different ways of classifying chemical reactions that have been discussed in the text are listed below, along with an example of each type of reaction:

   formation of a solid (precipitation):
   $FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$
   formation of water (acid-base):
   $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
   transfer of electrons (oxidation-reduction):
   $2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$
   combustion:
   $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g) + \text{energy}$
   synthesis (combination):
   $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$
   decomposition:
   $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$
   single displacement:
   $Mg(s) + 2AgNO_3(aq) \rightarrow Mg(NO_3)_2(aq) + 2Ag(s)$
   double displacement:
   $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow 2NaCl(aq) + BaSO_4(s)$

