Chapter 18
Practice Test

Name: ______________________

Date:________________ Per:________

1. Reactions that may proceed in a forward or reverse reaction are called ____reversible_____ reactions. They eventually reach a point where the forward and reverse reactions occur at the same ________rate___________. This point is called _____equilibrium____. Once the reactions reach this point, the __concentrations______ of reactants and products become _____constant___, but rarely equal.

2. An _equilibrium expression___ is a ratio of products to reactants at ______equilibrium____. Its value is represented by the symbol ____K_{eq}______ . The ratio is expressed by placing the ___concentrations of products_____ of the reaction in the numerator and the ____concentrations of reactants_____ in the denominator. The ______coefficients____ from the balanced equation are used as ______exponents_____. Because their concentrations never change at a given temperature, _____solids_____ and _____liquids____ are not included in the __equilibrium expression____.

3. Write the equilibrium expression for each of the following unbalanced equations.
   a. $H_2(g) + Cl_2(g) \leftrightarrows 2HCl(g)$

   $$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]}$$

   b. $N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$

   $$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

   c. $S_8(s) + 8O_2(g) \leftrightarrows 8SO_2(g)$

   $$K_{eq} = \frac{[SO_2]^8}{[O_2]^8}$$

   [S_8] doesn't appear because it's a solid

   d. $NH_4NO_2(s) \leftrightarrows N_2(g) + 2H_2O(g)$

   $$K_{eq} = [N_2][H_2O]^2$$

   [NH_4NO_2] doesn't appear because it's a solid

4. From the data provided below, calculate the value of the equilibrium constant for the reaction.
   a. $H_2(g) + Cl_2(g) \leftrightarrows HCl(g)$; $[H_2] = [Cl_2] = 1.0 \times 10^{-2}$; $[HCl] = 1.0 \times 10^{-4}$

   $$K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]} \qquad K_{eq} = \frac{[1.0 \times 10^{-4}]^2}{[1.0 \times 10^{-2}][1.0 \times 10^{-2}]} = 0.0001$$

   b. $N_2(g) + 3 H_2(g) \leftrightarrows 2 NH_3(g)$; $[N_2] = 4.4 \times 10^{-2}$; $[H_2] = 1.2 \times 10^{-1}$; $[NH_3] = 3.4 \times 10^{-3}$

   $$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3} \qquad K_{eq} = \frac{[3.4 \times 10^{-3}]^2}{[4.4 \times 10^{-2}][1.2 \times 10^{-1}]^3} = 0.152$$

5. To check if reaction has reached equilibrium a _reaction quotient____ , represented by the symbol ___Q___, may be calculated. If _______$Q < K_{eq}$___, then the reaction will continue to proceed forward, if _________$Q > K_{eq}$___, the reaction will proceed in reverse, and if ___________$Q = K_{eq}$_______, the reaction is at equilibrium.

6. For the reaction $N_2O_4(g) \leftrightarrows NO_2(g)$, $K_{eq} = 0.2$. At a particular time, the following concentrations are measured; $[N_2O_4] = 2.0$ *M*, $[NO_2] = 0.2$ *M*. Is this reaction at equilibrium? If not, in which direction will the reaction proceed?

   $N_2O_4(g) \leftrightarrows 2NO_2(g)$

   $[N_2O_4] = 2.0$ *M* , $[NO_2] = 0.2$ *M*

   $$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]} \qquad Q = \frac{[0.2]^2}{[2.0]} = 0.02$$

   $K_{eq} = 0.2$ $Q < K_{eq}$, so the reaction will proceed to the right (forward).

7. For the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$, the equilibrium concentrations of the sulfur oxides are $[SO_2] = 1.60$ and $[SO_3] = 2.70$. What is the concentration of oxygen when the $K_{eq} = 1.87$ for the reaction?

   $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

   $[SO_2] = 1.60$ *M* , $[O_2] = ?$ *M* , $[SO_3] = 2.70$ *M* , $K_{eq} = 1.87$

   $$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]} \qquad O_2 = \frac{[SO_3]^2}{[SO_2]^2K_{eq}} \qquad O_2 = \frac{[2.70]^2}{[1.60]^2 1.87} = 1.52\ M$$

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8. Name the 3 stresses that may be applied to a chemical equilibrium to cause it to shift.

_________temperature________ _________pressure or volume______ _____concentration___________

9. Describe LeChatelier's Principle.

When a system at equilibrium is subjected to a stress (a change in concentration, temperature, or pressure), the equilibrium will shift in the direction that tends to counteract the effect of the stress.

10. For the reaction below, mark whether the stress listed will cause the reaction to move forward or in reverse.

$$Zn(s) + 2HCl(l) \leftrightarrows ZnCl_2(l) + H_2(g);\ \Delta H = -235kJ$$

a. Increase Heat	_____	e. Increase $[ZnCl_2]$	_____	i. Decrease $[H_2]$	_____
b. Increase Pressure	_____	f. Increase [HCl]	_____	j. Decrease Pressure	_____
c. Increase $[H_2]$	_____	g. Decrease [HCl]	_____	k. Decrease Heat	_____
d. Increase [Zn]	_____	h. Decrease $[ZnCl_2]$	_____		

a. Forward reaction is exothermic (heat is a product), so increase in heat drives reaction in reverse.
b. Reverse reaction requires less volume (no gases produced), so reaction driven in reverse.
c. Reverse to consume added H_2.
d. Forward to consume Zn.
e. Reverse to consume $ZnCl_2$.
f. Forward to consume HCl.
g. Reverse to replace HCl.
h. Forward to replace $ZnCl_2$.
i. Forward to replace H_2.
j. Forward.
k. Forward.

11. What is the common-ion effect? Shift in equilibrium that occurs because the concentration of an ion that is part of the equilibrium is changed. For example, if you dissolve $MgCl_2$ in water, this is what happens:

$$MgCl_2(s) \leftrightarrows Mg^{2+}(aq) + 2Cl^{-}(aq)$$

Ions of both Mg^{2+} and Cl^{-} would be floating around in the water. If you then add a second solution that has NaCl [$NaCl(s) \leftrightarrows Na^{+}(aq) + Cl^{-}(aq)$], the added Cl^{-} would cause the above reaction to shift in reverse and $MgCl_2(s)$ would precipitate out of the solution. The "common ion" affects the equilibrium of one of the reactions (LeChatelier's Principle).

12. Solution formation also reaches a dynamic _____equilibrium____. The formation of a solution can be described using a _____dissociation____ equation. In these equations, the reactant is always a _____solid_____ and the products are always _____aqueous ions___. The solubility of a substance can be described using a ___solubility product____ expression, which is just an equilibrium expression for a dissolution reaction. In these expressions there will never be a _____denominator____________. Like K_{eq}, the value of K_{sp} is ___temperature____ dependent.

13. Write the balanced dissociation equations for the following salts.

a. $CaCl_2$ $CaCl_2(s) \leftrightarrows Ca^{2+}(aq) + 2Cl^{-}(aq)$

b. $(NH_4)_2SO_4$ $(NH_4)_2SO_4(s) \leftrightarrows 2NH_4^{+}(aq) + SO_4^{2-}(aq)$

c. NaI $NaI(s) \leftrightarrows Na^{+}(aq) + I^{-}(aq)$

d. $Al(NO_3)_3$ $Al(NO_3)_3(s) \leftrightarrows Al^{3+}(aq) + 3NO_3^{-}(aq)$

14. Q_{sp} is called the _____solubility product____. If $Q_{sp} > K_{sp}$, the solution is ___supersaturated______. If $Q_{sp} < K_{sp}$, the solution is _____unsaturated_____________. If $Q_{sp} = K_{sp}$, the solution is ___saturated __________.

15. Write the dissociation equations and solubility product expressions for the following:

a. $SrSO_4$

$SrSO_4(s) \leftrightarrows Sr^{2+}(aq) + SO_4^{2-}(aq)$

$K_{sp} = [Sr^{2+}][SO_4^{2-}]$

b. $Al_2(SO_4)_3$

$Al_2(SO4)_3(s) \leftrightarrows 2Al^{3+}(aq) + 3SO_4^{2-}(aq)$

$K_{sp} = [Al^{3+}]^2[SO_4^{2-}]^3$

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16. A sample of $PbBr_2(s)$ is added to pure water and allowed to come to equilibrium at 25°C. The concentration of Pb^{2+} is 0.0118*M* at equilibrium. What is the value of K_{sp} for $PbBr_2$? *(6.60 x 10⁻⁶)*

$$PbBr_2\ (s) \leftrightarrow Pb^{2+}(aq) + 2Br^{-}(aq)$$

$$K_{sp} = [Pb^{2+}][Br^{-}]^2$$

The concentrations of Br^- will be 2x the concentration of Pb^{2+} because the mole ratio of Br^- to Pb^{2+} is 2:1

$$K_{sp} = [0.0118][0.0236]^2$$

$$Ksp = 6.6 \times 10^{-6}$$

17. A sample of $BaSO_4(s)$ is added to pure water and allowed to come to equilibrium at 25°C. The concentration of Ba^{2+} is 1.05×10^{-5}*M* at equilibrium. What is the value of K_{sp} for $BaSO_4$? *(1.1×10^{-10})*

$$BaSO_4(s) \leftrightarrow Ba^{2+}(aq) + SO_4^{2-}(aq)$$

$$K_{sp} = [Ba^{2+}][SO_4^{2-}]$$

The concentrations of Ba^{2+} and SO_4^{2-} will be equal because their mole ratio is 1:1

$$K_{sp} = [1.05 \times 10^{-5}][1.05 \times 10^{-5}]$$

$$Ksp = 1.10 \times 10^{-10}$$

18. What will be the equilibrium concentration of dissolved ions in a saturated solution of $Pb(OH)_2$ at 25°C? K_{sp} for the reaction is 1.2×10^{-15}. *($[Pb^{2+}] = 6.69 \times 10^{-6}M$ and $[OH^-] = 1.34 \times 10^{-5}M$)*

$$Pb(OH)_2(s) \leftrightarrow Pb^{2+}(aq) + 2OH^{-}(aq)$$

$$K_{sp} = [Pb^{2+}][OH^{-}]^2$$

Setting the concentration of Pb^{2+} to *x*, and the concentration of OH^- to *2x*,
(the mole ratio of Pb^{2+} to OH^- is 1:2), we get the following:

$$1.2 \times 10^{-15} = [x][2x]^2$$

$$1.2 \times 10^{-15} = 4x^3$$

$$3.0 \times 10^{-16} = x^3$$

$$6.69 \times 10^{-6} = x$$

So, $[Pb^{2+}]$ is $6.69 \times 10^{-6}M$ and $[OH^-]$ would be twice that or $1.34 \times 10^{-5}M$.

19. What will be the equilibrium concentration of dissolved ions in a saturated solution of $SrSO_4$ at 25°C? K_{sp} for the reaction is 3.44×10^{-7}. *($[Sr^{2+}] = [SO_4^{2-}] = 5.87 \times 10^{-4}M$)*

$$SrSO_4(s) \leftrightarrow Sr^{2+}(aq) + SO_4^{2-}(aq)$$

$$K_{sp} = [Sr^{2+}][SO_4^{2-}]$$

Setting the concentration of Sr^{2+} to *x*, and the concentration of SO_4^{2-} also to *x*,
(the mole ratio of Sr^{2+} to SO_4^{2-} is 1:1), we get the following:

$$3.44 \times 10^{-7} = [x][x]$$

$$3.44 \times 10^{-7} = x^2$$

$$5.87 \times 10^{-4} = x$$

So, $[Sr^{2+}]$ is $5.87 \times 10^{-4}M$ and $[SO_4^{2-}]$ would be the same.