Gr. 12 Chemistry

Exam Review

Solutions

UNIT 1 – KINETICS

1. List the factors that affect the rate of a reaction and state what affect each has.
   1. ***Nature of Reactants*** *– reactions that involve ionic compounds and simple ions are usually faster than reactions involving molecular compounds.*
   2. ***Surface Area*** *– a larger surface area allows more of the reactants to be in contact with the other reactants which leads to more collisions which increases the rate.*
   3. ***Concentration*** *– increase in concentration increases the reaction rate because there are more molecules in the reactant that can collide.*
   4. ***Temperature*** *– increasing the temperature increases the amount of molecules that have sufficient energy to react which increases the reaction rate.*
   5. ***Catalyst*** *– lowers the activation energy of a reaction, allowing more molecules to react which increases the reaction rate.*
   6. ***Pressure*** *– as the pressure increases, the concentration of the reactants will increase allowing for more collisions which increases the reaction rate.*
   7. ***Volume*** *– as the volume decreases, the concentration of the reactants will increase allowing for more collisions which increases the reaction rate.*

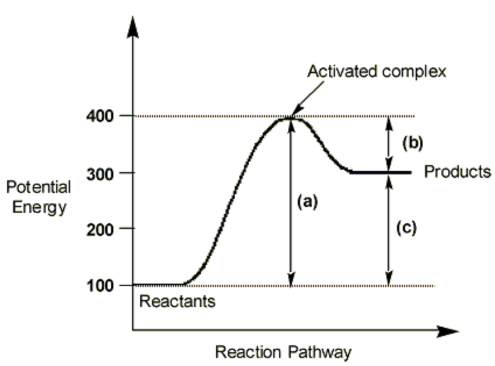
2. Using the collision theory, explain how the temperature of a reaction mixture influences the reaction rate. (Give two reasons)

   By increasing the temperature you are increasing the K.E. of the system. This speeds up the molecules, which causes more collisions. The more collisions, the faster the rate. By heating up the system, more molecules have sufficient energy to react which increases the reaction rate.

3. Using the collision theory, explain how the surface area of contact between reactant phases in a heterogeneous reaction influences the rate of reaction. Provide an example.

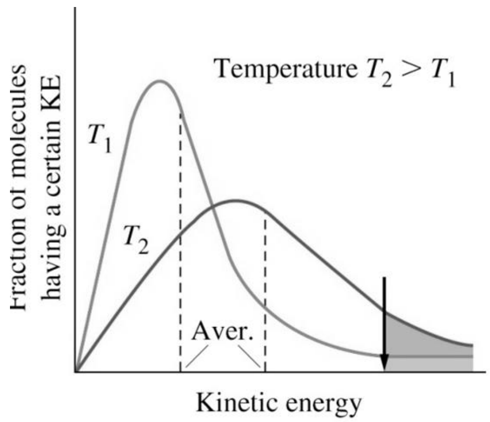
   By having a larger surface area, you are increasing the amount of collisions that can occur. One example would be the reaction of HCl with $CaCO_3$. The powder of calcium carbonate reacted faster than a solid piece because there was more surface available for contact, which increased the collisions.

4. Label the following potential energy (PE) diagram:



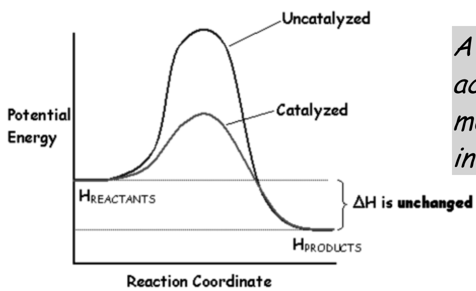
a.) Activation energy of forward reaction
b.) Activation energy of the reverse reaction
c.) Change in enthalpy of forward reaction

5. Using a kinetic energy (KE) diagram, explain the effect of increasing the temperature on the rate of reaction.

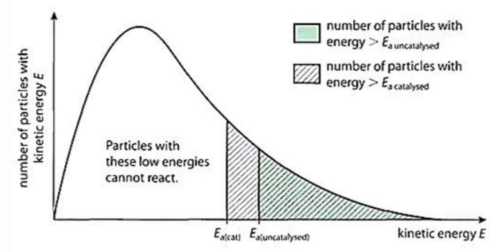


Increasing the temperature means that more particles have sufficient energy to react.

6. Explain the effect of a catalyst on the rate of a chemical reaction using a reaction coordinate diagram as well as a kinetic energy distribution curve.



A catalyst will lower the activation energy, allowing more molecules to react increasing the reaction rate



7. Explain the concept of a reaction mechanism. What is the rate-determining step?
Reaction Mechanism *is a sequence of steps by which a reaction occurs at the molecular level.*
Rate determining step *is the slowest of the elementary steps in a reaction, and it determines the rate of the reaction.*

8. Define entropy and enthalpy.
Entropy is the degree of disorder in a system
Enthalpy is the measured heat content of a system

9. Given the following experimental data, use the method of initial rates to determine the rate law and rate constant for the reaction $2ClO_{(aq)} + 2OH^-_{(aq)} \rightarrow ClO_3^-{}_{(aq)} + ClO_2^-{}_{(aq)} + H_2O_{(l)}$

Initial $[ClO_2]$ (M)	**Initial $[OH^-]$ (M)**	**Initial Rate (mol/L·min)**
0.0500	0.200	6.90
0.100	0.200	27.6
0.100	0.100	13.8

rate = $k[ClO_2]^2[OH^-]$

$$k = \frac{\text{rate}}{[ClO_2]^2[OH^-]} = \frac{6.90}{(0.05)^2(0.2)} = 13\,800\ L^2/mol^2s$$

UNIT 2 – EQUILIBRIUM PART I

1. Define chemical equilibrium at the macroscopic and microscopic levels.
   ***Macroscopic (what you see) Definition**: A reaction occurring in a closed system, all reactants and products are present and the observable properties remain the same*
   ***Microscopic (what you can't see) Definition:** The reactants are forming products at the same rate as the products are forming reactants.*

2. Identify the conditions required for chemical equilibrium.
   Constant observable macroscopic properties
   A closed system
   Constant temperature and pressure
   Reversibility

3. What statement is TRUE about a system at chemical equilibrium?
   a) observable changes occur during equilibrium
   b) the [ ]'s of reactants and products are equal
   c) the forward and reverse reaction rates are equal
   d) there are no reactions during equilibrium

4. 2.00 moles of NOCl are placed in a 2.00 L container. The following reaction occurred:

$$2NOCl_{(s)} \leftrightarrows 2NO_{(g)} + Cl_{2(g)}$$

At equilibrium, the [NOCl] was 0.34 mol/L. What is the Keq. [NOCl] = 2mol/2L = 1 mol/L

	2 NOCl	⇆	2 NO	+	Cl_2
I	1		0		0
C	- 2X		+ 2X		+ X
E	0.34		0.66		0.33

$K_{eq} = [NO]^2[Cl_2]$
$K_{eq} = (0.66)^2(0.33)$
$K_{eq} = 0.14$

1 – 2X = 0.34
X = 0.33

5. a) Write the equilibrium constant expression for the following balanced equation:

$$2SO_{3(g)} + 188.1\text{ kJ} \leftrightarrows 2SO_{2(g)} + O_{2(g)}$$

$$K_{eq} = \frac{[SO_2]^2[O_2]}{[SO_3]^2}$$

   b) What can be done to maximize the amount of $SO_{2(g)} + O_{2(g)}$ produced.
   increase volume, decrease pressure, increase temperature, increase [SO_3], decrease either/both [SO_2] and [O_2]

6. Consider the following equilibrium equation:

$$CO_{(g)} + H_2O_{(g)} \leftrightarrows CO_{2(g)} + H_{2(g)} \qquad \Delta H = -41\text{ kJ}$$

   Which of the following would cause a shift to the left?
   a) increasing the temperature
   b) adding CO
   c) removing H_2
   d) increasing pressure

7. Consider the following aqueous reaction at equilibrium:

$$2CrO_4^{2-} + 2H_3O^+ \leftrightarrows Cr_2O_7^{2-} + 3H_2O$$

For the above reaction, increasing the $[CrO_4^{2-}]$ will:

a) shift the equilibrium to the left
b) increase the $[Cr_2O_7^{2-}]$
c) increase the $[H_3O^+]$
d) cause no change in the position of equilibrium

8. Consider the following equilibrium reaction for an indicator in water:

$$\underset{\text{red}}{HMo_{(aq)}} + H_2O_{(aq)} \leftrightarrows H_3O^+_{(aq)} + \underset{\text{yellow}}{Mo^-_{(aq)}}$$

What is the effect of adding $NaOH_{(aq)}$ to the above system?

a) makes it more acidic
b) makes it more red
c) makes it more yellow
d) no change in the system

UNIT 2 –EQUILIBRIUM PART II

1. The solubility of calcium hydroxide can be represented by the equation: $Ca(OH)_{2(s)} \leftrightarrows Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$. What is the solubility product expression?

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

2. $Ca(OH)_{2(s)}$ is a sparingly soluble compound. If its solubility is 1.0×10^{-2} mol/L, what is the value of the Ksp?

   $Ca(OH)_2 \leftrightarrows Ca^{2+} + 2OH^-$
   X 2X

   $K_{sp} = [Ca^{2+}][OH^-]^2$
   $K_{sp} = [1x10^{-2}][2\ (1x10^{-2})]^2$
   $K_{sp} = 4x10^{-6}$

3. In a saturated solution of $SrCO_3$, the $[CO_3^{2-}] = 3.0 \times 10^{-5}$ mol/L. What is the value of Ksp for $SrCO_3$?

   $SrCO_{3(s)} \rightarrow Sr^{2+} + CO_3^{2-}$

   $K_{sp} = [Sr^{2+}][CO_3^{2-}]$
   $K_{sp} = (3x10^{-5})(3x10^{-5})$
   $K_{sp} = 9x10^{-10}$

4.. What is the Ksp value for $AgBrO_3$ if 1.0 L of a saturated solution contains 7.3×10^{-3} mol of this salt at 25°C? $AgBrO_{3(s)} \leftrightarrows Ag^+_{(aq)} + BrO_3^-{}_{(aq)}$

   $K_{sp} = [Ag^+][BrO_3^-]$
   $K_{sp} = (7.3 \times 10^{-3})(\ 7.3 \times 10^{-3})$
   $K_{sp} = 5.3 \times 10^{-5}$

5. What is the solubility of MgF_2 if its Ksp value is 1.6×10^{-8}?

$K_{sp} = [Mg^{2+}][F^-]^2$
$1.6 \times 10^{-8} = (x)(2x)^2$
$1.6 \times 10^{-8} = 4x^3$
$x^3 = 4 \times 10^{-9}$
$x = 1.59 \times 10^{-3}$ ➔ solubility is 1.59×10^{-3} mol/L

6.. The solubility of CuCl is 5.7×10^{-4} mol/L. What is the Ksp of CuCl?

$K_{sp} = [Cu^+][Cl^-]$
$K_{sp} = (5.7 \times 10^{-4})(5.7 \times 10^{-4})$
$K_{sp} = 3.2 \times 10^{-7}$

7. The Ksp for barium sulfate is 1.1×10^{-10}. What is the solubility of barium sulfate?

$K_{sp} = [Ba^{2+}][SO_4^{2-}]$
$1.1 \times 10^{-10} = (x)(x)$
$1.1 \times 10^{-10} = x^2$
$x = 1.05 \times 10^{-5}$ ➔ solubility is 1.05×10^{-5} mol/L

8. The Ksp for $BaCO_3$ is 8.1×10^{-9}. If Ba^{2+} ions are added to a solution which has a $[CO_3^{2-}] = 2.0 \times 10^{-4}$ mol/L, what $[Ba^{2+}]$ ions will start to precipitate?

	$BaCO_{3\,(s)}$	⇆	Ba^{2+}	+	CO_3^{2-}
I	---		0		2×10^{-4}
C	---		+ X		+ X
E	---		X		2×10^{-4} + X

$K_{sp} = [Ba^{2+}][CO_3^{2-}]$
$8.1 \times 10^{-9} = (x)(\,2\times10^{-4} + x)$ (the "+ x" is struck through)
$8.1 \times 10^{-9} = 2\times10^{-4}\,x$
$x = 4.05 \times 10^{-5}$ ➔ $[Ba^{2+}] = 4.05 \times 10^{-5}$ mol/L

9. The next two questions refer to the following information: Consider the Ksp values for 4 different salts.

calcium sulfate	2.4×10^{-5}	lead sulfate	1.1×10^{-8}
silver chloride	1.8×10^{-10}	barium chromate	8.5×10^{-11}

a) Which salt is LEAST soluble? *Barium chromate*

b) If 100 mL of 2.0×10^{-4} mol/L lead(II) nitrate solution is mixed with 250 mL of 4.0×10^{-5} mol/L sodium sulfate solution, what is Q? Will a ppt form?

(0.100L) (2,0 × 10^{-4}mol/L) =2 × 10^{-5} mol Pb^{2+} (0.250L)(4,0 × 10^{-4} mol/L) = 1 × 10^{-5} mol SO_4^{2-}
Total volume = 0,100L + 0,250 L = 0,350 L
$[Pb^{2+}] = 5.7 \times 10^{-5}$ mol/L $[SO_4^{2-}] = 2.9 \times 10^{-5}$ mol/L

$Kps_{exp} = [Pb^{2+}][SO_4^{2-}]$
$Kps_{exp} = (5,7 \times 10^{-5})(2.9 \times 10^{-5})$
$Kps_{exp} = 1.7 \times 10^{-9}$ $Kps_{exp} < K_{sp}$ ∴ pas de précipité

10. The slightly soluble salt $BaCO_3$ is added to an aqueous solution of Na_2CO_3. Which statement is TRUE?

a) No $BaCO_3$ will dissolve in the solution.
b) The solubility of $BaCO_3$ will be reduced.
c) A saturated solution of $BaCO_3$ is impossible.
d) Na_2CO_3 will precipitate from the solution.

11. If 100 mL of 2.4×10^{-4} M $AgNO_3$ solution is added to 100 mL of 2.4×10^{-5} NaCl solution, determine whether or not a precipitate will form. (Ksp for AgCl = 1.7×10^{-10}). Use the ***correct significant figures*** when answering.

(0,100L) ($2,4 \times 10^{-4}$mol/L) =$2,4 \times 10^{-5}$ mol Ag^+ (0,100L)($2,4 \times 10^{-5}$ mol/L) = $2,4 \times 10^{-6}$ mol Cl^-
Total volume = 0,100L + 0,100 L = 0,200 L
$[Ag^+] = 1,2 \times 10^{-4}$ mol/L $[Cl^-] = 1,2 \times 10^{-5}$ mol/L

$Kps_{exp} = [Ag^+][Cl^-]$
$Kps_{exp} = (1,2 \times 10^{-4})(1,2 \times 10^{-5})$
$Kps_{exp} = 1,44 \times 10^{-9}$ $Kps_{exp} > K_{sp}$ ∴ précipité

12. Calculate the mass of $SrCO_3$ that will dissolve in 2.00 L of water to form a saturated solution. Ksp of $SrCO_3 = 1.00 \times 10^{-8}$.

$K_{sp} = [Sr^{2+}][CO_3^{2-}]$
$1.00 \times 10^{-8} = (x)(x)$
$1.00 \times 10^{-8} = x^2$
$x = 1 \times 10^{-4}$
$[SrCO_3] = 1 \times 10^{-4}$ mol/L

n = CV
n = $(1 \times 10^{-4})(2.00)$
n = 2×10^{-4} mol

$$2\times10^{-4}\text{mol} \times \frac{147.61 \text{ g}}{1 \text{ mol}} = 0.030 \text{ g}$$

UNIT 3 – ACIDS AND BASES

1. Define electrolyte and non-electrolyte
   electrolyte – substance that conducts electricity when dissolved in water
   non-electrolyte –substance that does not conduct electricity when dissolved in water

2. Explain the differences between a strong and weak electrolyte.

strong electrolytes:	*weak electrolytes:*
• *good conductor when dissolved*	• *poor conductor when dissolved*
• *dissociates 100%*	• *partially dissociates in water*
• *example: NaCl*	• *example: Acetic Acid*

3. Define and compare Arrhenius and Bronsted-Lowry acid/base theories.
   Arrhenius acid: substance that produces H^+
   Arrhenius base: substance that produces OH^-
   Bronsted-Lowry acid: proton donator
   Bronsted-Lowry base: proton acceptor

4. For the following reaction: $NH_4^+ + H_2O \leftrightharpoons NH_3 + H_3O^+$, state which are the acids.
   A B CB CA

5. Consider the reaction: $H_2PO_4^-{}_{(aq)} + HCO_3^-{}_{(aq)} \leftrightharpoons HPO_4^{2-}{}_{(aq)} + H_2CO_{3(aq)}$. The CORRECT statement is:
   a) HPO_4^{2-} reacts as a B.L. base
   b) H_2CO_3 reacts as a B.L. base
   c) HCO_3^- reacts as a B.L. acid

d) $H_2PO_4^-$ reacts as both a B.L. acid and a B.L. base.

6. What is the conjugate acid of H_2O? $H3O^+$

7. Name the two conjugate acid/base pairs in this system:

$$HSO_3^-{}_{(aq)} + H_2BO_3^-{}_{(aq)} \leftrightarrows SO_3^{2-}{}_{(aq)} + H_3BO_{3(aq)}.$$

A *B* *CB* *CA*

8. $HS^-{}_{(aq)}$ ions behave like a base when they react in the presence of $HSO_3^-{}_{(aq)}$ ions. According to the Bronsted-Lowry theory, which of the following equations represents this reaction?
   a) $HS^-{}_{(aq)} + HSO_3^-{}_{(aq)} \leftrightarrows S^{2-}{}_{(aq)} + H_2SO_{3(aq)}$
   b) $HS^-{}_{(aq)} + HSO_3^-{}_{(aq)} \leftrightarrows H_2S_{(aq)} + SO_3^{2-}{}_{(aq)}$
   c) $HS^-{}_{(aq)} + H_2SO_{3(aq)} \leftrightarrows H_2S_{(aq)} + HSO_{3(aq)}$
   d) $HS^-{}_{(aq)} + SO_3^{2-}{}_{(aq)} \leftrightarrows S^{2-}{}_{(aq)} + HSO_3^-{}_{(aq)}$

9. In the reaction: $H_2O + CO_3^{2-} \leftrightarrows HCO_3^- + OH^-$, the CO_3^{2-} is a Bronsted:
   a) acid that donates protons.
   b) base that accepts protons.
   c) acid that accepts protons.
   d) base that donates protons.
   e) none of the above.

10. When used to describe an acid, the word “weak” means that the acid:
   a) has a low pH
   b) has a low [ ]
   c) shows incomplete ionization
   d) is monoprotic

11. The property that acids and bases have in common is that they:
   a) conduct electricity
   b) turn red litmus blue
   c) turn blue litmus red
   d) taste sour

12. HX is a weak acid. In a 0.10 M $HX_{(aq)}$ solution, the species present in highest [ ] is:
   a) $HX_{(aq)}$
   b) $X_{(aq)}$
   c) $H_3O^+{}_{(aq)}$
   d) $OH^-{}_{(aq)}$

13. Which of the following statements is the correct description of strong acids?
   a) high $[H_3O^+]$ and high pH
   b) high $[H_3O^+]$ and low pH
   c) high $[OH^-]$ and high pH
   d) high $[OH^-]$ and low pH

14. What is the expression for Kw? $K_w = [H_3O^+][OH^-]$

15. The $[OH^-]$ in a solution is 1.0×10^{-4} mol/L. What is the pH of the solution?

$pOH = -\log[OH^-]$	$pH + pOH = 14$
$pOH = -\log(1 \times 10^{-4})$	$pH + 4 = 14$

pOH = 4 pH = 10

16. What is the $[H_3O^+]$ in a 0.20 mol/L NaOH solution?
    $[OH^-]$ = 0.20M *NaOH is a strong base ∴ dissociates 100%
    $K_w = [H_3O^+][OH^-]$
    $1.0 \times 10^{-4} = [H_3O^+][0.20]$
    $[H_3O^+] = 5 \times 10^{-14}$ mol/L

17. A solution has a $[H^+]$ of 0.0010 M. What is the $[OH^-]$?
    $K_w = [H_3O^+][OH^-]$
    $1.0 \times 10^{-4} = [0.0010][OH^-]$
    $[OH^-] = 1 \times 10^{-11}$ mol/L

18. What is the $[H_3O^+]$ in a 0.020 M $Ba(OH)_{2(aq)}$ solution?
    $[OH^-]$ = 0.04M
    $K_w = [H_3O^+][OH^-]$
    $1.0 \times 10^{-4} = [H_3O^+][0.04]$
    $[H_3O^+] = 2.5 \times 10^{-13}$ mol/L

19. A base added to a neutral solution will:
    a) decrease $[OH^-]$
    b) decrease $[H_3O^+]$
    c) decrease the pH of the solution
    d) donate protons to another substance in the solution

20. A 0.10 M solution containing a single dissolved substance is a very good conductor of electricity and turns blue litmus paper red. The pH of the solution is approximately:
    a) 1.4
    b) 5.6
    c) 7.0
    d) 12.3

21. A solution with a pH of 5.0 is:
    a) basic with $[OH^-] = 1.0 \times 10^{-5}$ M
    b) acidic with $[H_3O^+] = 1.0 \times 10^{-5}$ M
    c) acidic with $[OH^-] = 1.0 \times 10^{-5}$ M
    d) basic with $[H_3O^+] = 1.0 \times 10^{-5}$ M

22. A 0.100 M nitrous acid solution, $HNO_{2(aq)}$ has a $[H^+]$ of 5.0×10^{-3} M. What is the Ka for this acid?
    $HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{(5\times10^{-3})(5\times10^{-3})}{(0.100)} = 2.5\times10^{-4}$$

23. A 0.17 M solution of a weak base, MOH, has a $[OH^-]$ of 0.012 M. Calculate the % dissociation of the base.

$$\%\ \text{diss} = \frac{[OH^-]}{[MOH]} \times 100 = \frac{0.012}{0.17} \times 100 = 7.06\ \%$$

24. A weak acid has a Ka value of 6.5×10^{-10}. Calculate the $[H^+]$ in a 0.20 M solution.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \Rightarrow 6.5\times10^{-10} = \frac{(x)(x)}{(0.20)} \Rightarrow x^2 = 1.3\times10^{-10} \Rightarrow x = 1.14\times10^{-5}$$

$\therefore [H_3O^+] = 1.14\times10^{-5}$ mol/L

25. What is the Ka of a weak acid, HA, if an aqueous 0.3 mol/L solution contains 1.0×10^{-3} mol/L H_3O^+ ions?

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(1\times10^{-3})(1\times10^{-3})}{(0.3)} = 3.3\times10^{-6}$$

26. A solution is prepared by dissolving 14.4 g of a weak acid, HClO in enough water to produce 1.0 L of solution. Ka = 2.3×10^{-11}. Calculate the $[H_3O^+]$.

$$14.4\text{ g} \times \frac{1\text{ mol}}{52.45\text{ g}} = 0.27\text{ mol} \qquad [HClO] = 0.27\text{ mol/L}$$

	HClO	+	H_2O	⇆	H_3O^+	+	ClO^-
I	0.27				0		0
C	- X				+ X		+ X
E	0.27 - X				X		X

$$K_a = \frac{[H_3O^+][ClO^-]}{[HClO]} \Rightarrow 2.3\times10^{-11} = \frac{(x)(x)}{(0.27 \cancel{- x})} \Rightarrow x^2 = 6.21\times10^{-12} \Rightarrow x = 2.49\times10^{-6} \Rightarrow [H_3O^+] = 2.49\times10^{-6}\text{ mol/L}$$

27. A neutralization reaction is a reaction between:
   a) an acid and a base to produce a salt and water.
   b) an acid and a metal to produce a salt and hydrogen gas.
   c) two aqueous solutions in which a precipitate forms.
   d) an oxidizing agent and a reducing agent in which a metal is produced.

28. What volume of 0.150 mol/L H_2SO_4 is required to neutralize 30.0 mL of 0.250 mol/L KOH solution? $H_2SO_4 + 2\ KOH \rightarrow 2\ H_2O + K_2SO_4$

$$n = CV = (0.250)(0.030) = 0.0075\text{ mol KOH}$$

$$0.0075\text{ mol KOH} \times \frac{1\text{ mol } H_2SO_4}{2\text{ mol KOH}} = 0.00375\text{ mol } H_2SO_4$$

$$V = \frac{n}{C} = \frac{0.00375\text{ mol}}{0.150\text{ mol/L}} = 0.025\text{ L (25 mL)}$$

29. How many mL of 0.15 M NaOH solution is required to neutralize 36 mL of 0.50 M HNO_3 solution? $NaOH + HNO_3 \rightarrow H_2O + NaNO_3$

$$n = CV = (0.50)(0.036) = 0.018\text{ mol } HNO_3$$

$$0.018\text{ mol } HNO_3 \times \frac{1\text{ mol NaOH}}{1\text{ mol } HNO_3} = 0.018\text{ mol NaOH}$$

$$V = \frac{n}{C} = \frac{0.018\text{ mol}}{0.150\text{ mol/L}} = 0.12\text{ L (120 mL)}$$

31. A student titrating a 25.0 mL sample of NaOH requires 20.0 mL of 0.50 M H_2SO_4 to reach the equivalence point. The original [ ] of NaOH is:
   a) 0.28 M
   b) 0.40 M
   c) 0.56 M
   d) 0.80 M

32. What is a titration curve?
   A titration curve is a curve that plots the pH of a titration of an acid and base.

33. What is the difference between the end point and the equivalence point?
   Equivalence point: *the point at which the amount of standard acid or base solution added just neutralizes the unknown sample.*
   Endpoint: *the point in the titration at which an indicators desired colour forms, want the endpoint and equivalence point to coincide.*

UNIT 4 – ATOMIC STRUCTURE

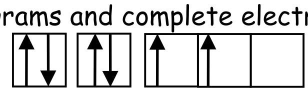
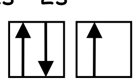
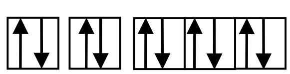
1. What is a quantum of energy?
   A quantum of energy is the minimum amount of energy that can be gained or lost by an atom

2. What is the continuous spectrum of white light?
   The continuous spectrum covers all wavelengths and frequencies in white light.

3. What is the atomic emission spectrum of an element?
   The atomic emission spectrum is a set of frequencies of electromagnetic waves given off by atoms of an element. It consists of a series of fine lines of individual colours.

4. What are the four blocks of the periodic table?
   The four blocks are s (left side), p (right side), d (middle), and f (bottom)

5. Which group has the most stable electron configuration? What is it?
   end p^6

6. Write orbital diagrams and complete electron configurations for the following elements:
   a. Carbon 
   1s $2s^2$ $2p^2$
   b. Lithium 
   $1s^2$ $2s^1$
   c. Neon 
   $1s^2$ $2s^2$ $2p^6$

7. Use noble-gas notation to describe the electron configurations of the elements represented by the following symbols:
   a. B = [He]$2s^2$ $2p^1$
   b. Mo = [Kr]$5s^2$ $4d^4$
   c. I = [Kr]$5s^2$ $4d^{10}$ $5p^5$

d. Gd = [Xe]$6s^2$ $4f^8$
e. Cd = [Kr] $5s^2$ $4d^{10}$
f. Pa = [Rn]$7s^2$ $5f^3$

8. What is a polar covalent bond?
   A polar bond is a bond between two non-metals where the electrons are not shared equally.

9. Explain what happens to the atomic radius as one moves across a row? Down a group? Why?
   Across the row, the atomic radii decreases, and this id due to an increase in the nuclear charge.
   Down the group, the atomic radii increases due to the increased size of the orbitals.

10. Explain what happens to the ionization energy as one moves across a row? Down a group? Why?
    Across the row, the ionization energy increases due to the increase in nuclear charge.
    Down the group, the ionization energy decreases due to the increase in atomic size.

UNIT 5 – ELECTROCHEMISTRY

1. Define oxidation and reduction in terms of:
   a) gain and loss of electrons
      oxidation – refers to a loss of electrons
      reduction – refers to a gain of electrons

   b) increase and decrease in oxidation number
      oxidation – element will have a negative number
      reduction – element will have a positive number

2. Be able to determine oxidation numbers for atoms in simple compounds and ions.
   (See oxidation rules)

3. Define and correctly use the terms oxidation, reduction, oxidizing agent, reducing agent.
   oxidizing agent – element that causes the oxidation of another.
   reducing agent – element that cause the reduction of another.

4. What is the oxidation number of nitrogen in $\mathbf{N}H_4OH$? *-3*

5. Which incomplete half-reaction is an oxidation reaction? (L.E.O.)
   a) $2H^+_{(aq)} + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$
   b) $Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$
   c) $K^+_{(aq)} \rightarrow K_{(s)}$
   d) $2I^-_{(aq)} \rightarrow I_{2(s)}$

6. In redox reactions,
   a) oxidizing agents lose electrons and are oxidized
   b) reducing agents lose electrons and are reduced
   c) oxidizing agents gain electrons and are reduced
   d) reducing agents gain electrons and are oxidized

7. An example of an oxidation-reduction reaction is:
   a) $H_3O^+_{(aq)} + HS^-_{(aq)} \rightarrow H_2S_{(aq)} + H_2O_{(l)}$
   b) $F^-_{(aq)} + HF_{(aq)} \rightarrow HF_2^-{}_{(aq)}$
   c) $2Br^-_{(aq)} + Cl_{2(g)} \rightarrow Br_{2(aq)} + 2Cl^-_{(aq)}$
   d) $2OH^-_{(aq)} + SO_{2(g)} \rightarrow SO_3^{2-}{}_{(aq)} + H_2O_{(l)}$

8. Which species is reduced in the following reaction:

$$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 3Sn^{2+}_{(aq)} \rightarrow 3Sn^{4+}_{(aq)} + 2Cr^{3+}_{(aq)} + 7H_2O_{(l)}$$

9. Which of the following reactions involves ***neither*** oxidation nor reduction?
   a) $2SO_3 \rightarrow 2SO_2 + O_2$
   b) $H_2O + Br_2 \rightarrow HBr + HOBr$
   c) $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$
   d) $CaO + SO_3 \rightarrow CaSO_4$

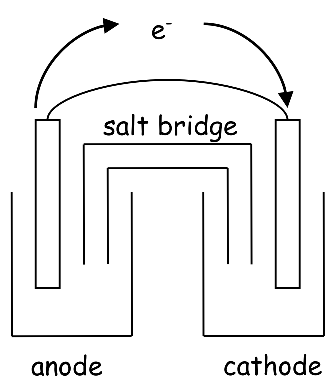
10. What is the oxidation number for the Mn in MnO_4^- ? *+7*

11. What is the coefficient for ClO_4^- when the following oxidation-reduction equation is balanced?

$$2Au_{(s)} + 3ClO_4^- + 12H^+ \leftrightarrows 3ClO_2^+ + 2Au^{3+} + 6H_2O$$

12. Describe how redox reactions generate an electric current.
Electrons will flow through a wire from the oxidation half-reaction to the reduction half-reaction while positive ions move through a salt bridge. The flow of electrons is the electric current.

13. Diagram and label an electrochemical (voltaic) cell – describing its components (i.e. cathode, anode, and salt bridge) and the direction of electron and ion flow.



14. Define anode, cathode, half-cell, and half-cell reaction.
anode – electrode where oxidation takes place
cathode – electrode where reduction takes place
half-cell – two parts of electrochemical cell in which the separate redox reactions take place
half-cell reaction – separate redox reactions

15. Differentiate between voltaic (galvanic) cells and electrolytic cells.

voltaic – produces electric energy

- *requires chemical reactions*
- *spontaneous*
- *require a salt bridge*
- *electrodes are involved in half reactions*
- *2 beakers*

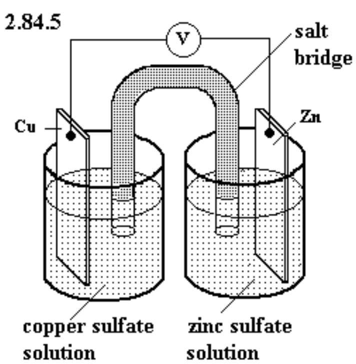
electrolytic – require electrical energy

- *produces chemical reactions*
- *no salt bridge in most cases*
- *electrodes are inert*
- *non-spontaneous*
- *1 beaker*

16. Reduction potentials are a measure of the tendency of:

a) $H^{+}_{(aq)}$ to gain electrons
b) a reducing agent to gain electrons
c) an oxidizing agent to lose electrons
d) an oxidizing agent to gain electrons

17. In this cell, which beaker will the cations flow into? (diagram)



The cations will flow from Cu to Zn
The electrons will flow from Zn to Cu

ll,

a) the strongest oxidizing agent reacts at the anode
b) gain of electrons occurs at the cathode
c) anions migrate towards the cathode
d) reduction occurs at the anode

19. Which reactants will result in a spontaneous reaction?

a) $Fe^{2+}_{(aq)} + Pb^{2+}_{(aq)}$
b) $Cr^{2+}_{(aq)} + Zn^{2+}_{(aq)}$
c) $Sn^{2+}_{(aq)} + I_{2(s)}$
d) $Na^{+}_{(aq)} + Pb_{(s)}$

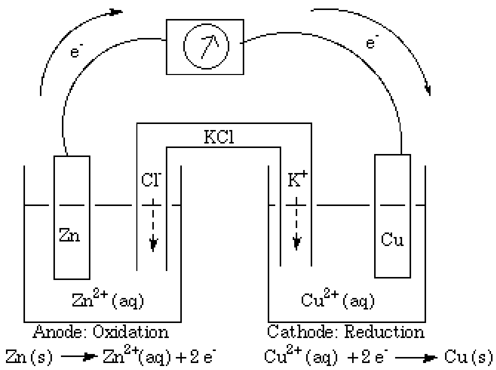
20. Given:

$Co^{2+} + 2e^{-} \rightarrow Co$ E° = -0.28 V
$Fe^{2+} + 2e^{-} \rightarrow Fe$ E° = -0.41 V

For the standard cell, $Fe/Fe^{2+}//Co^{2+}/Co$, what is the predicted cell voltage?

$Fe \rightarrow Fe^{2+} + 2e^{-}$	+0.41 V
$Co^{2+} + 2e^{-} \rightarrow Co$	-0.28 V
	+0.13 V

21. Refer to the following diagram to answer this question:



a) What is E° for the cell? *1.10 V*

b) Which of the following statements describes what is occurring in the zinc half-cell?

i. $Zn_{(s)}$ is oxidized, and the Zn electrode increases in mass.
ii. $Zn_{(s)}$ is oxidized, and the Zn electrode decreases in mass.
iii. $Zn^{2+}_{(aq)}$ is reduced, and the Zn electrode increases in mass.
iv. $Zn^{2+}_{(aq)}$ is reduced, and the Zn electrode decreases in mass.

An electrochemical cell is composed of chromium metal in a $Cr(NO_3)_{3(aq)}$ solution and magnesium metal in a $Mg(NO_3)_{2(aq)}$ solution. ***The next four questions refer to this cell.***

22. Balance the overall electrochemical equation for the reaction of this cell. When balanced, what is the coefficient of the Cr^{3+} in the net equation? *2*

$Mg/Mg^{2+}//Cr^{3+}/Cr$
$2Cr \rightarrow 6e^- + 2Cr^{3+}$
$3Mg^{2+} + 6e^- \rightarrow 3Mg$

$3Mg + 2Cr^{3+} \rightarrow 3Mg^{2+} + 2Cr$

23. Which material was the cathode?

a) $Cr_{(s)}$
b) $Cr(NO_3)_3$
c) $Mg_{(s)}$
d) $Mg(NO_3)_2$

24. The electrons in the external circuit move:

a) from $Mg_{(s)}$ to $Cr_{(s)}$
b) from $Cr_{(s)}$ to $Mg_{(s)}$
c) from Cr^{3+} to $Cr_{(s)}$
d) From Mg^{2+} to $Mg_{(s)}$

25. What is the E° for the reaction? *1.628 V*

$Mg \rightarrow Mg^{2+} + 2e^-$	+2.37 V
$Cr^{3+} + 3e^- \rightarrow Cr$	-0.74 V
	+1.63 V

26. A current of 8.00 A passes through molten $AlCl_3$ for 4.10 hrs. The number of moles of $Al_{(s)}$ deposited is:
   a) 0.0199
   b) 0.408
   c) 1.22
   d) 3.67

27. If an electrochemical cell produces 2.00 mol of electrons in 24.0 hrs, what is the amperage of this cell? *2.2A*

$$2\ \text{mol } e^- \times \frac{96500\ C}{1\ \text{mol } e^-} = 192\ 970\ C$$

$$Q = It$$
$$192\ 970 = I\ (86\ 400)$$
$$I = 2.2\ A$$

28. A 0.28 A current is passes through molten $MgBr_2$ for a period of 2.50 hrs.
   a) Write the half reaction which would occur at each electrode.

   $Mg^{2+} + 2e^- \rightarrow Mg$

   b) What mass of magnesium metal would be produced?

$$Q = It$$
$$Q = (0.28)(9000)$$
$$Q = 2520\ C$$

$$2\ 520\ C \times \frac{1\ \text{mol } e^-}{96500 C} \times \frac{1\ \text{mol Mg}}{2\ \text{mol } e^-} \times \frac{24.31\ g}{1\ \text{mol}} = 0.316\ g\ Mg$$

29. Consider the diagram below to answer the four parts of this question.
   a) Which electrode is the anode? *Zn*
   b) Indicate the direction of electron flow.
   c) At which electrode does reduction occur? *Cathode*
   d) Write the cathode half reaction. $Ni^{2+} + 2e^- \rightarrow Ni$