# Quarter 3 Review Worksheet-Accel

Name:_____ Period: _____

## UNIT 7- KMT and Gas Laws

1. What is an elastic collision?
2. How does the kinetic energy of particles change when temperature changes?
3. Use KMT to describe how gases are compressible and expandable.
4. What are the 3 assumptions of KMT?
5. Use your knowledge of gas properties to describe what is happening in the following situations:
    a. Your car tires appear flat in the morning after a cold night.

    b. Your ears "pop" as you are driving up a mountain and increasing in elevation at quick rate.

6. If I initially have a gas at a pressure of 10 atm, a volume of 15 liters, and a temperature of 300 K, and then I raise the pressure to 12 atm and decrease the temperature to 200 K, what is the new volume of the gas?

7. If I have 20 liters of gas at a temperature of 30 $^{\circ}$C and a pressure of 20 atm, what will be the pressure of the gas if I raise the temperature to 45$^{\circ}$C and decrease the volume to 15 liters?

8. What units do the following variables need to be in for the ideal gas law?
    a. pressure: _____
    b. volume: _____
    c. "n": _____
    d. temperature: _____
9. If I contain 20 g of chlorine gas in a container with a volume of 40 liters and at a temperature of 300 K, what is the pressure inside the container?

10. If I have 15 moles of gas held at a pressure of 10 atm and a temperature of 500 K, what is the volume of the gas?

## Unit 8 Thermochemistry
## Part 1: Enthalpy and Specific Heat

11. The symbol used to represent change in enthalpy is _____

12. When a chemical reaction gains heat, it is an _____ reaction, and its enthalpy value is (positive/negative)

13. When a substance goes from a solid to a liquid, it is an (endothermic/exothermic) process. The sign for ΔH is _____.

14. $H_2O$ (g) → $H_2O$(s) is an (endothermic/exothermic) process. The sign for ΔH is _____.

15. When a chemical reaction releases heat, it is an _____ reaction, and its enthalpy value is (positive/negative)

16. Describe the difference between temperature and heat.

17. The amount of energy required to raise the temperature of 1 gram of a substance 1$^{\circ}$C is called _____.
## Calculate the following and show your work!
18. What is the mass of a piece of iron that absorbs 1100 J of energy when its temperature changes from 25$^{\circ}$C to 175$^{\circ}$C? (specific heat of iron: 0.450 J/g $^{\circ}$C)

19. How many joules of heat are needed to raise the temperature of 15.0 g of aluminum from 20°C to 55°C, if the specific heat of aluminum is 0.90 J/g°C?

20. The products in an exothermic reaction have a (greater/smaller) enthalpy than the reactants.

21. How much heat will be released when 10.0g of sulfur reacts with excess $O_2$ according to the following equation?

$$2S + 3O_2 \rightarrow 2SO_3 \qquad\qquad \Delta H = -791.4 \text{ kJ}$$

22. Use the standard enthalpies of formation (pg. 975) to calculate the overall change in enthalpy for the reaction.

$$CH_4(g) + 2 O_2(g) \text{ ---> } CO_2(g) + 2 H_2O(l)$$
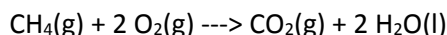
| | | | | | |
|---|---|---|---|---|---|
| H–H | 436 kJ/mol | C–H | 413 kJ/mol | C=C | 614 kJ/mol |
| H–Cl | 431 kJ/mol | C–C | 348 kJ/mol | C≡C | 839 kJ/mol |
| H–F | 567 kJ/mol | C–N | 293 kJ/mol | C=O | 799 kJ/mol |
| N–H | 391 kJ/mol | C–O | 358 kJ/mol | O=O | 495 kJ/mol |
| N–O | 201 kJ/mol | C–F | 485 kJ/mol | C≡O | 1072 kJ/mol |
| O–H | 463 kJ/mol | C–Cl | 328 kJ/mol | C=N | 615 kJ/mol |
| O–O | 146 kJ/mol | C–S | 259 kJ/mol | N=N | 418 kJ/mol |
| F–F | 155 kJ/mol | Cl–Cl | 242 kJ/mol | N≡N | 941 kJ/mol |
| | | | | C≡N | 891 kJ/mol |

23. Estimate the enthalpy change ($\Delta H_{rxn}$) using bond energies.
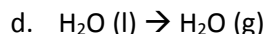
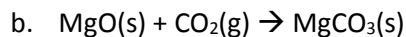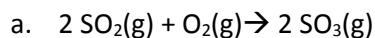$$H_2 + Cl_2 \rightarrow 2HCl$$
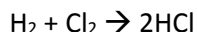
**Part 2: Entropy and Gibbs Free Energy**

24. Explain what a spontaneous process is.

25. Define entropy.

26. What is the symbol used to represent change in entropy? _____

27. Predict the sign on the change in entropy for the following equations:

a. $2 SO_2(g) + O_2(g) \rightarrow 2 SO_3(g)$ _____

b. $MgO(s) + CO_2(g) \rightarrow MgCO_3(s)$ _____

c. $2H_2O_2(l) \rightarrow H_2O(l) + O_2(g)$ _____

d. $H_2O(l) \rightarrow H_2O(g)$ _____

28. Copper (I) sulfide reacts with sulfur to produce copper (II) sulfide under standard conditions (298K). The process is exothermic ($\Delta H° = -26.7$ kJ/mol) with a decrease in disorder ($\Delta S° = -19.7$ J/(mol•K)). Determine the spontaneity of the reaction by calculating $\Delta G°$.

$$Cu2S (s) + 2 S (s) \rightarrow 2 CuS (s)$$
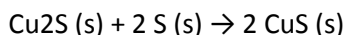
29. A reaction has $\Delta H° = -100.0$ kJ and $\Delta S° = -80.0$ J/K at 298 K. Is this reaction spontaneous?

**UNIT 9- Solutions**

30. What factors determine whether one substance will dissolve in another (solubility)?

31. What three factors influence the rate of solvation?

32. What are the two components of a solution? Define each component.

33. As the temperature increases, what happens to the solubility of a solid? Of a gas?

34. How is a supersaturated solution created?

35. Explain the meaning of the phrase "like dissolves like"?

36. How would you prepare 100.ml of a 0.500M $HNO_3$ solution if you have a 12.0M stock solution of $HNO_3$?

37. What volume of a 6.0M NaCl solution can be made from 3.51g of NaCl?

38. How many grams of solvent are necessary to dissolve 325g of lithium bromide at 50˚C if the solubility of LiBr is 203g/100g water at this temperature?

39. What is the molar concentration (molarity) of a 125ml solution made by dissolving 34.2g of sucrose, $C_{12}H_{22}O_{11}$, in water?

40. What mass of water is needed to make a 1.35$m$ solution with 8.20mol NaOH?

41. Calculate the molality of a solution of 50.0 g nickel (II) chloride in 100.0 g of water.

42. 12.5 g barium chloride dissolves in 250. ml of water. Calculate the mass percent concentration.

43. What is the new boiling point is 25.0 g of calcium chloride is dissolved in 500. ml of water?

44. Which of the following substances will affect the boiling point the most?  CH4, NaCl, or $MgCl_2$?