Name: Key Date: _____ Mods: _____

Unit 12: Ideal Gas Law → $PV = nRT$

1. How many moles of gas does it take to occupy 120 liters at a pressure of 2.3 atmospheres and a temperature of 340 K?

$n = ?$
$V = 120\ L$
$P = 2.3\ atm$
$T = 340\ K$

$$n = \frac{PV}{RT} = \frac{(2.3\ atm)(120\ L)}{(0.0821 \frac{L atm}{mol K})(340 K)}$$

$n = 9.89\ mols$

2. If I have a 50 liter container that holds 45 moles of gas at a temperature of 200° C, what is the pressure inside the container?

$V = 50\ L$
$n = 45\ mols$
$T = 200°C = 473K$
$P = ?$

$$P = \frac{nRT}{V} = \frac{(45\ mols)(0.0821 \frac{L atm}{mol K})(473 K)}{50\ L}$$

$P = 34.95\ atm$ or $3537.57\ kPa$

3. It is not safe to put aerosol canisters in a campfire, because the pressure inside the canisters gets very high and they can explode. If I have a 1.0 liter canister that holds 2 moles of gas, and the campfire temperature is 1400° C, what is the pressure inside the canister?

$V = 1.0\ L$
$n = 2\ mols$
$T = 1400°C = 1673K$
$P = ?$

$$P = \frac{nRT}{V} = \frac{(2\ mols)(0.0821 \frac{L atm}{mol K})(1673 K)}{1.0\ L}$$

$P = 274.7\ atm$ or $27805.26\ kPa$

4. How many moles of gas are in a 30 liter scuba canister if the temperature of the canister is 300 K and the pressure is 200 atmospheres?

$n = ?$
$V = 30\ L$
$T = 300\ K$
$P = 200\ atm$

$$n = \frac{PV}{RT} = \frac{(200\ atm)(30\ L)}{(0.0821 \frac{L atm}{mol K})(300 K)}$$

$n = 243.61\ mols$

5. I have a balloon that can hold 100 liters of air. If I blow up this balloon with 3 moles of oxygen gas at a pressure of 1 atmosphere, what is the temperature of the balloon?

$V = 100\ L$
$n = 3\ mols$
$P = 1\ atm$
$T = ?$

$$T = \frac{PV}{nR} = \frac{(1\ atm)(100\ L)}{(3\ mols)(0.0821 \frac{L atm}{mol K})}$$

$T = 406\ K$ or $133°C$

6. If 3.7 moles of propane are at a temperature of 28°C and are under 154.2 kPa of pressure, what volume does the sample occupy?

$n = 3.7$ mols
$T = 28°C = 301$ K
$P = 154$ kPa
$V = ?$

$$V = \frac{nRT}{P} = \frac{(3.7 \text{ mols})(8.31 \frac{\text{L kPa}}{\text{mol K}})(301 \text{ K})}{154.2 \text{ kPa}}$$

$V = 60.02$ L

7. A sample of carbon monoxide at 57°C and under 0.67 atm of pressure takes up 85.3 L of space. What mass of carbon monoxide is present in the sample?

$T = 57°C = 330$ K
$P = 0.67$ atm
$V = 85.3$ L
$n = ?$

$$n = \frac{PV}{RT} = \frac{(0.67 \text{ mols})(85.3 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{mol K}})(330 \text{ K})}$$

$$n = 2.11 \text{ mols CO} \left| \frac{28 \text{ g CO}}{1 \text{ mol CO}} \right. = 59.08 \text{ g CO}$$

8. At –45°C, 71 g of fluorine gas take up 6843 mL of space. What is the pressure of the gas, in kPa?

↳ diatomic; F_2

$T = -45°C = 228$ K
$V = 6843 \text{ mL} \left| \frac{1 \text{ L}}{1000 \text{ mL}} \right. = 6.843$ L
$P = ?$ kPa
$n = 71 \text{ g } F_2 \left| \frac{1 \text{ mol } F_2}{38 \text{ g } F_2} \right. = 1.868$ mols

$$P = \frac{nRT}{V} = \frac{(1.868 \text{ mols})(8.31 \frac{\text{L kPa}}{\text{mol K}})(228 \text{ K})}{6.843 \text{ L}}$$

$P = 517.21$ kPa

9. At 971 mm Hg, 145 g of carbon dioxide have a volume of 34.13 dm^3. What is the temperature of the sample, in °C?

$T = ?$
$P = 971 \text{ mm Hg} \left| \frac{1 \text{ atm}}{760 \text{ mm Hg}} \right. = 1.278$ atm
$V = 34.13 \text{ dm}^3 \left| \frac{(10 \text{ cm})^3}{(1 \text{ dm})^3} \right| \frac{1 \text{ mL}}{1 \text{ cm}^3} \left| \frac{1 \text{ L}}{1000 \text{ mL}} \right. = 34.13$ L
$n = 145 \text{ g } CO_2 \left| \frac{1 \text{ mol } CO_2}{44 \text{ g } CO_2} \right. = 3.295$ mols

$$T = \frac{PV}{nR} = \frac{(1.278 \text{ atm})(34.13 \text{ L})}{(3.295 \text{ mols})(0.0821 \frac{\text{L atm}}{\text{mol K}})}$$

$T = 161.24$ K $= -111.76°C$

↳ $1 \text{ mL} = 1 \text{ cm}^3$ & $1 \text{ L} = 1 \text{ dm}^3$

10. At 137°C and under a pressure of 3.11 atm, a 276 g sample of an unknown noble gas occupies 13.46 L of space. What is the identity of the gas?

$T = 137°C = 410$ K
$P = 3.11$ atm
$n = ?$
$V = 13.46$ L

$$n = \frac{PV}{RT} = \frac{(3.11 \text{ atm})(13.46 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{mol K}})(410 \text{ K})}$$

$n = 1.244$ mols

$$\text{molar mass} = \frac{\text{mass}}{\text{mols}} = \frac{276 \text{ g}}{1.244 \text{ mols}} = 221.86 \text{ g/mol}$$

Radon gas.